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CHEMISTRY FORM 1 complete

CHEMISTRY FORM ONE NOTES

Introduction to chemistry

Chemistry is a branch of Science. Science is basically the study of living and non-living things. The branch of science that study living things is called Biology. The branch of science that study non-living things is called Physical Science. Physical Science is made up of:

(i) Physics- the study of matter in relation to energy
(ii) Chemistry- the study of the composition of matter.

Chemistry is thus defined as the branch of science that deals with the structure composition, properties and behavior of matter.

Basic Chemistry involves studying:

(a) States/phases of matter

Matter is anything that has weight/**mass** and occupies space/**volume.** Naturally, there are basically **three** states of matter.

(i) **Solid**-e.g. soil, sand, copper metal, bucket, ice.
(ii)**Liquid-** e.g. water, Petrol, ethanol/alcohol, Mercury (liquid metal).
(iii) **gas-** e.g. Oxygen, Nitrogen ,Water vapour.

A solid is made up of particles which are very closely packed. It thus has a definite/fixed shape and fixed/definite volume /occupies definite space. It has a very high density.

A liquid is made up of particles which have some degree of freedom. It thus has no definite/fixed shape. It takes the shape of the container it is put. A liquid has fixed/definite volume/occupies definite space.

A gas is made up of particles free from each other. It thus has no definite/fixed shape. It takes the shape of the container it is put. It has no fixed/definite volume/occupies every space in a container.

(b) Separation of mixture

A mixture is a combination of two or more substances that can be separated by physical means. Simple methods of separating mixtures at basic chemistry level include:

i) Sorting/picking-this involve physically picking one pure substance from a mixture with another/other. e. g. sorting maize from maize beans mixture.

ii) **Decantation-**this involve pouring out a liquid from a solid that has settled /sinking solid in it. e. g. Decanting water forms sand**.**

iii)**Filtration**-this involves sieving /passing particles of a mixture through a filter containing small holes that allow smaller particle to pass through but do not allow bigger particle to pass through.

iv) **Skimming-**this involve scooping floating particles. E.g. cream from milk

(c) Metals and non-metals

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Metals are shiny, ductile(able to form wires), malleable(able to form sheet) and coil without breaking. E.g. Iron, gold, silver, copper. Mercury is the only **liquid metal** known.

Non-metals are dull, not ductile (do not form wires), not malleable (do not form sheet) and break on coiling/brittle. E.g. Charcoal, Sulphur, pla-stics.

(d) Conductors and non-conductors

A conductor is a solid that allow electric current to pass through. A non-conductor is a solid that do not allow electric current to pass through.

All metals conduct electricity. All non-metals do not conduct electricity except carbon **graphite**.

(e) Drugs

A drug is a natural or synthetic/man-made substance that when taken changes/alter the body functioning. A natural or synthetic/man-made substance that when taken changes/alter the abnormal body functioning to normal is called **medicine**.

Medicines are thus drugs intended to correct abnormal body functions. . Medicines should therefore be taken on **prescription** and **dosage**.

A prescription is a medical instruction to a patient/sick on the correct type of medicine to take and period/time between one intake to the other.

A dosage is the correct quantity of drug required to alter the abnormal body function back to normal. This is called **treatment**. It is the professional work of qualified doctors/pharmacists to administer correct prescription and dosage of drugs/medicine to the sick. Prescription and dosage of drugs/medicine to the sick use medical language.

Example

(i) **2 x 4** ; means "**2**" tablets for **solid** drugs/spoonfuls for **liquid** drugs taken "**4**" times for a duration of one day/24 hours and then repeated and continued until all the drug given is finished.

(ii) **1 x 2** ; means "**1**" tablets for **solid** drugs/spoonfuls for **liquid** drugs taken "**2**" times for a duration of one day/24 hours and then repeated and continued until all the drug given is finished.

Some drugs need minimal prescription and thus are available without pharmacist/ doctor's prescription. They are called **O**ver **T**he **C**ounter (OTC) drugs. OTC drugs used to treat mild headaches, stomach upsets, common cold include:

(i) Painkillers
(ii) Anti-acids
(iii) cold/flu drugs.

All medicine requires correct intake dosage. When a prescription dosage is not followed, this is called drug **misuse/abuse.** Some drugs are used for other purposes other than that intended. This is called **drug abuse**.

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Drug abuse is when a drug is intentionally used to alter the normal functioning of the body. The intentional abnormal function of the drug is to make the victim have false feeling of well being. The victim lack both mental and physical coordination. Some drugs that induce a false feeling of well being are illegal. They include heroin, cocaine, bhang, Mandrax and morphine.
Some abused drugs which are not illegal include: Miraa, alcohol, tobacco, sleeping pills.

The role of chemistry in society

(a) Chemistry is used in the following:
(i) Washing/cleaning with soap:
Washing/cleaning is a chemical process that involves interaction of water, soap and dirt so as to remove the dirt from a garment.
(ii) Understanding chemicals of life
Living thing grow, respire and feed. The formation and growth of cells involve chemical processes in living things using carbohydrates, proteins and vitamins.
(iii) Baking:
Adding baking powder to dough and then heating in an oven involves interactions that require understanding of chemistry.
(iv) Medicine:
Discovery, test, prescription and dosage of drugs to be used for medicinal purposes require advanced understanding of chemistry
(v) Fractional distillation of crude oil:
Crude oil is fractional distilled to useful portions like petrol, diesel, kerosene by applying chemistry.
(vi) Manufacture of synthetic compounds/substances
Large amounts of plastics, glass, fertilizers, insecticides, soaps, cements, are manufactured worldwide. Advanced understanding of the chemical processes involved is a requirement.
(vii) Diagnosis/test for abnormal body functions.
If the body is not functioning normally, it is said to be sick/ill. Laboratory test are done to diagnose the illness/sickness.

(b) The following career fields require Chemistry as one of subject areas of advanced/specialized study:

(i) Chemical engineering/chemical engineer
(ii) Veterinary medicine/Veterinary doctor
(iii) Medicine/Medical doctor/pharmacist/nurse
(iv) Beauty/Beautician
(v) Teaching/Chemistry teacher.

The School Chemistry Laboratory

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Chemistry is studied mainly in a science room called a school chemistry **laboratory**. The room is better ventilated than normal classroom. It has electricity, gas and water **taps**. A school chemistry laboratory has a qualified professional whose called Laboratory technician/assistant.

All students user in a school chemistry laboratory must consult the Laboratory technician/assistant for all their laboratory work. A school chemistry laboratory has chemicals and apparatus.

A chemical is a substance whose composition is known. All chemical are thus labeled as they are. This is because whereas physically a substance may appear similar, chemically they may be different.

All Chemicals which are not labeled should never be used. Some chemicals are toxic/poisonous, explosive, corrosive, caustic, irritants, flammable, oxidizing, carcinogenic, or radioactive.

Care should always be taken when handling any chemical which have any of the above characteristic properties.

Common school chemistry laboratory chemicals include:

(i) Distilled water

(ii) Concentrated mineral acid which are very corrosive (on contact with skin they cause painful open wounds)

(iii) Concentrated alkali/bases which are caustic (on contact with skin they cause painful blisters)

(iv) Very many types of salts

The following safety guideline rules should be followed by chemistry laboratory users:

(i) Enter the laboratory with permission in an orderly manner without rushing/pushing/scrabbling.

(ii) Do not try unauthorized experiments. They may produce flammable, explosive or toxic substances that affect your health.

(iii) Do not taste any chemical in the laboratory. They may be poisonous.

(iv) Waft gas fumes to your nose with your palm. Do not inhale/smell gases directly. They may be highly poisonous/toxic.

(v) Boil substances with mouth of the test tube facing away from others and yourself. Boiling liquids spurt out portions of the hot liquid. Products of heating solids may be a highly poisonous/toxic gas.

(vi) Wash with lots of water any skin contact with chemicals immediately. Report immediately to teacher/laboratory technician any irritation, cut, burn, bruise or feelings arising from laboratory work.

(vii) Read and follow safety instruction. All experiments that evolve/produce poisonous gases should be done in the open or in a fume chamber.

(viii)Clean your laboratory work station after use. Wash your hand before leaving the chemistry laboratory.

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(ix) In case of fire, remain calm, switch of the source of fuel-gas tap. Leave the laboratory through the emergency door. Use fire extinguishers near the chemistry laboratory to put of medium fires. Leave strong fires wholly to professional fire fighters.

(x) Do not carry unauthorized item from a chemistry laboratory.

An apparator /apparatus are scientific tools/equipment used in performing scientific experiments. The conventional apparator used in performing scientific experiments is called **standard** apparator/apparatus. If the conventional standard apparator/apparatus is not available, an **improvised** apparator/apparatus may be used in performing scientific experiments. An improvised apparator/apparatus is one used in performing a scientific experiment **for** a standard apparator/apparatus. Most standard apparatus in a school chemistry laboratory are made of **glass** because:

(i)Glass is transparent and thus reactions /interactions inside are clearly visible from outside

(ii) Glass is comparatively cheaper which reduces cost of equipping the school chemistry laboratory

(iii) Glass is comparatively easy to clean/wash after use.

(iv) Glass is comparatively unreactive to many chemicals.

Apparatus are designed for the purpose they are intended in a school chemistry laboratory:

(a) Apparatus for measuring volume

1. Measuring cylinder

Measuring cylinders are apparatus used to measure volume of liquid/ solutions. They are calibrated/ graduated to measure any volume required to the maximum. Measuring cylinders are named according to the maximum calibrated/graduated volume e.g.

"10ml" measuring cylinder is can hold maximum calibrated/graduated volume of "10mililitres" /"10 cubic centimetres"

"50ml" measuring cylinder is can hold maximum calibrated/graduated volume of "50mililitres" /"50 cubic centimetres"

"250ml" measuring cylinder is can hold maximum calibrated/graduated volume of "250mililitres" /"250 cubic centimetres"

"1000ml" measuring cylinder is can hold maximum calibrated/graduated volume of "1000mililitres" /"1000 cubic centimetres"

2. Burette

Burette is a long and narrow/thin apparatus used to measure small accurate and exact volumes of a liquid solution. It must be clamped first on a stand before being used. It has a tap to run out the required amount out. They are calibrated/ graduated to run out small volume required to the maximum 50ml/50cm3.

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The maximum 50ml/50cm3 calibration/ graduation reading is at the **bottom** .This ensure the amount run **out** from a tap **below** can be determined directly from **burette reading** before and after during volumetric analysis.
Burettes are expensive and care should be taken when using them.
3. (i) Pipette
Pipette is a long and narrow/thin apparatus that widens at the middle used to measure and transfer small very accurate/exact volumes of a liquid solution.
It is open on either ends.
The maximum 25ml/25cm3 calibration/ graduation mark is a visible **ring** on one thin end.
To fill a pipette to this mark, the user must suck up a liquid solution upto a level above the mark then adjust to the mark using a finger.
This requires practice.
(ii) Pipette filler
Pipette filler is used to suck in a liquid solution into a pipette instead of using the mouth. It has a suck, adjust and eject button for ensuring the exact volume is attained. This requires practice.
4. Volumetric flask.
A volumetric flask is thin /narrow but widens at the base/bottom. It is used to measure very accurate/exact volumes of a liquid solution.
The maximum calibration / graduation mark is a visible **ring.**
Volumetric flasks are named according to the maximum calibrated/graduated volume e.g.
"250ml" volumetric flask has a calibrated/graduated mark at exact volume of "250mililitres" /"250centimetres"
"1l" volumetric flask has a calibrated/graduated mark at exact volume of "one litre" /"1000 cubic centimeters"
"2l" volumetric flask has a calibrated/graduated mark at exact volume of "two litres" /"2000 cubic centimeters"
5. Dropper/teat pipette
A dropper/teat pipette is a long thin/narrow glass/rubber apparatus that has a flexible rubber head.
A dropper/teat pipette is used to measure very small amount/ drops of liquid solution by pressing the flexible rubber head. The numbers of drops needed are counted by pressing the rubber gently at a time
(b)Apparatus for measuring mass
1. Beam balance
A beam balance has a pan where a substance of unknown mass is placed. The scales on the opposite end are adjusted to "balance" with the mass of the unknown substance. The mass from a beam balance is in **grams**.
2. Electronic/electric balance.

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An electronic/electric balance has a pan where a substance of unknown mass is placed. The mass of the unknown substance in **grams** is available immediately on the screen.

(c)Apparatus for measuring temperature

A thermometer has alcohol or mercury trapped in a bulb with a thin enclosed outlet for the alcohol/mercury in the bulb.

If temperature rises in the bulb, the alcohol /mercury expand along the thin narrow enclosed outlet.

The higher the temperature, the more the expansion

Outside, a calibration /graduation correspond to this expansion and thus changes in temperature.

A thermometer therefore determines the temperature when the bulb is fully dipped in to the substance being tested. To determine the temperature of solid is thus very difficult.

(d)Apparatus for measuring time

The stop watch/clock is the standard apparatus for measuring time. Time is measured using hours, minutes and second.

Common school stop watch/clock has start, stop and reset button for determining time for a chemical reaction. This requires practice.

(e) Apparatus for scooping

1. Spatula

A spatula is used to **scoop** solids which do not require accurate measurement. Both ends of the spatula can be used at a time.

A solid scooped to the **brim** is "one spatula end full" A solid scooped to **half brim** is "half spatula end full".

2. Deflagrating spoon

A deflagrating spoon is used to **scoop** solids which do not require accurate measurement mainly for heating. Unlike a spatula, a deflagrating spoon is longer.

(f) Apparatus for putting liquids/solid for heating.

1. Test tube.

A test tube is a narrow/thin glass apparatus open on one side. The end of the opening is commonly called the "the mouth of the test tube".

2. Boiling/ignition tube.

A boiling/ignition tube is a wide glass apparatus than a test tube open on one side. The end of the opening is commonly called the "the mouth of the boiling/ignition tube".

3. Beaker.

Beaker is a wide calibrated/graduated lipped glass/plastic apparatus used for transferring liquid solution which do not normally require very accurate measurements

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Beakers are named according to the maximum calibrated/graduated volume they can hold e.g.
"250ml" beaker has a maximum calibrated/graduated volume of "250mililitres" /"250 cubic centimeters"
"1l" beaker has a maximum calibrated/graduated volume of "one litre" /"1000 cubic centimeters"
"5 l" beaker has a maximum calibrated/graduated volume of "two litres" /"2000 cubic centimeters"

4. Conical flask.

A conical flask is a moderately narrow glass apparatus with a wide base and no calibration/graduation. Conical flasks thus carry/hold exact volumes of liquids that have been measured using other apparatus. It can also be put some solids. The narrow mouth ensures no spillage.
Conical flasks are named according to the maximum volume they can hold e.g.
"250ml" Conical flasks hold a maximum volume of "250mililitres" /"250 cubic centimeters"
"500ml" Conical flasks hold a maximum volume of "500ml" /"1000 cubic centimeters"

5. Round bottomed flask

A round bottomed flask is a moderately narrow glass apparatus with a wide round base and no calibration/graduation. Round bottomed flask thus carry/hold exact volumes of liquids that have been measured using other apparatus. The narrow/thin mouth prevents spillage. The flask can also hold (weighed) solids. A round bottomed flask must be held/ clamped when in use because of its wide narrow base.

6. Flat bottomed flask

A flat bottomed flask is a moderately narrow glass apparatus with a wide round base with a small flat bottom. It has no calibration/graduation.
Flat bottomed flasks thus carry/hold exact volumes of liquids that have been measured using other apparatus. The narrow/thin mouth prevents spirage. They can also hold (weighed) solids. A flat bottomed flask must be held/ clamped when in use because it's flat narrow base is not stable.

(g) Apparatus for holding unstable apparatus (during heating).

1. Tripod stand

A tripod stand is a three legged metallic apparatus which unstable apparatus are placed on (during heating).Beakers. Conical flasks, round bottomed flask and flat bottomed flasks are placed on top of tripod stand (during heating).

2. Wire gauze/mesh

Wire gauze/mesh is a metallic/iron plate of wires crossings. It is placed on top of a tripod stand:

(i) Ensure even distribution of heat to prevent cracking glass apparatus
(ii) Hold smaller apparatus that cannot reach the edges of tripod stand

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3 Clamp stand
A clamp stand is a metallic apparatus which tightly hold apparatus at their "neck" firmly.
A clamp stand has a wide metallic base that ensures maximum stability. The height and position of clamping is variable. This require practice
4. Test tube holder
A test tube holder is a hand held metallic apparatus which tightly hold test/boiling/ignition tube at their "neck" firmly on the other end.
Some test tube holders have wooden handle that prevent heat conduction to the hand during heating.
5. Pair of tong.
A pair of tong is a scissor-like hand held metallic apparatus which tightly hold firmly a small solid sample on the other end.
6. Gas jar
A gas jar is a long wide glass apparatus with a wide base.
It is open on one end. It is used to collect/put gases.
This requires practice.

(h) Apparatus for holding/directing liquid solutions/funnels (to avoid spillage).

1. Filter funnel
A filter funnel is a wide mouthed (mainly plastic) apparatus that narrow drastically at the bottom to a long extension.
When the long extension is placed on top of another apparatus, a liquid solution can safely be directed through the wide mouth of the filter funnel into the apparatus without spirage.
Filter funnel is also used to place a filter paper during filtration.

2. Thistle funnel
A thistle funnel is a wide mouthed glass apparatus that narrow drastically at the bottom to a very long extension.
The long extension is usually drilled through a stopper/cork.
A liquid solution can thus be directed into a stoppered container without spillage

3. Dropping funnel
A dropping funnel is a wide mouthed glass apparatus with a tap that narrow drastically at the bottom to a very long extension.
The long extension is usually drilled through a stopper/cork.
A liquid solution can thus be directed into a stoppered container without spillage at the rate determined by adjusting the tap.

4. Separating funnel
A separating funnel is a wide mouthed glass apparatus with a tap at the bottom narrow extension.
A liquid solution can thus be directed into a separating funnel without spillage. It can also safely be removed from the funnel by opening the tap.

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It is used to separate two or more liquid solution mixtures that form layers/immiscible. This requires practice.

(h) Apparatus for heating/Burners

1. Candle, spirit burner, kerosene stove, charcoal burner/jiko are some apparatus that can be used for heating.

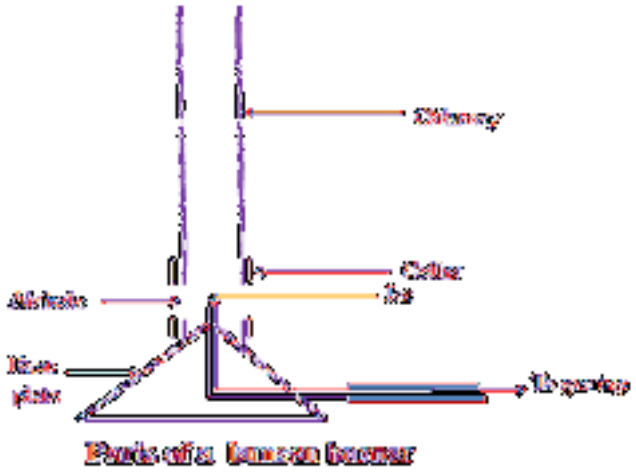
Any flammable fuel when put in a container and ignited can produce some heat.

2. Bunsen burner

The Bunsen burner is the **standard** apparatus for heating in a Chemistry school laboratory.

It was discovered by the German Scientist Robert Wilhelm Bunsen in1854.

(**a**)Diagram of a Bunsen burner



A Bunsen burner uses butane/laboratory gas as the fuel. The butane/laboratory gas is highly flammable and thus usually stored safely in a secure chamber outside Chemistry school laboratory. It is tapped and distributed into the laboratory through gas pipes.

The gas pipes end at the gas tap on a chemistry laboratory bench .If opened the gas tap releases butane/laboratory gas. Butane/laboratory gas has a characteristic odor/smell that alerts leakages/open gas tap.

The Bunsen burner is fixed to the gas tap using a strong rubber tube.

The Bunsen burner is made up of the following parts:

(i) Base plate –to ensure the burner can stand on its own

(ii)Jet-a hole through which laboratory gas enters the burner

(iii)Collar/sleeve-adjustable circular metal attached to the main chimney/burell with a side hole/entry. It controls the amount of air entering used during burning.

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(iv)Air hole- a hole/entry formed when the collar side hole is in line with chimney side hole. If the collar side hole is **not** in line with chimney side hole, the air hole is said to be "closed" If the collar side hole is **in line** with chimney side hole, the air hole is said to be "open"

(v)Chimney- tall round metallic rod attached to the base plate.

(b)Procedure for lighting/igniting a Bunsen burner

1. Adjust the collar to ensure the air holes are closed.
2. Connect the burner to the gas tap using a rubber tubing. Ensure the rubber tubing has no side leaks.
3. Turn on the gas tap.
4. Ignite the top of the chimney using a lighted match stick/gas lighter/wooden splint.
5. Do not delay excessively procedure (iv) from (iii) to prevent highly flammable laboratory gas from escaping/leaking.

(c)Bunsen burner flames

A Bunsen burner produces two types of flames depending on the amount of air entering through the air holes.

If the air holes are **fully open**, a **non luminous** flame is produced. If the air holes are **fully closed**, a **luminous flame** is produced. If the air holes are **partially** open/ closed, a **hybrid** of non luminous and luminous flames is produced.

Characteristic differences between luminous and non-luminous flame

Luminous flame	**Non-luminous flame**
1. Produced when the air holes are fully/completely **closed**.	**1.** Produced when the air holes are fully/completely **open**.
2. when the air holes are fully/ completely closed there is **incomplete** burning/ combustion of the laboratory gas	**2.**when the air holes are fully/ completely open there is **complete** burning/ combustion of the laboratory gas
3. Incomplete burning/ combustion of the laboratory gas produces fine unburnt carbon particles which make the flame **sooty/smoky**	**3.** Complete burning/ combustion of the laboratory gas does not produce carbon particles. This makes the flame non-**sooty /non- smoky.**
4. Some carbon particles become white hot and emit light. This flame is thus bright **yellow** in colour producing **light.** This makes luminous flame useful for **lighting**	**4.** Is mainly **blue** in colour and is **hotter** than luminous flame. This makes non-luminous flame useful for **heating**
5. Is **larger**, **quiet** and **wavy**/easily	**5.**Is **smaller**, **noisy** and **steady**

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swayed by wind	
Luminous flame has **three** main regions: (i)the top yellow region where there is incomplete combustion/burning (ii)the region of unburnt gas below the yellow region where the gas does not burn (iii) blue region on the sides of region of unburnt gas where there is complete burning	Non-luminous flame has **four** main regions: (i)the top colourless region (ii) Blue region just below where there is complete burning. It is the hottest region (iii) green region surrounded by the blue region where there is complete burning (Ii) The region of unburnt gas at the innermost surrounded by green and blue regions. No burning takes place here

Scientific apparatus are drawn:
(i) Using a proportional **two** dimension (**2D**) cross-sections. Three dimensions (3D) are not recommended.
(ii) Straight edges of the apparatus on a scientific diagram should be drawn using ruler.
(iii) Curved edges of the apparatus on a scientific diagram should be drawn using free hand.
(iv)The bench, tripod or clamp to support apparatus which cannot stand on their own should be shown.

CLASSIFICATION OF SUBSTANCES

Substances are either pure or impure. A pure substance is one which contains only one substance.
An impure substance is one which contains two or more substances. A pure substance is made up of a pure solid, pure liquid or pure gas.
A mixture is a combination of two or more pure substances which can be separated by physical means. The three states of matter in nature appear mainly as mixtures of one with the other.
Common mixtures include:

(a)Solutions/solid-liquid dissolved mixture

Experiment:

To make a solution of copper (II) sulphate (VI)/Potassium magnate(VII) /sodium chloride
Procedure

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Put about 100 cm3 of water in three separate beakers. Separately place a half spatula end full of copper (II) sulphate (VI), Potassium manganate (VII) and sodium chloride crystals to each beaker. Stir for about two minutes.
Observation
Copper (II) sulphate (VI) crystals dissolve to form a blue solution
Potassium manganate (VII) crystals dissolve to form a purple solution
Sodium chloride crystals dissolve to form a colourless solution
Explanation
Some solids, liquids and gases dissolve in some other liquids.
A substance/liquid in which another substance dissolves is called solvent.
A substance /solid /gas which dissolves in a solvent is called solute.
When a solute dissolves in a solvent it forms a uniform mixture called **solution**.
A solute dissolved in water as the solvent exists in another state of matter called **aqueous state.** Water is referred as the **universal solvent** because it dissolves many solutes. A solute that dissolves in a solvent is said to be **soluble**. Soluble particles uniformly spread between the particles of water/solvent and cannot be seen.

Solute + Solvent -> solution
Solute + Water -> aqueous solution of solute

The solute dissolved in water gives the **name** of the solution e. g.
1. Sodium chloride solution is a solution formed after dissolving sodium chloride crystals/solid in water. Sodium chloride exists in aqueous state after dissolving.

Sodium chloride + Water -> Sodium chloride solution
NaCl(**s**) + (aq) -> NaCl(**aq**)

2. Ammonia solution is a solution formed after dissolving ammonia gas in water. Ammonia exists in aqueous state after dissolving.

Ammonia gas + Water -> aqueous ammonia
NH_3 (**g**) + (aq) -> NH_3 (**aq**)

3. Copper (II) sulphate (VI) solution is a solution formed after dissolving Copper (II) sulphate (VI) crystals/solid in water. Copper (II) sulphate (VI) exists in aqueous state after dissolving.

Copper (II) sulphate (VI) + Water -> Copper (II) sulphate (VI) solution
$CuSO_4$(**s**) + (aq) -> $CuSO_4$ (**aq**)

4. Potassium manganate(VII) solution is a solution formed after dissolving Potassium manganate(VII) crystals/solid in water.
Potassium manganate(VII)exist in aqueous state after dissolving.
Potassium manganate(VII) + Water -> Potassium manganate(VII) solution
$KMnO_4$(**s**) + (aq) -> $KMnO_4$ (**aq**)

(b)Suspension/ precipitates/solid-liquid mixture which do not dissolve

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Experiment: To make soil, flour and Lead (II) Iodide suspension/precipitate
Procedure
Put about 100 cm3 of water in three separate beakers. Separately place a half spatula end full of soil, maize and lead (II) Iodide to each beaker. Stir for about two minutes.

Observation
Some soil, maize and lead (II) Iodide float in the water
A brown suspension/precipitate/particles suspended in water containing soil
A white suspension/precipitate/particles suspended in water containing flour
A yellow suspension/precipitate/particles suspended in water containing Lead (II) iodide. Some soil, maize and lead (II) Iodide settle at the bottom after some time.
Explanation
Some solid substances do not dissolve in a liquid. They are said to be **insoluble** in the solvent .When an insoluble solid is put in liquid:
(i) Some particles remain **suspended/floating** in the liquid to form a **suspension /precipitate**.
(ii) Some particles **sink/settle** to the bottom to form **sediments** after being allowed to stand.
An **insoluble** solid acquire the colour of the suspension/precipitate .e.g.
1. A white suspension /precipitate have some fine white particles suspended /floating in the liquid. **Not** "white solution"
2. A blue suspension /precipitate has some fine **blue** particles suspended /floating in the liquid.
3. A green suspension /precipitate has some fine **green** particles suspended /floating in the liquid.
4. A **brown** suspension /precipitate has some fine **brown** particles suspended /floating in the liquid.
4. A yellow suspension /precipitate has some fine yellow particles suspended /floating in the liquid.
(c) (i) Miscibles /Liquid-liquid mixtures
To form water-ethanol and Kerosene-turpentine miscibles
Procedure
(i)Measure 50cm3 of ethanol into 100cm3 beaker. Measure 50cm3 of water. Place the water into the beaker containing ethanol. Swirl for about one minute.
(ii)Measure 50cm3 of kerosene into 100cm3 beaker. Measure 50cm3 of turpentine oil. Place the turpentine oil into the beaker containing kerosene. Swirl for about one minute.
Observation
Two liquids do not form layers.
Ethanol and water form a uniform mixture.

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Kerosene and turpentine oil form uniform mixture

Explanation

Ethanol is miscible in Water. Kerosene is miscible in turpentine oil. Miscible mixture form uniform mixture. They do not form layers. The particles of one liquid are smaller than the particles of the other. The smaller particles occupy the spaces between the bigger particles.

(iii) Immiscibles /Liquid-liquid mixtures

To form water-turpentine oil and Kerosene-water miscibles

Procedure

(i)Measure 50cm3 of water into 100cm3 beaker. Measure 50cm3 of turpentine oil. Place the oil into the beaker containing water. Swirl for about one minute.

(ii) Measure 50cm3 of water into 100cm3 beaker. Measure 50cm3 of kerosene. Place the kerosene into the beaker containing water. Swirl for about one minute.

Observation

Two liquids form layers.

Turpentine and water do not form a uniform mixture.

Water and kerosene do not form uniform mixture

Explanation

Kerosene is immiscible in Water. Water is immiscible in turpentine oil. Immiscible mixtures do not form uniform mixtures. They form layers. The size of the particles of one liquid is almost equal to the particles of the other. The particles of one liquid cannot occupy the spaces between the particles of the other. The heavier particles settle at the bottom. The less dense particles settle on top.

(d)Solid-solid mixtures/Alloys

Before solidifying, some heated molten/liquid metals dissolve in another metal to form a uniform mixture of the two. On solidifying, a uniform mixture of the metals is formed. A uniform mixture of two metals on solidifying is called **alloy**. In the alloy, one metallic particle occupies the spaces between the metallic particles of the other.

c) Common alloys of metal.

Alloy name	**Constituents of the alloy**	**Uses of the alloy**
Brass	Copper and Zinc	Making screws and bulb caps
Bronze	Copper and Tin	Making clock springs, electrical contacts and copper coins
Soldier	Lead and Tin	Soldering, joining electrical contacts because of its low melting points and high thermal conductivity

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Duralumin	Aluminum, Copper and Magnesium	Making aircraft, utensils, and windows frames because of its light weight and corrosion resistant.
Steel	Iron, Carbon ,Manganese and other metals	Railway lines, car bodies girders and utensils.
Nichrome	Nichrome and Chromium	Provide resistance in electric heaters and ovens
German silver	Copper, Zinc and Nickel	Making coins

METHODS OF SEPARATING MIXTURES

Mixtures can be separated from applying the following methods:

(a) Decantation

Sediments can be separated from a liquid by pouring out the liquid. This process is called **decantation.**

Experiment

Put some sand in a beaker. Add about 200cm3 of water. Allow sand to settle. Pour off water carefully into another beaker.

Observation

Sand settles at the bottom as sediments.

Less clean water is poured out.

Explanation

Sand does not dissolve in water. Sand is denser than water and thus settles at the bottom as **sediment**. When poured out, the less dense water flows out.

(b)Filtration

Decantation leaves suspended particles in the liquid after separation. Filtration is thus improved decantation.Filtration is the method of separating insoluble mixtures/particles/solids from a liquid.

Experiment: To separate soil and water using filtration

Fold a filter paper to fit well into a filter funnel. Place the funnel in an empty 250 cm3 beaker.

Put one spatula end full of soil into 50cm3 of water. Stir. Put the soil/water mixture into the filter funnel.

Observations

Clean water is collected **below** the filter funnel.

Soil remains **above** the filter paper.

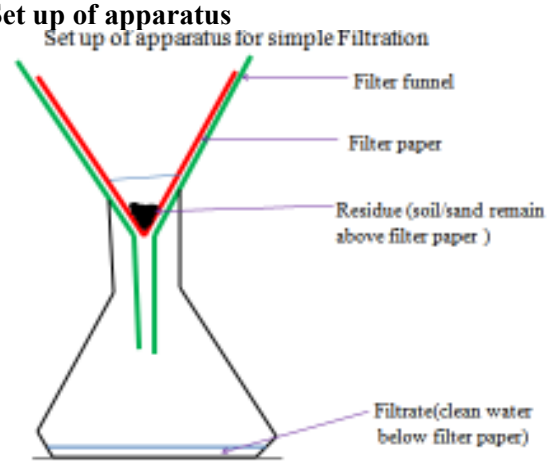
Explanation

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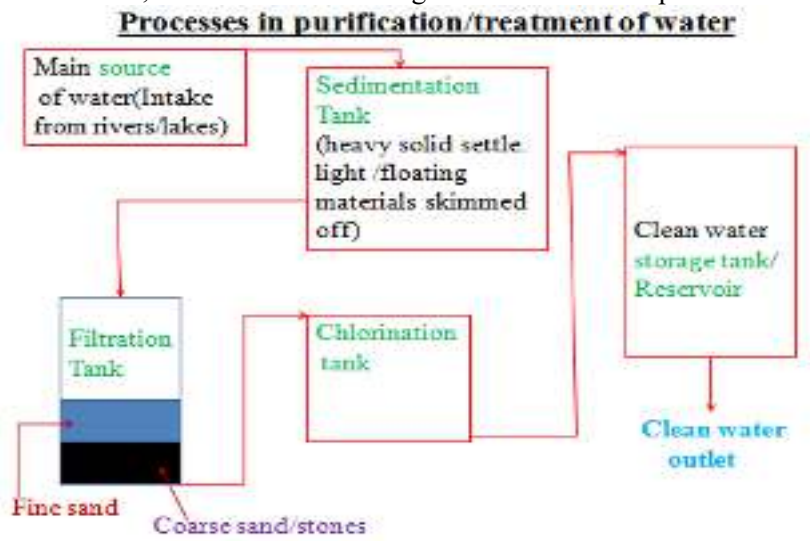
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A filter paper is **porous** which act like a fine sieve with very small **holes**. The holes allow smaller water particles to pass through but do not allow bigger soil particles. The liquid which passes through is called **filtrate**. The solid which do not pass through is called **residue**.

Set up of apparatus



In industries, filtration is used in engine filters to clean up air.



(c)Evaporation

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Evaporation is a method of separating a solute/solid from its solution. This involves heating a solution (solvent and solute)to vapourize the solvent out of the solution mixture leaving pure solute/solid. If a mixture contain insoluble solid, they are filtered out.

Experiment: **To separate a mixture of soil and salt (sodium chloride).**

Procedure:

Put one spatula end full of soil on a filter paper.
Put one spatula full of common salt/sodium chloride into the same filter paper. Mix well using the spatula,.
Place about 200cm3 of water into a beaker.
Put the contents of the filter paper into the water. Stir thoroughly using a glass/stirring rod for about one minute.
Fold a filter paper into a filter funnel.
Pour half portion of the contents in the beaker into the filter funnel.
Put the filtrate into an evaporating dish. Heat on a water bath.

Observation

(i)On mixing
Colourless crystals and brown soil particles appear on the filter paper.
(ii)On adding water
Common soil dissolves in water. Soil particles do not dissolve in water.
(iii)On filtration
Colourless liquid collected as filtrate below the filter funnel/paper.
Brown residue collected above the filter funnel/paper.
(iv)On evaporation
Colourless crystals collected after evaporation

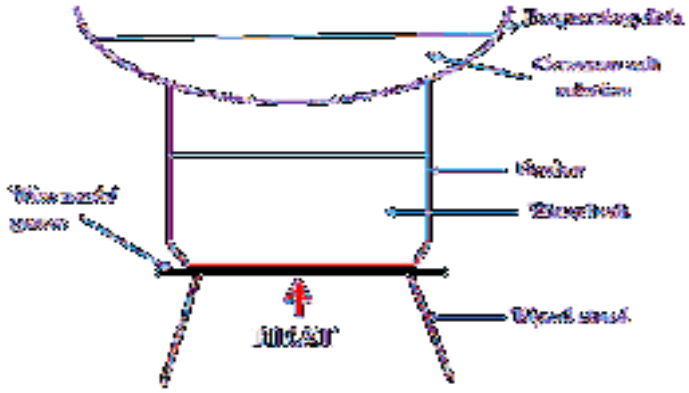
Explanation

Solid mixture of sand and common salt take the colors of the two.
On adding water, common salt dissolves to form a solution.
Soil does not because it is insoluble in water and thus forms a suspension.
On filtration, a residue of insoluble soil does not pass through the filter paper.
It is collected as residue.
Common salt solution is collected as filtrate.
On heating the filtrate, the solvent/water evaporate/vaporize out of the evaporating dish leaving common salt crystals.
Vapourization/evaporation can take place even without heating.
This is the principle/process of drying wet clothes on the hanging line.

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Set up of apparatus



(d) Distillation

Distillation is an improved evaporation where both the solute and the solvent in the solution are separated /collected. Distillation therefore is the process of separating a solution into constituent solid solute and the solvent. It involves heating the solution to evaporate/vaporize the solvent out. The solvent vapour is then condensed back to a liquid.

Experiment: To obtain copper (II) sulphate (VI) crystals and water from copper (II) sulphate (VI) solution.

Procedure:

Put one spatula end full of copper (II) sulphate (VI) crystals into a 250cm3 beaker.
Place about 200cm3 of water into the beaker.
Stir thoroughly using a glass/stirring rod for about one minute.
Pour half portion of the contents in the beaker into a round bottomed/flat/conical flask broken porcelain/sand/glass into the flask.
Put a few pieces of b Stopper the flask.
Connect the flask to a Liebig condenser using delivery tube.
Place a 200cm3 clean empty beaker/conical flask as a receiver at the end of the Liebig condenser.
Circulate water in the Liebig condenser.
Heat the flask strongly on a tripod stand with wire mesh/gauze until there is no more visible **boiling bubbles** in the flask.

Observation

Copper (II) sulphate (VI) crystals dissolve in water to form a blue solution.
On heating, colourless liquid is collected in the receiver.
Blue crystals are left in the flask.
(If gently heated further, the blue crystals turn to white powder)

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Explanation

On heating blue Copper (II) sulphate (VI) solution, the colourless liquid solvents evaporate/vaporize.

The liquid vapour/gas passes through the delivery tube to the Liebig condenser.

The Liebig condenser has a cold water **inlet** near the receiver and cold water **out** let.

This ensures efficient cooling. If the cold water **outlet/inlet** is reversed, the water circulation would be less efficient.

The water in the receiver would be warm. In the Liebig condenser, the cold water condenses the liquid vapour into liquid.

The condensed liquid collects in the receiver as **distillate**.

The solute of blue Copper (II) sulphate (VI) crystals is left in the flask as **residue**.

During simple distillation, therefore, the solution is heated to vaporize /evaporate the solvent/one component which is condensed at a different part of the apparatus.

The purpose of pieces of broken porcelain/porous pot/glass/sand/ is to:

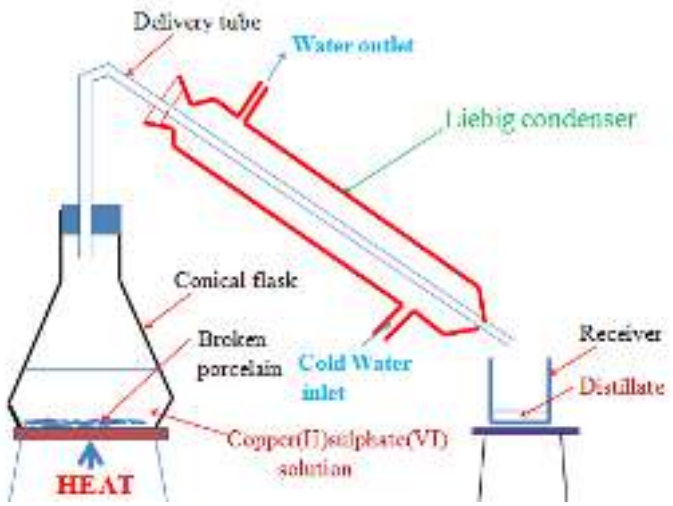
(i) Prevent bumping of the solution during boiling.

(ii) Ensure smooth and even boiling.

Salty sea water can be made pure through simple distillation.

Any mixture with a large difference /40°C in boiling point can be separated using simple distillation.

Set up of apparatus



e)Fractional distillation

Fractional distillation is an improved simple distillation used specifically to separate miscible mixtures with very **close /near** boiling points.

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Fractional distillation involves:

(**i**) Heating the mixture in a conical/round bottomed /flat bottomed flask.

The pure substance with a lower boiling point and thus more volatile evaporates/boils/vaporize first.e.g. Pure ethanol has a boiling point of 78ºC.Pure water has a boiling point of 100 ºC at sea level/one atmosphere pressure.

When a miscible mixture of ethanol and water is heated, ethanol vaporizes /boils/ evaporates first because it is more volatile.

(**ii**)The conical/round bottomed /flat bottomed flask is connected to a long glass tube called **fractionating column**.

The purpose of the fractionating column is to offer areas of condensation for the less volatile pure mixture.

The fractionating column is packed with glass beads/broken glass/ porcelain/ shelves to increase the surface area of condensation of the less volatile pure mixture.

(**iii**)When the vapors rise they condense on the glass beads/broken glass /porcelain / shelves which become hot.

When the temperature of the glass beads/broken glass/porcelain/shelves is beyond the boiling point of the less volatile pure substance, the pure substance rise and condensation take place on the glass beads/broken glass/porcelain/shelves at a higher level on the fractionating column.

The less volatile pure substance trickles/drips back down the fractionating column or back into the conical/round bottomed /flat bottomed flask to be heated again. e.g. If the temperature on glass beads/broken glass/porcelain/shelves is beyond 78ºC, the **more volatile** pure ethanol rise to condense on the glass beads/broken glass /porcelain/shelves **higher** in the fractionating column.

Water condenses and then drip/trickle to the glass beads/broken glass /porcelain /shelves **lower** in the fractionating column because it is **less volatile.**

(**iv**) The fractionating column is connected to a Liebig condenser. The Liebig condenser has a cold water inlet and outlet circulation.

The more volatile mixture that reach the top of the fractionating column is condenses by the Liebig condenser into a receiver. It is collected as the first fraction.

(**v**)At the top of the fractionating column, a thermometer is placed to note/monitor the temperature of the boiling mixtures.

Pure substances have constant/fixed boiling point. When one mixture is completely separated, the thermometer reading rises.

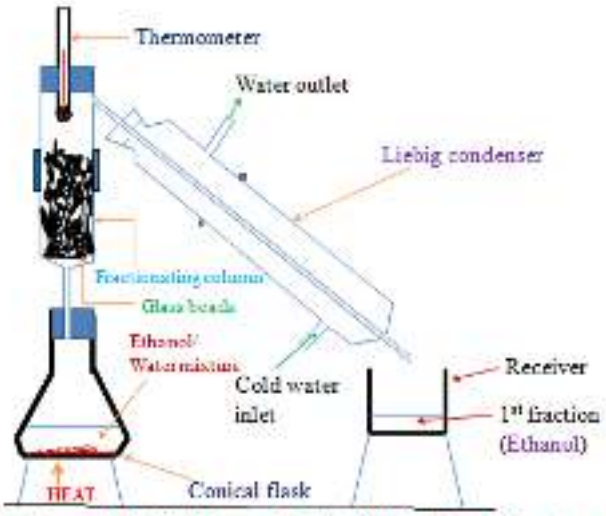
E.g. the thermometer reading remains at78ºC when ethanol is being separated. When no more ethanol is being separated, the mercury/alcohol level in the thermometer rises.

(**vi**)The second /subsequent fractions are collected in the receiver after noting a rise the mercury/alcohol level in the thermometer.

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E.g. the thermometer reading rises to 100ºC when water is being separated. It is passed through the Liebig condenser with the cold water inlet and outlet circulation. It is collected different receiver as the second/subsequent fraction.
(vii)Each fraction collected should be confirmed from known physical/chemical properties/characteristic.
Example
Ethanol
Ethanol is a colourless liquid that has a characteristic smell .When it is put in a watch glass then ignited, it catches fire and burn with a blue flame.
Water
Water is a colourless liquid that has no smell/odour .When it is put in a watch glass then ignited, it does not catch fire.
Set up of apparatus



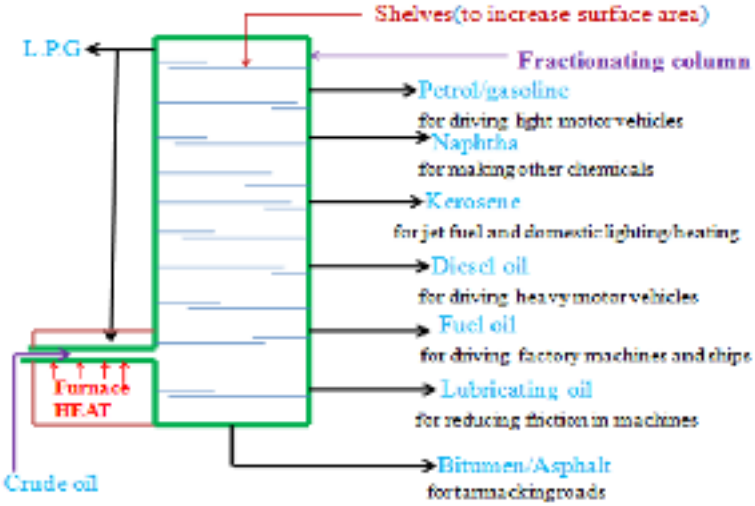
Fractional Distillation of miscible ethanol/water mixture

Industrial application of Fractional distillation
On a large scale,fractional distillation is used:
(i) In fractional distillation of crude oil in an oil refinery.
Crude oil is a mixture of many fractions. When heated in a furnace, the different fractions separate out according to their boiling point. In Kenya,fractional distillation takes place at Changamwe in Mombasa.
(ii) In fractional distillation of air.

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Air contain a mixture of three main useful gases which are condensed by cooling to very low temperature (-200ºC) to form a liquid. The liquid is then heated. Nitrogen is the most volatile (-196 ºC) and thus comes out as the first fraction. Argon (at -186 ºC) is the second fraction. Oxygen (at -183 ºC) is the last fraction. The three gases are very useful industrial gases.



Industrial fractional distillation of crude oil in an oil refinery

(f)Separation of immiscibles (Using a separating funnel)

Two or more liquids that form layers on mixing are immiscible. Immiscible mixture arrange themselves according to their densities

i.e. The denser liquid sink to the bottom. The less dense liquid floats on the denser one. Immicible mixtures can be separated from each other by using a **separating funnel**.

Experiment: To separate an immiscible mixture of paraffin and water.

Procedure

Place about 100cm3 of water into a 250cm3 beaker. Add about 100cm3 of paraffin into the beaker. Stir.

Transfer the mixture into a separating funnel. Allow to settle for about one minute. Open the tap, run out the lower layer out slowly into a clean beaker. Close the tap when the upper layer is very close to the tap.

Run out the intermediate small amount of the mixture near the tap into a beaker. Discard it.

Run out the remaining upper layer into a fresh beaker.

Place a portion of upper and lower layer into a watch glass separately after separating each. Ignite.

Observation

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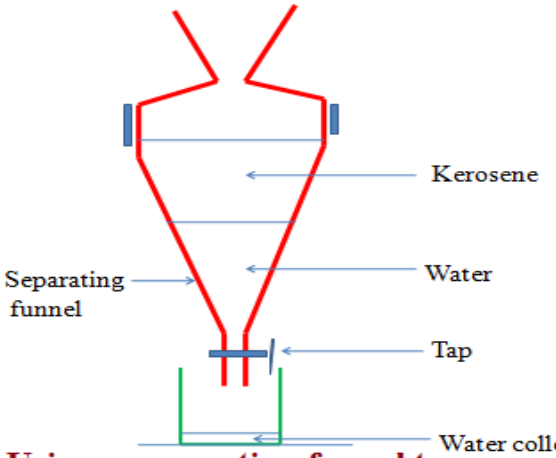
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Water and paraffin are both colourless liquids.
Two layers are formed on mixing.
Colourless odorless liquid collected first. It does not catch fire.
A colourless liquid with characteristic smell collected later/second. It catches fire and burn with a yellow smoky flame.

Explanation

Water and paraffin are immiscible. Water is denser than paraffin. When put in a separating funnel, paraffin float on water. On opening the tap, water runs out. A mixture of water and paraffin at the junction of the two is discarded. It is not pure.

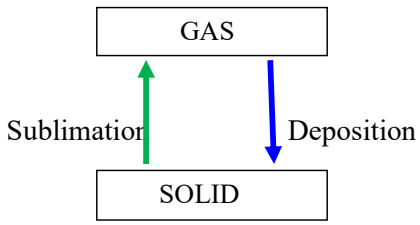
Set up of apparatus



Using a separating funnel to separate a mixture of water/ paraffin

(g)Sublimation/deposition

Some solids on heating do not melt to a liquid but change directly to a gas. The process by which a solid changes to a gas is called **sublimation**. The gas cools back and changes directly to a solid. The process by which a gas changes to a solid is called **deposition.** Sublimation and deposition therefore are the same but opposite processes.



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Some common substances that undergo sublimation/ deposition include:
(i)Iodine (ii)Carbon(IV)oxide (iii)Camphor (iv) ammonium chloride (v)Iron(III)chloride (vi)Aluminum(III)chloride (vii) benzoic acid

If a mixture has any of the above as a component, then on heating it will change to a gas and be deposited away from the source of heating.

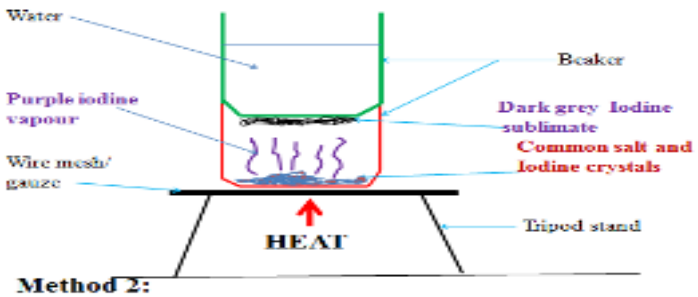
Procedure

Place about one spatula full of ammonium chloride crystals into a clean dry 100cm3 beaker. Add equal amount of sodium chloride crystals into the beaker. Swirl to mix.

Place the beaker on a tripod stand.

Put about 100cm3 of water into another beaker. Place carefully the beaker containing water on top of the beaker containing the solid mixture. Light/ignite a burner and heat the solid.

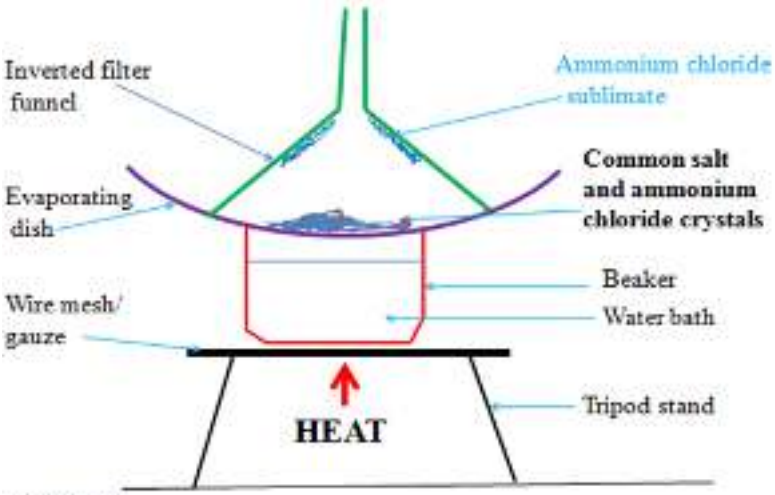
Set up of apparatus:



Method 2:
Using sublimation to separate common salt and Iodine crystals

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Method 1

Using sublimation to separate common salt and ammonium chloride

Observation

(i)With ammonium chloride/common salt mixture

White fumes produced.
White sublimate deposited
Colourless residue left

(ii)With Iodine/common salt mixture

Purple fumes produced.
Dark grey sublimate deposited
Colourless residue left

Explanation

(i)On heating a mixture of ammonium chloride and common salt, a white fume of ammonium chloride is produced. The white fumes solidify as white sublimate on the cooler parts. Common salt remains as residue.

Chemical equation:

Ammonium chloride **solid** $\rightleftharpoons$ Ammonium chloride **gas**

$NH_4Cl(s) \rightleftharpoons NH_4Cl(g)$

(ii)On heating a mixture of Iodine and common salt, a purple fume of Iodine vapour is produced. The purple fumes solidify as dark grey sublimate on the cooler parts. Common salt remains as residue.

Chemical equation:

Iodine **solid** $\rightleftharpoons$ Iodine **gas**

$I_2(s) \rightleftharpoons I_2(g)$

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(h)Chromatography

Chromatography is a method of separating components of a solution mixture by passing it through a medium where the different components move at different rates. The medium through which the solution mixture is passed is called **absorbent material**.

Paper chromatography is a method of separating colored dyes by using paper as the absorbent material.

Since dyes are insoluble/do not dissolve in water, ethanol and propanone are used as suitable solvents for dissolving the dye.

Practically, a simple paper chromatography involve placing a dye/material on the absorbent material, adding slowly a suitable soluble solvent on the dye/material using a dropper, the solvent spread out on the absorbent material carrying the soluble dye away from the origin.

The spot on which the dye is initially/originally placed is called **baseline**. The farthest point the solvent spread is called **solvent front**.

The farthest a dye can be spread by the solvent depend on:

(i) Density of the dye-the denser the dye, the less it spread from the basely ne by the solvent.

(ii) Stickiness of the dye-some dyes sticks on the absorbent material more than other thus do not spread far from baseline.

Experiment: To investigate the colors in ink

Procedure

Method 1

Place a filter paper on an empty beaker. Put a drop of black/blue ink in the centre of the filter paper. Wait for about one minute for the ink drop to spread. Using a clean teat pipette/dropper add one drop of ethanol/propanone. Wait for about one minute for the ink drop to spread further. Add about twenty other drops of ethanol waiting for about one minute before each addition. Allow the filter paper to dry.

Experiment: To investigate the colors in ink

Procedure

Method 2

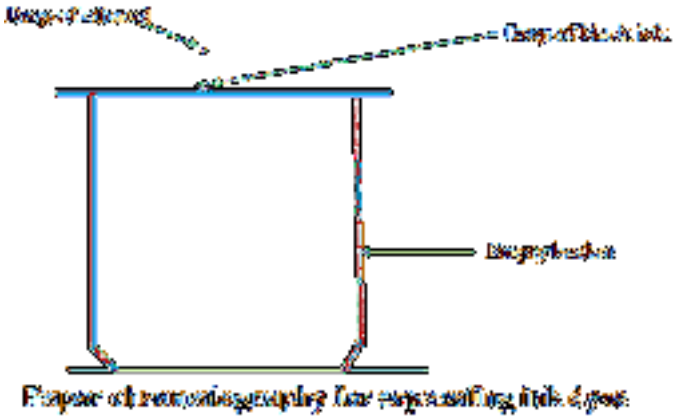
Cut an 8 centimeter thin strip of a filter paper. At about 3cm on the strip, place a drop of ink. Place the filter paper in a 10cm length boiling tube containing 5cm3 of ethanol. Ensure the cut strip of the filter paper just dips into the ethanol towards the ink mark. Cover the boiling tube. Wait for about twenty minutes. Remove the boiling tube and allow the filter paper to dry.

Set up of apparatus

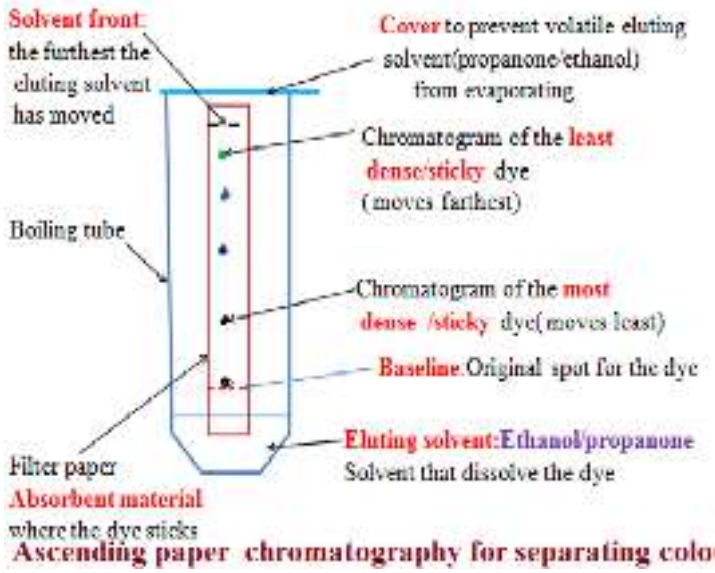
Method 1

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Set up of apparatus
Method 2



Ascending paper chromatography for separating coloured dyes

Explanation

When a drop of ink is placed on an absorbent material it sticks. On adding an eluting solvent, it dissolves the dye spread out with it. The denser and sticky pure dye move least. The least dense/sticky pure dye move farthest. A pure dye will produce the same chromatogram/spot if the same eluting solvent is used on the same absorbent material. Comparing the distance moved by a pure dye with a mixture, the coloured dyes in a mixture can be deduced as below:

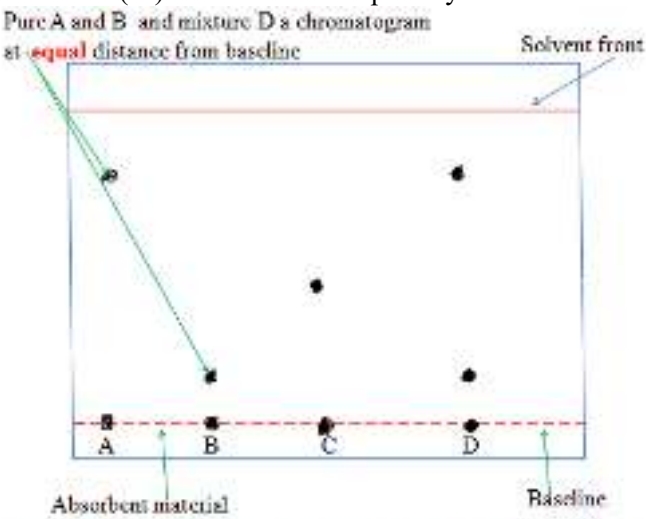
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Example 1

The chromatogram of pure dyes A, B ,C and a dye mixture D is shown below
Determine the pure dyes present in D. On the diagram show:

(i)the solvent front
(ii) Baseline
(Iii) the most soluble pure dye



Chromatogram showing pure A,B,C and mixture D

(i) Solvent extraction

Solvent extraction is a method of separating oil from nuts/seeds. Most nuts contain oil. First the nuts are crushed to reduce their size and increase the surface area. A suitable volatile solvent is added. The mixture is filtered. The filtrate solvent is then allowed to crystallize leaving the oil/fat. If a filter paper is rubbed/smeared with the oil/fat, it becomes translucent. This is the test for the presence of oil/fat.

Experiment: To extract oil from Macadamia nut seeds

Procedure

Crush Macadamia nut seeds form the hard outer cover .Place the inner soft seed into a mortar. Crush (add a little sand to assist in crushing).

Add a little propanone and continue crushing. Continue crushing and adding a little propanone until there is more liquid mixture than the solid. Decant/filter. Put the filtrate into an evaporating dish. Vapourize the solvent using solar energy /sunlight. Smear/rub a portion of the residue left after evaporation on a clean dry filter paper.

Observation /Explanation

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Propanone dissolve fat/oil in the macadamia nuts. Propanone is more volatile (lower boiling point) than oil/fat. In sunlight/solar energy, propanone evaporate/vaporize leaving oil/fat(has a higher boiling point).Any seed like corn, wheat , rice, soya bean may be used instead of macadamia seed. When oil/fat is rubbed/ smeared on an opaque paper, it becomes translucent.

(j) Crystallization

Crystallization is the process of using solubility of a solute/solid to obtain the solute/solid crystals from a saturated solution by cooling or heating the solution.
A crystal is the smallest regular shaped particle of a solute. Every solute has unique shape of its crystals.
Some solutions form crystals when heated. This is because less solute dissolves at higher temperature. Some other solutions form crystals when cooled. This is because less solute dissolves at lower temperature.

Experiment; To crystallize copper (II) sulphate (VI) solution

Procedure:

Place about one spatula full of hydrated copper sulphate (VI) crystals into 200cm3 of distilled water in a beaker. Stir. Continue adding a little more of the hydrated copper sulphate (VI) crystals and stirring until no more dissolve. Decant/filter. Cover the filtrate with a filter paper. Pierce and make small holes on the filter paper cover. Preserve the experiment for about seven days.

Observation/Explanation

Large blue crystals formed
When hydrated copper (II) sulphate crystals are placed in water, they dissolve to form copper (II) sulphate solution. After some days water slowly evaporate leaving large crystals of copper (II) sulphate. If the mixture is heated to dryness, small crystals are formed.

Physical/Temporary and Chemical changes

A physical/temporary change is one which **no new** substance is formed and is **reversible** back to original.
A chemical/permanent change is one which **a new** substance is formed and is **irreversible** back to original.

The following experiments illustrates physical and chemical changes

(a)Heating ice

Place about 10g of pure ice in a beaker. Determine its temperature. Record it at time "0.0" in the table below. Heat the ice on a strong Bunsen flame and determine its temperature after every 60seconds/1minute to complete the table below:

Time/minutes	0	1	2	3	4	5	6	7	8
Temperature (°C)	-2	0	0	40	80	90	95	95	96

Plot a graph of time against Temperature (y-axes)
Explain the shape of your graph

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Melting/freezing/fusion/solidification and **boiling /vaporization /evaporation** are the two physical processes.
Melting /freezing point of pure substances is fixed /constant.
The boiling point of pure substance depends on **external** atmospheric **pressure**.
Melting/fusion is the physical change of a **solid** to **liquid**.
Freezing is the physical change of a **liquid** to **solid**.
Melting/freezing/fusion/solidification is therefore two **opposite** but **same** reversible physical processes i.e.

A (**s**) A (**l**)

Boiling/vaporization/evaporation is the physical change of a **liquid** to **gas**.
Condensation/ liquidification is the physical change of **gas** to **liquid**.
Boiling/vaporization/evaporation and condensation/ liquidification are therefore two **opposite** but **same** reversible physical processes i.e.

B (**l**) B(**g**)

Practically
(i) Melting/liquidification/fusion involves **heating** a solid to **weaken** the strong bonds holding the solid particles together.
Solids are made up of very strong bonds holding the particles **very close** to each other (**Kinetic Theory of matter**).
On heating these particles gain energy/heat from the surrounding heat source to form a liquid with **weaker** bonds holding the particles close together but with some degree of **freedom**.
(ii)Freezing/fusion/solidification involves cooling a liquid to reform /rejoin the very strong bonds to hold the particles **very close** to each other as solid and thus lose their degree of **freedom** (**Kinetic Theory of matter**).
Freezing /fusion / solidification is an **exothermic** (**-**ΔH) process that require particles holding the liquid together to lose energy to the surrounding.
(iii)Boiling/vaporization/evaporation involves **heating** a liquid to completely **break/free** the bonds holding the liquid particles together.
Gaseous particles have high degree of **freedom** (**Kinetic Theory of matter**).
Boiling /vaporization / evaporation is an **endothermic** (+ΔH) process that require/absorb energy from the surrounding.
(iv)Condensation/liquidification is **reverse** process of boiling /vaporization / evaporation.
It involves gaseous particles losing energy to the surrounding to form a liquid.

AIR OXYGEN AND COMBUSTION
A.THE ATMOSPHERE
1. The atmosphere is made up of air. Air is a mixture of colourless, odorless gases which is felt as wind (air in motion).All living things breath in air for respiration. Plants use air for respiration and photosynthesis.

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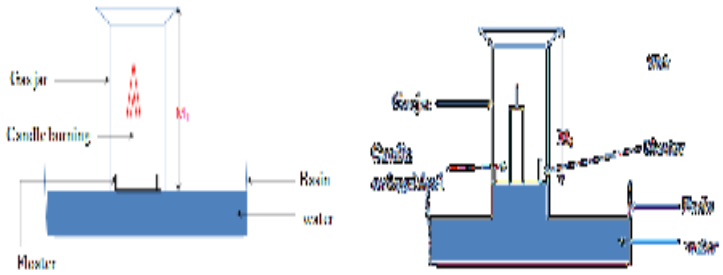
2. The main gases present in the atmosphere/air:

Gas	Approximate % composition by volume
Nitrogen	78.0
Oxygen	21.0
Carbon(IV)oxide	0.03
Noble gases	1.0
Water vapour	Vary from region

3. The following experiments below shows the presence and composition of the gases in air/atmosphere

(a)To find the composition of air supporting combustion using a candle stick

Procedure
Measure the length of an empty gas jar M_1. Place a candle stick on a Petri dish. Float it on water in basin/trough. Cover it with the gas jar. Mark the level of the water in the gas jar M_2. Remove the gas jar. Light the candle sick. Carefully cover it with the gas jar. Observe for two minutes. Mark the new level of the water M_3.
Set up of apparatus



Sample observations
Candle continues to burn then extinguished/goes off
Level of water in the gas jar rises after igniting the candle
Length of empty gas jar = M_1= 14cm
Length of gas jar **without** water before igniting candle = M_2= 10 cm
Length of gas jar **with** water before igniting candle = M_1 - M_2= 14- 10 = 4 cm
Length of gas jar **with** water after igniting candle = M_3 = 8 cm
Length of gas jar **without** water after igniting candle = M_1 - M_3 = 10 -8 = 2 cm

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Explanation

Candle burns in air. In a closed system (vessel), the candle continues to burn using the part of air that support burning/combustion. This is called the **active part of air**. The candle goes off/extinguished when all the active part of air is used up. The level of the water rises to occupy the space /volume occupied by the used active part of air.

The experiment is better when very dilute **sodium/potassium hydroxide** is used instead of water. Dilute Potassium/ sodium hydroxide absorb **Carbon (IV) oxide** gas that comes out from burning/combustion of candle stick.

From the experiment above the % composition of the:

(i) Active part of air can be calculated:

$\frac{M_2 - M_3}{M_2} \times 100\%$ => $\frac{10 - 8}{10cm} \times 100\% = \mathbf{20\%}$

(ii) Inactive part of air can be calculated:

100% **-20% = 80%** // $\frac{M_3}{M_2}$ => $\frac{8}{10cm} \times 100\% = \mathbf{80\%}$

(b)To find the composition of active part of air using heated copper turnings.

Procedure

Clamp a completely packed/filled open ended glass tube with copper turnings. Seal the ends with glass/cotton wool.

Label two graduated syringes as "A" and "B" Push out air from syringe "A". Pull in air into syringe "B".

Attach both syringe "A" and "B" on opposite ends of the glass tube.

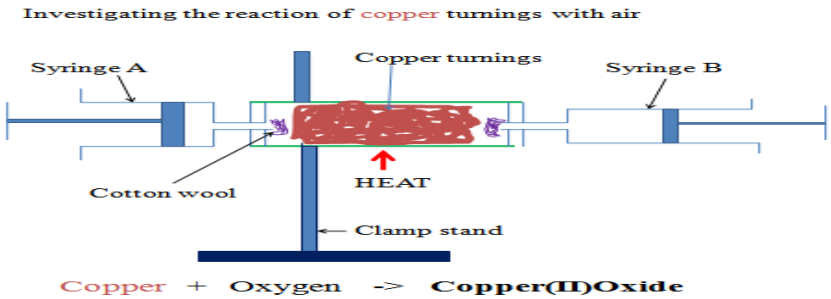
Determine and record the volume of air in syringe "B" V_1.

Heat the glass tube strongly for about three minutes.

Push all the air slowly from syringe "B" to syringe "A" as heating continues. Push all the air slowly from syringe "A" back to syringe "B" and repeatedly back and forth.

After about ten minutes, determine the new volume of air in syringe "B" V_2

Set up of apparatus



Sample observations

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Colour change from brown to black
Volume of air in syringe "B" before heating V_1 = 158.0cm3
Volume of air in syringe "B" after heating V_2 = 127.2cm3
Volume of air in syringe "B" used by copper $V_1 - V_2$ = 30.8cm3

Sample questions

1. What is the purpose of:

(i) glass/cotton wool

To prevent/stop copper turnings from being blown into the syringe/out of the glass tube

(ii) Passing air through the glass tube repeatedly

To ensure all the active part of air is used up

(iii) Passing air through the glass tube slowly

To allow enough time of contact between the active part of and the heated copper turnings

2. State and explain the observations made in the glass tube.

Colour change from brown to black

Brown copper metal reacts with the active part of air/oxygen to form black copper (II) oxide.

Chemical equation

Copper + Oxygen -> Copper (II) oxide
2Cu(s) + O_2 (g) -> 2CuO(s)

The reaction reduces the amount/volume of oxygen in syringe "B" leaving the inactive part of air. Copper only react with oxygen when heated.

3. Calculate the % of

(i) Active part of air

$$\% \text{ active part of air} = \frac{V_1 - V_2}{V_1} \times 100\% \Rightarrow \frac{30.8\text{cm3}}{158.0\text{cm3}} \times 100\% = \mathbf{19.493}\%$$

(ii) Inactive part of air

Method 1

$$\% \text{ inactive part of air} = \frac{V_2}{V_1} \times 100\% \Rightarrow \frac{127.2\text{cm3}}{158.0\text{cm3}} \times 100\% = \mathbf{80.506}\%$$

Method 2

% inactive part of air = 100% -% active part of air
=> 100 % - 19.493 % = **80.507**%

4. The % of active part of air is theoretically higher than the above while % of inactive part of air is theoretically lower than the above. Explain.

Not all the active part of air reacted with copper

5. State the main gases that constitute:

(a)active part of air.

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Oxygen

(b) **Inactive part of air**

Nitrogen, carbon (IV) oxide and noble gases

6. If the copper turnings are replaced with magnesium shavings the % of active part of air obtained is extraordinary very high. Explain.

Magnesium is more reactive than copper. The reaction is highly exothermic. It generates enough heat for magnesium to react with both oxygen and nitrogen in the air.

A white solid/ash mixture of Magnesium oxide and Magnesium nitride is formed. This considerably reduces the volume of air left after the experiment.

Chemical equation

Magnesium + Oxygen -> magnesium (II) oxide

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

Magnesium + Nitrogen -> magnesium (II) nitride

$3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$

(c)To find the composition of active part of air using alkaline pyrogallol

Procedure

Measure about 2cm3 of dilute sodium hydroxide into a graduated gas jar. Record the volume of the graduated cylinder V_1.

Place about two spatula end full of pyrogallol/1, 2, 3-trihydroxobenzene into the gas jar. Immediately place a cover slip firmly on the mouth of the gas jar. Swirl thoroughly for about two minutes.

Invert the gas jar in a trough/basin containing water. Measure the volume of air in the gas jar V_2

Sample observations

Colour of pyrogallol/1, 2, 3-trihydroxobenzene change to **brown**.

Level of water in gas jar rises when inverted in basin/trough.

Volume of gas jar /air in gas jar V1= **800cm3**

Volume of gas jar /air in gas jar after shaking with alkaline pyrogallol/1, 2, 3-trihydroxobenzene V_2= **640 cm3**

Sample questions

1. Which gas is absorbed by alkaline pyrogallol/1,2,3-trihydroxobenzene

Oxygen

2. Calculate the

(i) % of active part of air

$\frac{V_1 - V_2}{V_1} \times 100\% \Rightarrow \frac{800cm3 - 640\ cm3}{800cm3} \times 100\% = \mathbf{20\%}$

(ii) % of inactive part of air

$\frac{V_2}{} \times 100\% \Rightarrow \frac{640\ cm3}{} \times 100\% = \mathbf{80\%}$

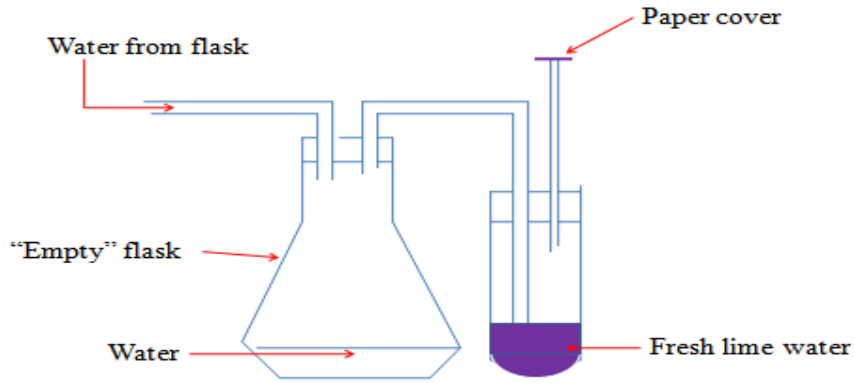
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V_1 800cm3

(d)To establish the presence of carbon (IV) oxide in air using lime water

Pass tap water slowly into an empty flask as in the set up below



Sample observation questions

1. **What is the purpose of paper cover?**

To ensure no air enters into the lime water.

2. **What happens when water enters the flask?**

It forces the air from the flask into the lime water.

3. What is observed when the air is bubbled in the lime water?

A white precipitate is formed. The white precipitate dissolves on prolonged bubbling of air.

4. (a) Identify the compound that form:

(i)lime water

Calcium hydroxide / $Ca(OH)_2$

(ii) White precipitate

Calcium carbonate/ $CaCO_3$

(iii) When the white precipitate dissolves

Calcium hydrogen carbonate/ $CaHCO_3$

(b)Write the chemical equation for the reaction that tale place when:

(i) White precipitate is formed

Calcium hydroxide + carbon (IV) oxide -> **Calcium** carbonate + water

$Ca(OH)_{2(aq)} + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

(ii) White precipitate dissolves

Calcium carbonate + water+ carbon (IV) oxide -> Calcium hydrogen carbonate

$CaCO_3(s) + H_2O(l) + CO_2(g) \rightarrow CaHCO_3(aq)$

5. State the chemical test for the presence of carbon (IV) oxide gas based on 4(a) and (b) above:

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Carbon (IV) oxide forms a white precipitate with lime water that dissolves in excess of the gas.

6. State the composition of carbon (IV) oxide gas by volume in the air.

About 0.03% by volume

B.OXYGEN

a) Occurrence.

1. Fifty 50% of the earth's crust consist of Oxygen combined with other elements e.g. oxides of metals
2. About 70% of the earth is water made up of Hydrogen and Oxygen.
3. About 20% by volume of the atmospheric gases is Oxygen that form the active part of air.

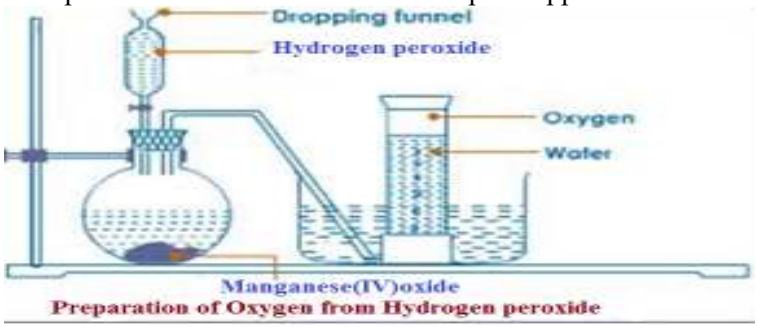
b) School laboratory preparation.

Oxygen was first prepared in 1772 by Karl Scheele and later in 1774 by Joseph Priestly. It was Antony Lavoisier who gave it the name "Oxygen"

Procedure

Method 1: Using Hydrogen peroxide

Half fill a trough/basin with tap water. Place a bee hive shelf/stand into the water. Completely fill the gas jar with water and invert in onto the bee hive shelf/stand. Clamp a round bottomed flask and set up the apparatus as below.



Preparation of Oxygen from Hydrogen peroxide

Collect several gas jars of Oxygen covering each sample.

Sample observation questions

1. **What is observed when the hydrogen peroxide is added into the flask?**

Rapid effervescence/bubbling/fizzing

2. Describe the colour and smell of the gas

Colourless and odorless

3. (a)Name the method of gas collection used.

-Over water
-Upward delivery
-Down ward displacement of water

(b)What property of Oxygen makes it to be collected using the method above?

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-Slightly soluble in water

4. What is the purpose of manganese (IV) oxide?

Manganese (IV) oxide is **catalyst**.

A catalyst is a substance that speeds up the rate of a chemical reaction but remain chemically unchanged at the end of the reaction.

Hydrogen peroxide decomposes slowly to form water and Oxygen gas.

A little Manganese (IV) oxide speeds up the rate of decomposition by **reducing** the time taken for a given volume of Oxygen to be produced.

5. Write the equation for the reaction.

Hydrogen peroxide -> Water + Oxygen

$2H_2O_2$ (aq) -> $2H_2O$ (l) + O_2 (g)

6. Lower a glowing splint slowly into a gas jar containing Oxygen gas. State what is observed.

The glowing splint relights/rekindles

Oxygen relights/rekindles a glowing splint. This is the confirmatory test for the presence of Oxygen gas

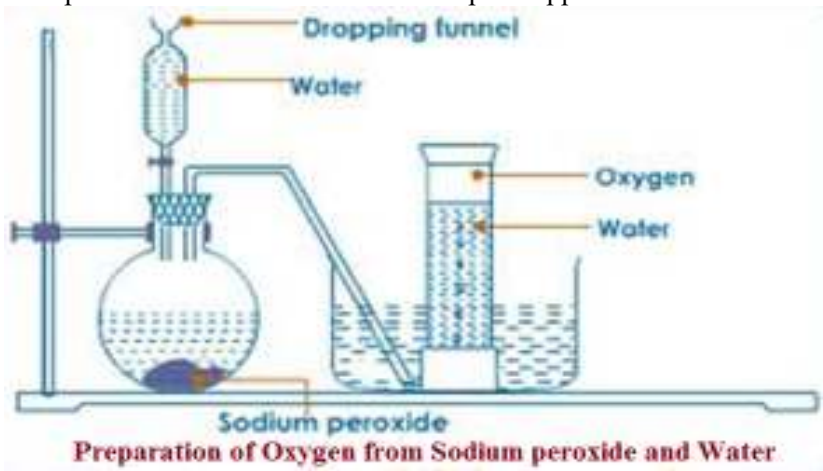
Method 1: **Using Sodium peroxide**

Half fill a trough/basin with tap water. Add four drops of phenolphthalein indicator.

Place a bee hive shelf/stand into the water.

Completely fill a gas jar with water and invert in onto the bee hive shelf/stand.

Clamp a round bottomed flask and set up the apparatus as below.



Preparation of Oxygen from Sodium peroxide and Water

Collect several gas jars of Oxygen covering each sample.

Sample observation questions

1. **What is observed when water is added?**

(i) Into the flask containing sodium peroxide

Rapid effervescence/bubbling/fizzing

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(ii) Phenolphththalein

Remains colourless /Phenolphthalein indicator is colourless in neutral solution

2. Describe the colour and smell of the gas

Colourless and odorless

3.(a)Name the method of gas collection used.

-Over water. Oxygen is slightly soluble in water.

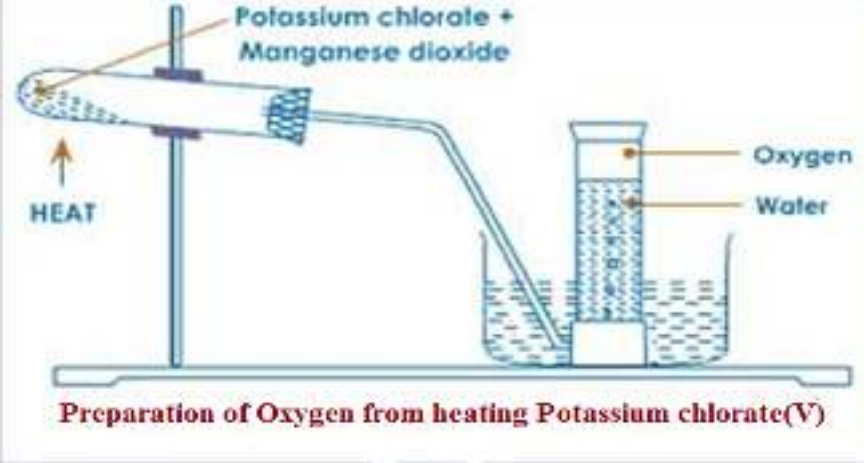
4. Test the gas by lowering a glowing splint slowly into a gas jar containing the prepared sample.

The glowing splint relights/rekindles. This confirms the presence of Oxygen gas

5. Write the equation for the reaction.

Sodium peroxide + Water -> Sodium hydroxide + Oxygen

$2Na_2O_2$ (aq) + 2H2O (l) -> 4NaOH (aq) + O_2 (g)



Preparation of Oxygen from heating Potassium chlorate(V)

1. Test the gas by lowering a glowing splint slowly into a gas jar containing the prepared sample.

The glowing splint relights/rekindles.

This confirms the presence of Oxygen gas

2. Write the equation for the reaction.

Potassium Chlorate (V) -> Potassium Chloride + Oxygen

$2KClO_3$ (aq) -> 2KCl (aq) + $3O_2$ (g)

3. What is the purpose of manganese (IV) oxide?

Manganese (IV) oxide is **catalyst**.

A catalyst is a substance that speeds up the rate of a chemical reaction but remain chemically unchanged at the end of the reaction.

Potassium Chlorate (V) decomposes slowly to form potassium chloride and Oxygen gas.

A little Manganese (IV) oxide speeds up the rate of decomposition by **reducing** the time taken for a given volume of Oxygen to be produced.

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(c)Uses of Oxygen

1. Oxygen is put in cylinders for use where natural supply is not sufficiently enough. This is mainly in:

(i)Mountain climbing/Mountaineering-at high altitudes, the concentration of air/oxygen is low. Mountain climbers must therefore carry their own supply of oxygen for breathing.

(ii) Deep sea diving-Deep sea divers carry their own supply of Oxygen.

(iii) Saving life in hospitals for patients with breathing problems and during anesthesia.

2. A mixture of oxygen and some other gases produces a flame that is very hot.

(i) **Oxy-acetylene/ethyne** flame is produced when Ethyne/acetylene gas is burnt in pure oxygen. The flame has a temperature of about 3000ºC.It is used for **welding /cutting metals**.

(ii)**Oxy-hydrogen** flame is produced when Hydrogen is burn in pure oxygen. The flame has a temperature of about 2000ºC.It is used also for **welding /cutting metals**.

3. **Oxy-hydrogen** mixture is used as rocket fuel

4. A mixture of charcoal, petrol and liquid Oxygen is an explosive.

(d) Chemical properties of Oxygen /combustion.

Oxygen is a very reactive non metal. Many elements react with oxygen through burning to form a group of compounds called **Oxides**.

Burning/combustion is the reaction of Oxygen with an element/substances.

Reaction in which a substance is added oxygen is called **Oxidation reaction**.

Burning/combustion are an example of an oxidation reaction.

Most **non metals** burn in Oxygen/air to form an Oxide which in solution / dissolved in water is **acidic** in nature. They turn blue litmus red.e.g. Carbon (IV) oxide/CO_2, Nitrogen (IV) oxide/ NO_2, Sulphur (IV) oxide/ SO_2

Some non metals burn in Oxygen/air to form an Oxide which in solution / dissolved in water is **neutral** in nature. They **don't** turn blue or red litmus. E.g. Carbon (II) oxide/CO, Water/ H_2O

All **metals** burns in Oxygen/air to form an Oxide which in solution/dissolved in water is **basic/alkaline** in nature. They turn red litmus blue.e.g.

Magnesium oxide/MgO, Sodium Oxide/ Na_2O, Copper (II) oxide/CuO

Elements/substances burn **faster** in pure Oxygen than in air

Air contains the inactive part of air that **slows** the rate of burning of substances/elements.

(i)Reaction of metals with Oxygen/air

The following experiments show the reaction of metals with Oxygen and air.

I. Burning Magnesium

Procedure

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(a)Cut a 2cm length piece of magnesium ribbon. Using a pair of tongs introduce it to a Bunsen flame. Remove it when it catches fire. Observe.
Place the products in a beaker containing about 5cm3 of water. Test the solution/mixture using litmus papers

(b)Cut another 2cm length piece of magnesium ribbon. Using a pair of tongs introduce it to a Bunsen flame. When it catches fire, lower it slowly into a gas jar containing Oxygen.

Place about 5cm3 of water into the gas jar. Test the solution/mixture using litmus papers. Test the solution/mixture using litmus papers

Observations

(a)In air

Magnesium burns with a bright blindening flame in air forming white solid/ash /powder. Effervescence/bubbles/ fizzing Pungent smell of urine. Blue litmus paper remains blue. Red litmus paper turns blue

(b) In pure Oxygen

Magnesium burns **faster** with a very bright blindening flame pure oxygen forming white solid/ash /powder. No effervescence/bubbles/ fizzing. No pungent smell of urine. Blue litmus paper remains blue. Red litmus paper turns blue

Explanation

Magnesium burns in air producing enough heat energy to react with both Oxygen and Nitrogen to form **Magnesium Oxide** and **Magnesium nitride.** Both Magnesium Oxide and Magnesium nitride are white solid/ash /powder.

Chemical equations

Magnesium + Oxygen -> Magnesium Oxide
$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

Magnesium + Nitrogen -> Magnesium **Nitr**ide
$3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$

Magnesium Oxide dissolves in water to form a basic/alkaline solution of Magnesium hydroxide

Chemical equations

Magnesium Oxide + Water -> Magnesium hydroxide
$2Mg(s) + O_2(l) \rightarrow 2MgO(s)$

Magnesium Nitride dissolves in water to form a basic/alkaline solution of Magnesium hydroxide and producing **Ammonia gas**. Ammonia is also an alkaline/basic gas that has a pungent smell of urine.

Chemical equations

Magnesium Nitride + Water -> Magnesium hydroxide + Ammonia gas
$Mg_3N_2(s) + 6H_2O(l) \rightarrow 3Mg(OH)_2(aq) + 2NH_3(g)$

II. Burning Sodium

Procedure

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(a)**Carefully** cut a very small piece of sodium. Using a deflagrating spoon introduce it to a Bunsen flame. Remove it when it catches fire. Observe.
Place the products in a beaker containing about 20cm3 of water. Test the solution/mixture using litmus papers

(b) **Carefully** cut another very small piece of sodium. Using a deflagrating spoon introduce it to a Bunsen flame. When it catches fire, lower it slowly into a gas jar containing Oxygen.

Place about 20 cm3 of water into the gas jar. Test the solution/mixture using litmus papers. Test the solution/mixture using litmus papers

Observations

(a)In air

Sodium burns with a **yellow** flame in air forming a **black** solid. Blue litmus paper remains blue. Red litmus paper turns blue

(b) In pure Oxygen

Sodium burns **faster** with a golden **yellow** flame in pure oxygen forming a **yellow** solid. Effervescence/bubbles/ fizzing. Gas produced relights glowing splint. Blue litmus paper remains blue. Red litmus paper turns blue.

Explanation

(a)Sodium burns in air forming black **Sodium Oxide**

Chemical equations

Sodium + Oxygen/air -> Sodium Oxide

$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$

Sodium Oxide dissolves in water to form a basic/alkaline solution of Sodium hydroxide

Chemical equations

Sodium Oxide + Water -> Sodium hydroxide

$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$

(b)Sodium burns in pure oxygen forming yellow **Sodium peroxide**

Chemical equations

Sodium + Oxygen -> Sodium peroxide

$2Na(s) + O_2(g) \rightarrow Na_2O_2(s)$

Sodium peroxide dissolves in water to form a basic/alkaline solution of Sodium hydroxide. Oxygen is produced.

Chemical equations

Sodium Oxide + Water -> Sodium hydroxide + Oxygen

$2Na_2O_2(s) + 2H_2O(l) \rightarrow 4NaOH(\mathbf{aq}) + O_2(l)$

III. Burning Calcium

Procedure

(a)Using a pair of tongs hold the piece of calcium on a bunsen flame. Observe.

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Place the products in a beaker containing about 2cm3 of water. Test the solution/mixture using litmus papers

(b)Using a pair of tongs hold another piece of calcium on a Bunsen flame. Quickly lower it into a gas jar containing Oxygen gas .Observe.

Place about 2cm3 of water. Swirl.

Test the solution/mixture using litmus papers

Observations

(a)In air

Calcium burns with difficulty producing a faint **red** flame in air forming a **white** solid. Blue litmus paper remains blue. Red litmus paper turns blue

(b) In pure Oxygen

Calcium burns with difficulty producing a less faint **red** flame Oxygen forming a **white** solid. Blue litmus paper remains blue. Red litmus paper turns blue

Explanation

(a)Calcium burns in air forming white **calcium Oxide.** Calcium Oxide coat/cover the calcium preventing further burning.

Chemical equations

Calcium + Oxygen/air -> calcium Oxide

$2Ca(s) + O_2(g) \rightarrow 2CaO(s)$

Small amount of Calcium Oxide dissolves in water to form a basic/alkaline solution of Calcium hydroxide. The common name of Calcium hydroxide is **lime water.**

Chemical equations

Calcium Oxide + Water -> Calcium hydroxide

$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$

IV. Burning Iron

Procedure

(a)Using a pair of tongs hold the piece of Iron wool/steel wire on a Bunsen flame. Observe.

Place the products in a beaker containing about 2cm3 of water. Test the solution/mixture using litmus papers

(b)Using a pair of tongs hold another piece of Iron wool/steel wire on a Bunsen flame.

Quickly lower it into a gas jar containing Oxygen gas .Observe.

Place about 2cm3 of water. Swirl. Test the solution/mixture using litmus papers

Observations

(a)In air

Iron wool/steel wire burns producing an Orange flame in air forming a **brown** solid. Blue litmus paper remains blue. Red litmus paper turns faint blue

(b) In pure Oxygen

Iron wool/steel wire burns producing a golden **Orange** flame in Oxygen forming a **Brown** solid. Blue litmus paper remains blue. Red litmus paper turns faint blue

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Explanation

(a)Iron burns in air forming brown **Iron (III) Oxide**

Chemical equations

Iron + Oxygen/air -> Iron (III) Oxide

$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

Very small amount of Iron (III) Oxide dissolves in water to form a weakly basic/alkaline **brown** solution of Iron (III) hydroxide.

Chemical equations

Calcium Oxide + Water -> Iron (III) hydroxide

$Fe_2O_3(s) + 3H_2O(l) \rightarrow 2Fe(OH)_3$ **(s)**

V. Burning Copper

Procedure

(a)Using a pair of tongs hold the piece of copper turnings/shavings on a Bunsen flame.

Observe.

Place the products in a beaker containing about 2cm3 of water. Test the solution/mixture using litmus papers

(b)Using a pair of tongs hold another piece of Copper turnings/shavings on a Bunsen flame. Quickly lower it into a gas jar containing Oxygen gas .Observe. Place about 2cm3 of water. Swirl. Test the solution/mixture using litmus papers

Observations

(a)In air

Copper turnings/shavings burns with difficulty producing a **green** flame in air forming a **black** solid. Blue litmus paper remains blue. Red litmus paper turns faint blue

(b) In pure Oxygen

Copper turnings/shavings burns less difficulty producing a **green** flame in Oxygen forming a **Brown** solid. Blue litmus paper remains blue. Red litmus paper turns faint blue

Explanation

(a)Copper burns in air forming black **Copper (II) Oxide**

Chemical equations

Copper + Oxygen/air -> Copper (II) Oxide

$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$

Very small amount of Copper (II) Oxide dissolves in water to form a weakly basic/alkaline **blue** solution of Copper (II) hydroxide.

Chemical equations

Copper (II) Oxide + Water -> Copper (II) hydroxide

$CuO(s) + H_2O(l) \rightarrow Cu(OH)_2$ **(s)**

(i)Reaction of non metals with Oxygen/air

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The following experiments show the reaction of non metals with Oxygen and air.

I. Burning Carbon

Procedure

(a)Using a pair of tongs hold a dry piece of charcoal on a Bunsen flame. Observe.

Place the products in a beaker containing about 2cm3 of water. Test the solution/mixture using litmus papers

(b)Using a pair of tongs hold another piece of dry charcoal on a Bunsen flame. Quickly lower it into a gas jar containing Oxygen gas .Observe.

Place about 2cm3 of water. Swirl. Test the solution/mixture using litmus papers

Observations

-Carbon **chars** then burns with a **blue** flame

-Colourless and odorless gas produced

-Solution formed turn blue litmus paper faint red.

Red litmus paper remains red.

Explanation

Carbon burns in air and faster in Oxygen with a blue non-sooty/non-smoky flame forming Carbon (IV) oxide gas.

Carbon burns in limited supply of air with a blue non-sooty/non-smoky flame forming Carbon (IV) oxide gas.

Carbon (IV) oxide gas dissolves in water to form weak acidic solution of Carbonic (IV) acid.

Chemical Equation

Carbon + Oxygen -> Carbon (IV) oxide
(excess air/oxygen)

$C(s) + O_2(g) \rightarrow CO_2(g)$ (in excess air)

Carbon + Oxygen -> Carbon (II) oxide
(limited air/oxygen)

$2C(s) + O_2(g) \rightarrow 2CO(g)$ (in limited air)

Carbon (IV) oxide + Water -> Carbonic (IV) acid

$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$ (very weak acid)

II. Burning Sulphur

Procedure

(a)Using a deflagrating spoon place sulphur powder on a Bunsen flame. Observe.

Place the products in a beaker containing about 3cm3 of water. Test the solution/mixture using litmus papers

(b) Using a deflagrating spoon place sulphur powder on a Bunsen flame. Slowly lower it into a gas jar containing Oxygen gas. Observe.

Place about 5cm3 of water. Swirl. Test the solution/mixture using litmus papers.

Observations

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-Sulphur burns with a **blue** flame
-Gas produced that has pungent choking smell
-Solution formed turn blue litmus paper faint red.
Red litmus paper remains red.

Explanation

Sulphur burns in air and faster in Oxygen with a blue non-sooty/non-smoky flame forming Sulphur (IV) oxide gas.
Sulphur (IV) oxide gas dissolves in water to form weak acidic solution of Sulphuric (IV) acid.

Chemical Equation

Sulphur + Oxygen -> Sulphur (IV) oxide
$S(s) + O_2(g) \rightarrow SO_2(g)$ (in excess air)
Sulphur (IV) oxide + Water -> Sulphuric (IV) acid
$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$ (very weak acid)

III. Burning Phosphorus

Procedure

(a)Remove a small piece of phosphorus from water and using a deflagrating spoon (with a lid cover) places it on a Bunsen flame.
Observe.
Carefully put the burning phosphorus to cover gas jar containing about 3cm3 of water. Test the solution/mixture using litmus papers

(b) Remove another small piece of phosphorus from water and using a deflagrating spoon (with a lid cover) place it on a Bunsen flame.
Slowly lower it into a gas jar containing Oxygen gas with about 5 cm3 of water.
Observe.
Swirl. Test the solution/mixture using litmus papers.

Observations

-Phosphorus catches fire before heating on Bunsen flame
-Dense white fumes of a gas produced that has pungent choking **poisonous** smell
-Solution formed turn blue litmus paper faint red.
Red litmus paper remains red.

Explanation

Phosphorus is stored in water. On exposure to air it instantaneously fumes then catch fire to burn in air and faster in Oxygen with a **yellow** flame producing dense white acidic fumes of Phosphorus (V) oxide gas.
Phosphoric (V) oxide gas dissolves in water to form weak acidic solution of Phosphoric (V) acid.

Chemical Equation

Phosphorus + Oxygen -> Phosphorous (V) oxide
$4P(s) + 5O_2(g) \rightarrow 2P_2O_5(s)$

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Phosphorous (V) oxide + Water -> Phosphoric (V) acid

$P_2O_5(s) + 3H_2O(l) \rightarrow 2H_3PO_4(aq)$ (very weak acid)

(e) Reactivity series/competition for combined Oxygen.

The reactivity series is a list of elements/metals according to their affinity for oxygen.

Some metals have higher affinity for Oxygen than others.

A metal/element with higher affinity for oxygen is placed higher/on top of the one less affinity.

The complete reactivity series of metals/elements

Element/Metal	**Symbol**
Potassium	K
Sodium	Na
Calcium	Ca
Magnesium	Mg
Aluminum	Al
Carbon	**C**
Zinc	**Zn**
Iron	**Fe**
Tin	**Sn**
Lead	**Pb**
Hydrogen	**H**
Copper	**Cu**
Mercury	**Hg**
Silver	**Ag**
Gold	**Au**
Platinum	**Pt**



Most reactive

Least reactive

Metals compete for combined Oxygen. A metal/element with higher affinity for oxygen removes Oxygen from a metal lower in the reactivity series/less affinity for Oxygen.

When a metal/element gains/acquire Oxygen, the **process** is called **Oxidation**.

When metal/element donate/lose Oxygen, the **process** is called **Reduction**.

An element/metal/compound that undergoes Oxidation is called **Reducing agent**.

An element/metal/compound that undergoes Reduction is called **Oxidizing agent**.

A reaction in which **both** Oxidation and Reduction take place is called a **Redox** reaction.

Redox reaction between Magnesium and copper (II) Oxide

Procedure

Place about 2g of copper (II) oxide in a crucible with a lid. Place another 2g of Magnesium powder into the crucible. Mix thoroughly.

Cover the crucible with lid. Heat strongly for five minutes.

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Allow the mixture to cool. Open the lid. Observe.

Observation

Colour change from black to brown. White solid power formed.

Explanation

Magnesium is higher in the reactivity series than Copper. It has therefore higher affinity for Oxygen than copper.

When a mixture of copper (II) oxide and Magnesium is heated, Magnesium reduces copper (II) oxide to brown copper metal and itself oxidized to Magnesium oxide. Magnesium is the reducing agent because it undergoes oxidation process.

Copper (II) oxide is the oxidizing agent because it undergoes **redox** reduction process.

The mixture should be cooled before opening the lid to prevent **hot** brown copper from being **reoxidized** back to black copper (II) oxide.

The reaction of Magnesium and Copper (II) oxide is a reaction

Chemical equation

Reduction process

Oxidation process

1. Copper (II) oxide + Magnesium -> Magnesium oxide + Copper
(black) (white ash/solid) (brown)
CuO(s) + Mg(s) -> Mg**O**(s) + **Cu**(s)
(Oxidizing Agent) (Reducing Agent)

2. Zinc (II) oxide + Magnesium -> Magnesium oxide + Zinc
(yellow when hot) (white ash/solid) (grey)
ZnO(s) + Mg(s) -> Mg**O**(s) + **Zn**(s)
(Oxidizing agent) (Reducing agent)

3. Zinc (II) oxide + Carbon -> Carbon (IV) oxide gas + Zinc
(yellow when hot) (colourless gas) (grey)
ZnO(s) + C(s) -> C**O₂** (g) + **Zn**(s)
(Oxidizing agent) (Reducing agent)

The reactivity series is used during extraction of metals from their ore. An ore is a rock containing mineral element which can be extracted for commercial purposes. Most metallic ores occur naturally as:

(i) **oxides** combined with Oxygen
(ii) **sulphides** combined with Sulphur
(iii) **carbonates** combined with carbon and Oxygen.

Metallic ores that naturally occur as metallic sulphides are first **roasted** in air to form the corresponding oxide. Sulphur (IV) oxide gas is produced. e.g.

Copper (I) sulphide + Oxygen -> Copper (I) Oxide + Sulphur (IV) oxide

$Cu_2S(s) + O_2(g) \rightarrow 2Cu(s) + SO_2(g)$

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Zinc (II) sulphide + Oxygen -> Zinc (II) Oxide + Sulphur (IV) oxide

$ZnS(s) + O_2(g) \rightarrow Zn(s) + SO_2(g)$

Lead (II) sulphide + Oxygen -> Lead (II) Oxide + Sulphur (IV) oxide

$PbS(s) + O_2(g) \rightarrow Pb(s) + SO_2(g)$

Iron (II) sulphide + Oxygen -> Iron (II) Oxide + Sulphur (IV) oxide

$FeS(s) + O_2(g) \rightarrow Fe(s) + SO_2(g)$

Metallic ores that naturally occur as metallic carbonates are first **heated** in air. They **decompose**/split to form the corresponding oxide and produce Carbon (IV) oxide gas. .e.g.

Copper (II) carbonate -> Copper (II) oxide + Carbon (IV) oxide

$CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$

Zinc (II) carbonate -> Zinc (II) oxide + Carbon (IV) oxide

$ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$

Lead (II) carbonate -> Lead (II) oxide + Carbon (IV) oxide

$PbCO_3(s) \rightarrow PbO(s) + CO_2(g)$

Iron (II) carbonate -> Iron (II) oxide + Carbon (IV) oxide

$FeCO_3(s) \rightarrow FeO(s) + CO_2(g)$

Metallic ores

WATER AND HYDROGEN

A.WATER

Pure water is a **colourless**, **odorless**, **tasteless**, neutral liquid. Pure water does not exist in nature but naturally in varying degree of purity. The main sources of water include rain, springs, borehole, lakes, seas and oceans:

Water is generally **used** for the following purposes:

(i) Drinking by animals and plants.
(ii) Washing clothes.
(iii) Bleaching and dyeing.
(iv) Generating hydroelectric power.
(v) Cooling industrial processes.

Water dissolves many substances/solutes.

It is therefore called **universal solvent**.

It contains about 35% **dissolved** Oxygen which support aquatic fauna and flora.

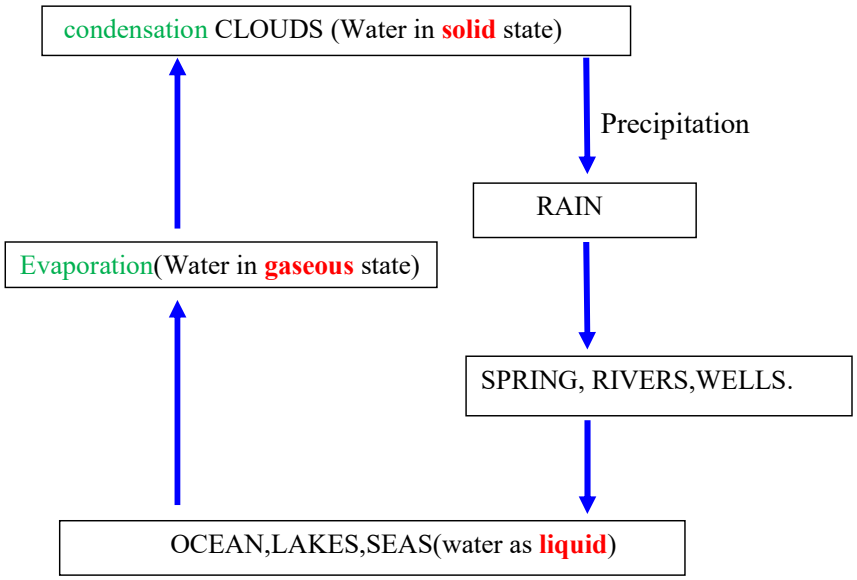
Water naturally exists in three phases/states **solid** ice, **liquid** water and **gaseous** water vapour.

The three states of water are naturally **interconvertible**.

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The natural interconvertion of the three phases/states of water forms the water cycle.

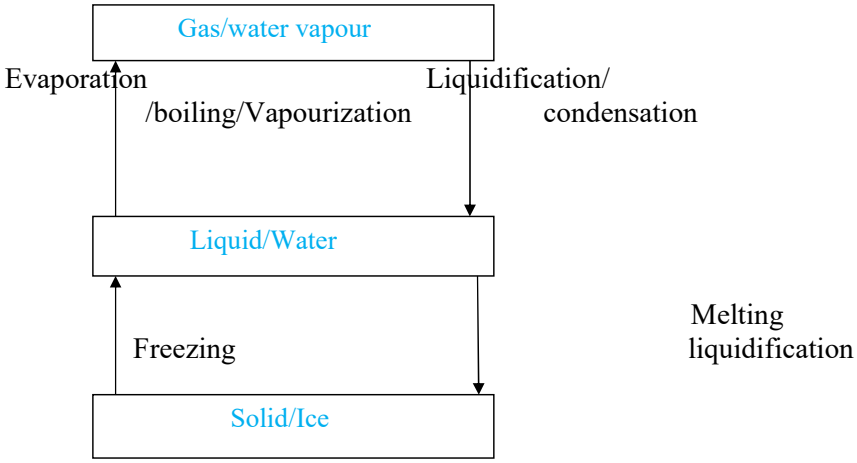


Liquid water in land, lakes, seas and oceans use the solar/sun **energy** to **evaporate/vapourize** to form water vapour/**gas**. Solar/sun energy is also used during transpiration by plants and respiration by animals.
During evaporation, the water vapour rises up the earth's surface. Temperatures decrease with height above the earth surface increase. Water vapour therefore cools as it rises up. At a height where it is cold enough to below 373Kelvin/100ºC Water vapour looses enough energy to form tiny droplets of liquid.
The process by which a gas/water vapour changes to a liquid is called **condensation/liquidification**.
On further cooling, the liquid looses more energy to form **ice/solid.** The process by which a liquid/water changes to a ice/solid is called **freezing/solidification**.
Minute/tiny ice/solid particles float in the atmosphere and coalesce/join together to form clouds. When the clouds become too heavy they fall to the earth's surface as rain/snow as the temperature increase with the fall.

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Interconversion of the three phases/states water



Solidification

Pure water has:

(i) fixed/constant/sharp freezing point/melting point of 273K/**0°C**

(ii) fixed/constant/sharp boiling point of 373K/**100°C** at **sea level**/1 atmosphere pressure

(iii) fixed density of **1gcm^{-3}**

This is the **criteria** of identifying pure/purity of water.

Whether a substance is water can be determined by using the following methods:

a) To test for presence of water using anhydrous copper (II) suphate (VI)

Procedure

Put about 2g of anhydrous copper (II) sulphate (VI) crystals into a clean test tube. Add three drops of tap water. Repeat the procedure using distilled water.

Observation

Colour changes from white to blue

Explanation

Anhydrous copper (II) sulphate (VI) is white. On adding water, anhydrous copper (II) sulphate (VI) gains/reacts with water to form **hydrated** copper (II) sulphate (VI).

Hydrated copper (II) sulphate (VI) is **blue**. Hydrated copper (II) sulphate (VI) contains water of crystallization.

The change of white **anhydrous** copper (II) sulphate (VI) to blue hydrated copper (II) sulphate (VI) is a confirmatory test for the **presence** of water

Chemical equation

Anhydrous copper (II) sulphate (VI) (white) + Water -> Hydrated copper (II) sulphate (VI) (**blue**)

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$CuSO_4(s)$ + 5H2O (l) -> **$CuSO_4.5H_2O(s)$**

b) To test for presence of water using anhydrous cobalt (II) chloride

Procedure

Put about 5cm3 of water into a clean test tube.
Dip a dry anhydrous cobalt (II) chloride **paper** into the test tube.
Repeat the procedure using distilled water.

Observation

Colour changes from blue to **pink**

Explanation

Anhydrous cobalt (II) chloride is **blue**. On adding water, **anhydrous** cobalt (II) chloride gains/reacts with water to form **hydrated** cobalt (II) chloride.
Hydrated cobalt (II) chloride is **pink**.
Hydrated cobalt (II) chloride contains water of crystallization.
The change of blue **anhydrous** cobalt (II) chloride to pink hydrated cobalt (II) chloride is a confirmatory test for the **presence** of water **Chemical equation.**

Anhydrous cobalt (II) chloride (**Blue**) + Water -> Hydrated cobalt (II) chloride (**pink**)

$CoCl_2$ (s) + 5H2O (l) -> **$CoCl_2.5H_2O(s)$**

Burning a candle in air

Most organic substances/fuels burn in air to produce water. Carbon (IV) oxide gas is also produced if the air is sufficient/excess.

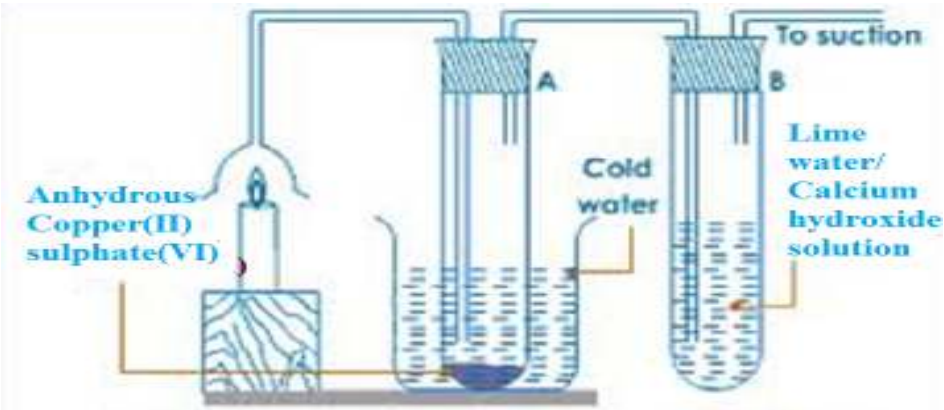
Procedure

Put about 2g of anhydrous copper (II) sulphate (VI) crystals in a boiling tube.
Put about 5cm3 of lime water in a boiling tube.
Light a small candle stick. Place it below an inverted thistle/filter funnel
Collect the products of the burning candle by setting the apparatus as below

Set up of apparatus

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To demonstrate that water and carbon(IV)oxide gas are formed when a candle burns in air

Observation

The sanction pump pulls the products of burning into the inverted funnel. Colour of anhydrous copper (II) sulphate (VI) changes from white to blue. A **white precipitate** is formed in the lime water/calcium hydroxide.

Explanation

When a candle burn it forms a water and carbon (IV) oxide.

Water turns anhydrous copper (II) sulphate (VI) changes from white to blue.

Carbon (IV) oxide gas forms **white precipitate** when bubbled in lime water/calcium hydroxide.

Since:

(i) hydrogen in the wax burn to form water

Hydrogen + Oxygen -> Water
(from candle) (from the air)

$2H_2$ (g) + O_2 (g) -> $2H_2O$ (g/l)

(ii) carbon in the wax burn to form carbon (IV) oxide

Hydrogen + Oxygen -> Water
(from candle) (from the air)

C(s) + O_2(g) -> CO_2 (g)

The candle before burning therefore contained only **Carbon** and **Hydrogen only.**

A compound made up of **hydro**gen and carbon is called **Hydro**carbon.

A candle is a hydrocarbon.

Other hydrocarbons include: Petrol, diesel, Kerosene, and Laboratory gas.

Hydrocarbons burn in air to form water and carbon (IV) oxide gas.

Hydrocarbons + Oxygen -> Water + Oxygen

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Water pollution

Water pollution takes place when undesirable substances are added into the water.
Sources of water pollution include:
(i)Industrial chemicals being disposed into water bodies like rivers, lakes and oceans.
(ii)Discharging untreated /raw sewage into water bodies.
(iii)Leaching of insecticides/herbicides form agricultural activities into water bodies.
(iv)Discharging non-biodegradable detergents after domestic and industrial use into water bodies.
(v)Petroleum oil spilling by ships and oil refineries
(vi)Toxic/poisonous gases from industries dissolving in rain.
(vii) Acidic gases from industries dissolving in rain to form "acid rain"
(viii)Discharging hot water into water bodies. This reduces the quantity of dissolved Oxygen in the water killing the aquatic fauna and flora.
Water pollution can be reduced by:
(i) Reducing the use of agricultural fertilizers and chemicals in agricultural activities.
(ii) Use of biological control method instead of insecticides and herbicides
(iii) Using biodegradable detergents

REACTION OF WATER WITH METALS.

Some metals react with water while others do not. The reaction of metals with water depends on the reactivity series. The higher the metal in the reactivity series the more reactive the metal with water .The following experiments shows the reaction of metals with cold water and water vapour/steam.

(a)Reaction of sodium/ potassium with cold water:

Procedure

Put about 500cm3 of water in a beaker. Add three drops of phenolphthalein indicator/litmus solution/universal indicator solution/methyl orange indicator into the water.
Cut a **very small** piece of sodium .Using a pair of forceps put the metal into the water.

Observation

Sodium melts to a silvery ball that floats and darts on the surface decreasing in size.
Effervescence/fizzing/ bubbles of colourless gas produced.
Colour of phenolphthalein turns **pink**
Colour of litmus solution turns **blue**
Colour of methyl orange solution turns **Orange**
Colour of universal indicator solution turns **blue**

Explanation

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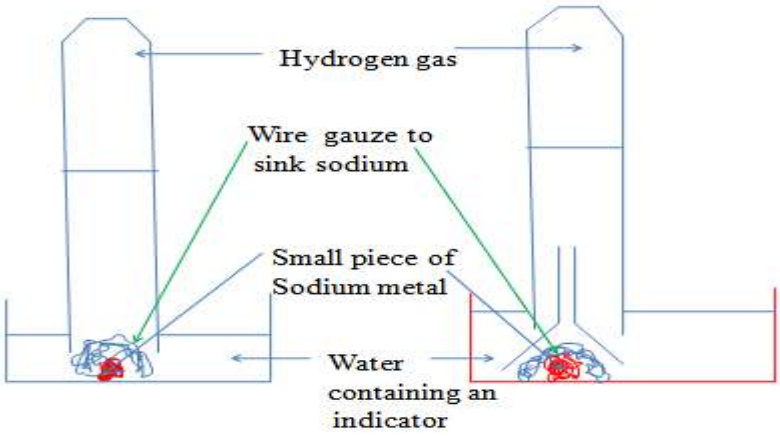
Sodium is less dense than water. Sodium floats on water and vigorously reacts to form an **alkaline** solution of sodium hydroxide and producing hydrogen gas. Sodium is thus stored in paraffin to prevent **contact** with water.

Chemical equation

Sodium + Water -> Sodium hydroxide + Hydrogen gas

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

To collect hydrogen gas, Sodium metal is forced to **sink** to the bottom of the trough/beaker by wrapping it in wire gauze/mesh.



Potassium is more reactive than Sodium. On contact with water it **explodes**/burst into flames. An alkaline solution of potassium hydroxide is formed and hydrogen gas

Chemical equation

Potassium + Water -> Potassium hydroxide + Hydrogen gas

$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$

Caution: Reaction of Potassium with water is very risky to try in a school laboratory.

(b)Reaction of Lithium/ Calcium with cold water:

Procedure

Put about 200cm3 of water in a beaker. Add three drops of phenolphthalein indicator/litmus solution/universal indicator solution/methyl orange indicator into the water.

Cut a small piece of Lithium .Using a pair of forceps put the metal into the water. Repeat with a piece Calcium metal

Observation

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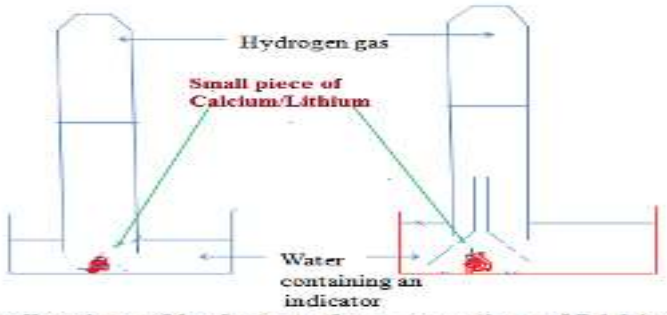
Lithium sinks to the bottom of the water. Rapid effervescence/fizzing/ bubbles of colourless gas produced.
Colour of phenolphthalein turns **pink**
Colour of litmus solution turns **blue**
Colour of methyl orange solution turns **Orange**
Colour of universal indicator solution turns **blue**

Explanation

Lithium and calcium are **denser** than water. Both sink in water and vigorously react to form an **alkaline** solution of Lithium hydroxide / calcium hydroxide and producing hydrogen gas. Lithium is more reactive than calcium. It is also stored in paraffin like Sodium to prevent **contact** with water.

Chemical equation

Lithium + Water -> Lithium hydroxide + Hydrogen gas
$2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$
Calcium + Water -> Calcium hydroxide + Hydrogen gas
$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$



Collection of hydrogen from reaction of Lithium /calcium with water

(c) Reaction of Magnesium/Zinc/ Iron with Steam/water vapour:

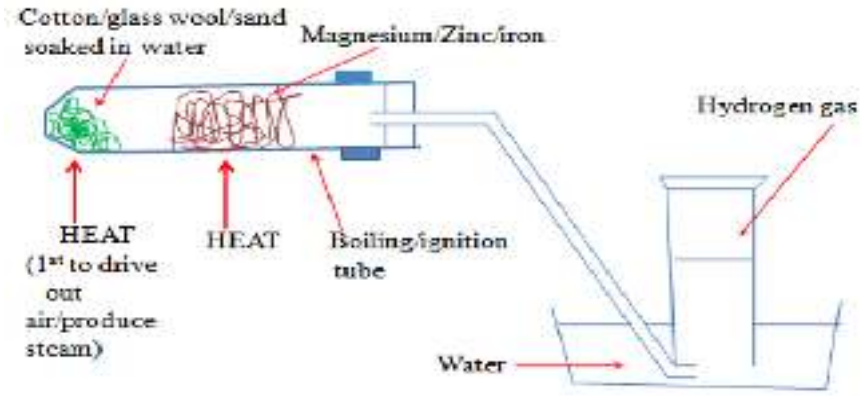
Procedure method1

Place some wet sand or cotton/glass wool soaked in water at the bottom of an ignition/hard glass boiling tube.
Polish magnesium ribbon using sand paper.
Coil it at the centre of the ignition/hard glass boiling tube.
Set up the apparatus as below.
Heat the wet sand or cotton/glass wool soaked in water gently to:
(i) Drive away air in the ignition/hard glass boiling tube.
(ii) Generate steam
Heat the coiled ribbon strongly using another burner. Repeat the experiment using Zinc powder and fresh Iron filings.

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Set up of apparatus



Reaction of Steam/water vapour with Magnesium /Zinc/iron

Observations

(i)With Magnesium ribbon:

The Magnesium glows with a bright flame (and continues to burn even if heating is stopped)

White solid /ash formed

White solid /ash formed dissolve in water to form a colourless solution

Colourless gas produced/collected that extinguish burning splint with "pop sound"

(ii) With Zinc powder:

The Zinc powder turns red hot on strong heating

Yellow solid formed that turn white on cooling

White solid formed on cooling does not dissolve in water.

(iii)With Iron fillings:

The Iron fillings turn red hot on strong heating

Dark blue solid formed

Dark blue solid formed does not dissolve in water.

Procedure method 2

Put some water in a round bottomed flask

Polish magnesium ribbon using sand paper.

Coil it at the centre of a hard glass tube

Set up the apparatus as below.

Heat water strongly to boil so as to:

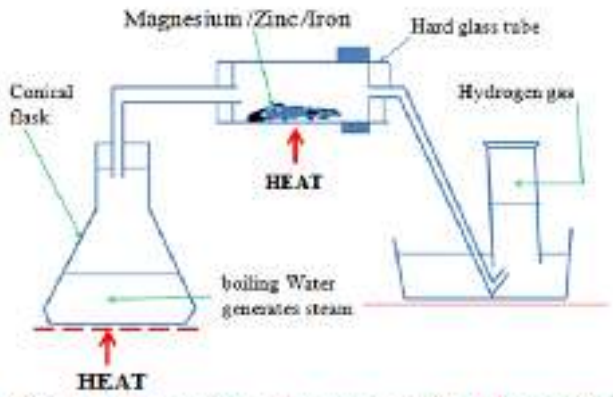
(i) drive away air in the glass tube.

(ii) generate steam

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Heat the coiled ribbon strongly using another burner. Repeat the experiment using Zinc powder and fresh Iron filings.



Method 2:Reaction of steam with Magnesium / Zinc/Iron

Observations

(i)With Magnesium ribbon:

The Magnesium glows with a bright flame (and continues to burn even if heating is stopped)

White solid /ash formed

White solid /ash formed dissolve in water to form a colourless solution

Colourless gas produced/collected that extinguish burning splint with "pop sound"

(ii) With Zinc powder:

The Zinc powder turns red hot on strong heating

Yellow solid formed that turn white on cooling

White solid formed on cooling does not dissolve in water.

(iii)With Iron fillings:

The Iron fillings turn red hot on strong heating

Dark blue solid formed

Dark blue solid formed does not dissolve in water.

Explanations

(**a**)Hot magnesium burn vigorously in steam. The reaction is highly exothermic generating enough heat/energy to proceed without further heating.

White Magnesium oxide solid/ash is left as residue.

Hydrogen gas is produced .It extinguishes a burning splint with a "pop sound".

Chemical Equation

Magnesium + Steam -> Magnesium oxide + Hydrogen

$Mg(s) + H_2O(\mathbf{g}) \rightarrow MgO(s) + H_2(g)$

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Magnesium oxide reacts /dissolves in water to form an alkaline solution of Magnesium oxide

Chemical Equation

Magnesium oxide + Water -> Magnesium hydroxide

$MgO(s) + H_2O(\mathbf{l}) \rightarrow Mg(OH)_2\ (\mathbf{aq})$

(**b**)Hot Zinc react vigorously in steam forming yellow Zinc oxide solid/ash as residue which cools to white.

Hydrogen gas is produced .It extinguishes a burning splint with a "pop sound".

Chemical Equation

Zinc + Steam -> Zinc oxide + Hydrogen

$Zn(s) + H_2O(\mathbf{g}) \rightarrow ZnO(s) + H_2(g)$

Zinc oxide does not dissolve in water.

(**c**)Hot Iron reacts with steam forming dark blue tri iron tetra oxide solid/ash as residue.

Hydrogen gas is produced .It extinguishes a burning splint with a "pop sound".

Chemical Equation

Iron + Steam -> Tri iron tetra oxide + Hydrogen

$2Fe(s) + 4H_2O(\mathbf{g} \rightarrow Fe_2O_4(s) + 4H_2(g)$

Tri iron tetra oxide does not dissolve in water.

(**d**)Aluminum reacts with steam forming an **insoluble coat**/cover of **impervious** layer of aluminum oxide on the surface preventing further reaction.

(**e**) Lead, Copper, Mercury, Silver, Gold and Platinum do **not** react with either water or steam.

HYDROGEN

Occurrence

Hydrogen does not occur free in nature. It occurs as Water and in Petroleum.

School laboratory Preparation

Procedure

Put Zinc granules in a round/flat/conical flask. Add dilute sulphuric (VI) /Hydrochloric acid.

Add about 3cm3 of copper (II) sulphate (VI) solution.

Collect the gas produced over water as in the set up below.

Discard the first gas jar. Collect several gas jars.

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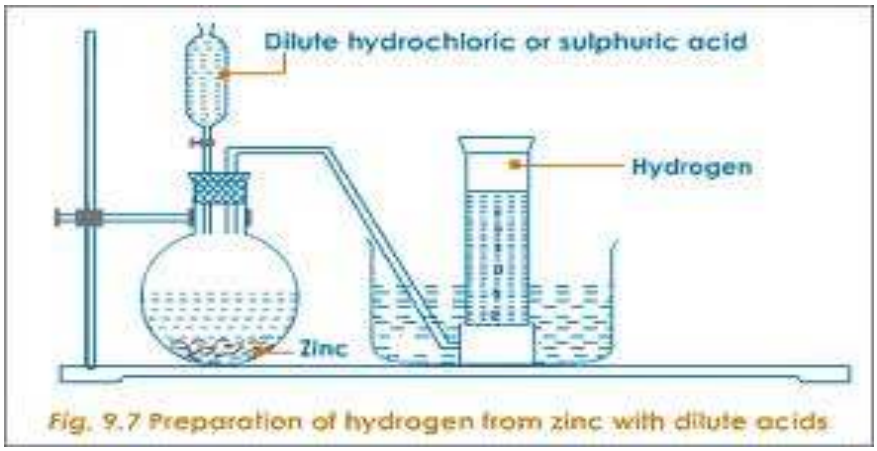


Fig. 9.7 Preparation of hydrogen from zinc with dilute acids

Observation/Explanation

Zinc reacts with dilute sulphuric (VI)/hydrochloric acid to form a salt and produce hydrogen gas.

When the acid comes into contact with the metal, there is rapid effervescence/ bubbles /fizzing are produced and a colourless gas is produced that is collected:

(i) Over water because it is insoluble in water

(ii) Through downward displacement of air/upward delivery because it is less dense than air.

The first gas jar is impure. It contains air that was present in the apparatus.

Copper (II) sulphate (VI) solution act as catalyst.

Chemical equation

(a) Zinc + Hydrochloric acid -> Zinc chloride + Hydrogen

$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Ionic equation

$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

Zinc + Sulphuric (VI) acid -> Zinc Sulphate (VI) + Hydrogen

$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

Ionic equation

$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

(b) Chemical equation

Magnesium + Hydrochloric acid -> Magnesium chloride + Hydrogen

$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

Ionic equation

$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

Magnesium + Sulphuric (VI) acid -> Magnesium Sulphate(VI) + Hydrogen

$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

Ionic equation

$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

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(c) Chemical equation

Iron + Hydrochloric acid -> Iron (II) chloride + Hydrogen

$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$

Ionic equation

$Fe(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$

Iron + Sulphuric (VI) acid -> Iron (II) Sulphate (VI) + Hydrogen

$Fe(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2(g)$

Ionic equation

$Fe(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$

Note

1. Hydrogen cannot be prepared from reaction of:

(i)Nitric (V) acid and a metal. Nitric (V) acid is a strong oxidizing agent. It **oxidizes** hydrogen gas to **water**.

(ii) Dilute sulphuric (VI) acid with calcium/Barium/Lead because Calcium sulphate (VI), Barium sulphate (VI) and Lead (II) sulphate (VI) salts formed are insoluble. Once formed, they **cover/coat** the **unreacted** calcium/Barium/Lead **stopping** further reaction and producing very small amount/volume of hydrogen gas.

(iii) Dilute acid with sodium/potassium. The reaction is **explosive.**

Properties of Hydrogen gas

(a)Physical properties

1. Hydrogen is a **neutral**, colourless and **odorless** gas. When mixed with air it has a characteristic pungent choking smell
2. It is insoluble in water thus can be collected over water.
3. It is the lightest known gas. It can be transferred by inverting one gas jar over another.

(b)Chemical properties

(i)Burning

I. Hydrogen does not support burning/combustion. When a burning splint is inserted into a gas jar containing Hydrogen, the flame is extinguished /put off.

II. Pure dry hydrogen burn with a blue quiet flame to form water. When a stream of pure dry hydrogen is ignited, it catches fire and continues to burn with a blue flame.

III. Impure (air mixed with) hydrogen burns with an explosion. Small amount/ volume of air **mixed** with hydrogen in a test tube produce a small explosion as a "pop" sound. This is the confirmatory test for the presence of Hydrogen gas. A gas that burns with a "pop" sound is confirmed to be Hydrogen.

(ii)Redox in terms of Hydrogen transfer

Redox can also be defined in terms of Hydrogen transfer.

(i)Oxidation is removal of Hydrogen

(ii)Reduction is addition of Hydrogen

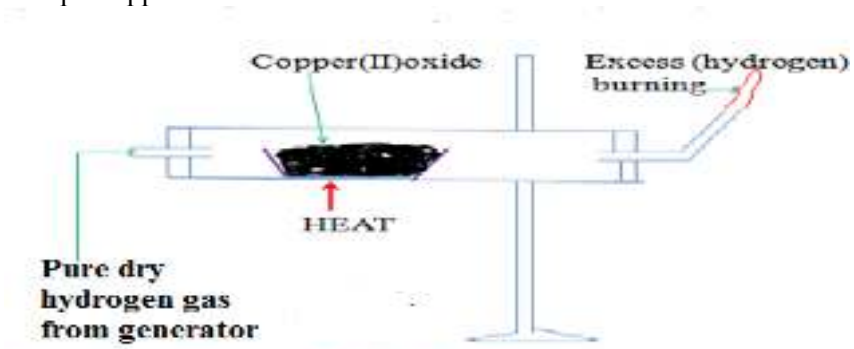
(iii)Redox is simultaneous addition and removal of Hydrogen

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Example

When a stream of dry hydrogen gas is passed through black copper (II) oxide, hydrogen gas gains the oxygen from copper (II) oxide.
Black copper (II) oxide is reduced to brown copper metal.
Black copper (II) oxide thus the Oxidizing agent.
Hydrogen gas is oxidized to Water. Hydrogen is the Reducing agent.
Set up of apparatus



Reduction of Copper(II)Oxide by Hydrogen gas

(**a**)Chemical equation
(i) In glass tube

Copper (II) Oxide + Hydrogen -> Copper + Hydrogen gas
(oxidizing agent) (reducing agent)
(**black**) (brown)

$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$

(ii) when excess Hydrogen is burning.

Oxygen + Hydrogen -> Water

$O_2(g) + 2H_2(g) \rightarrow 2H_2O(l)$

(**b**)Chemical equation
(i) In glass tube

Lead (II) Oxide + Hydrogen -> Lead + Hydrogen gas
(oxidizing agent) (reducing agent)
(brown when hot/ yellow when cool) (grey)

$PbO(s) + H_2(g) \rightarrow Pb(s) + H_2O(l)$

(ii) when excess Hydrogen is burning.

Oxygen + Hydrogen -> Water

$O_2(g) + 2H_2(g) \rightarrow 2H_2O(l)$

(**c**)Chemical equation
(i) In glass tube

Iron (III) Oxide + Hydrogen -> Iron + Hydrogen gas

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(oxidizing agent) (reducing agent)

(Dark grey) (grey)

Fe_2O_3 (s) + $3H_2$ (g) -> Fe(s) + $3H_2O$ (l)

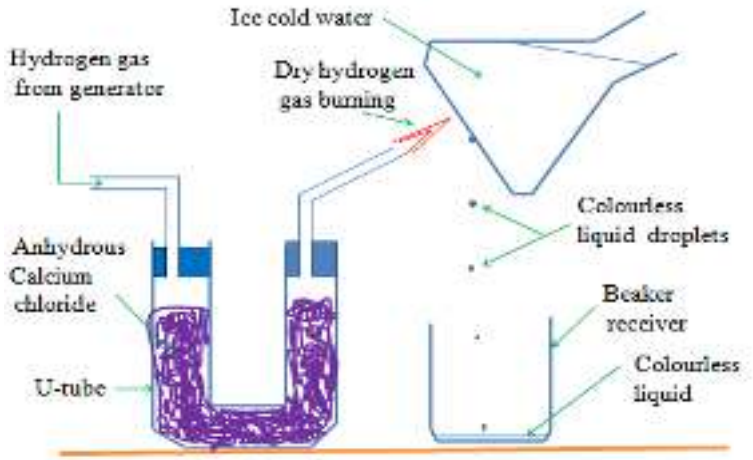
(ii) when excess Hydrogen is burning.

Oxygen + Hydrogen -> Water

O_2 (g) + $2H_2$ (g) -> $2H_2O$ (l)

(iii) Water as an Oxide as Hydrogen

Burning is a reaction of an element with Oxygen. The substance formed when an element burn in air is the oxide of the element. When hydrogen burns, it reacts/ combines with Oxygen to form the **oxide of Hydrogen**. The oxide of Hydrogen is called water. Hydrogen is first dried because a mixture of Hydrogen and air explode. The gas is then ignited .The products condense on a cold surface/flask containing a freezing mixture. A freezing mixture is a mixture of water and ice.



Burning Hydrogen:To show Water is an Oxide of Hydrogen

The condensed products are collected in a receiver as a colourless liquid.

Tests

(**a**) When about 1g of **white** anhydrous copper (II) sulphate (VI) is added to a sample of the liquid, it turns to **blue**. This confirms the liquid formed is water.

(**b**) When blue anhydrous cobalt (II) chloride paper is dipped in a sample of the liquid, it turns to **pink**. This confirms the liquid formed is water.

(**c**)When the liquid is heated to boil, its **boiling point** is **100°C** at sea level/one atmosphere pressure. This confirms the liquid is **pure water.**

Uses of Hydrogen gas

1. Hydrogenation/Hardening of unsaturated vegetable oils to saturated fats/margarine.

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When Hydrogen is passed through unsaturated compounds in presence of **Nickel** catalyst and about **150°C**, they become saturated. Most vegetable oil is unsaturated liquids at room temperature. They become saturated and hard through hydrogenation.

2. In weather forecast balloons.

Hydrogen is the lightest known gas**.** Meteorological data is collected for analysis by sending hydrogen filled weather balloons to the atmosphere. The data collected is then used to forecast weather conditions.

3. In the Haber process for the manufacture of Ammonia

Hydrogen is mixed with Nitrogen in presence of Iron catalyst to form Ammonia gas. Ammonia gas is a very important raw material for manufacture of agricultural fertilizers.

4. In the manufacture of Hydrochloric acid.

Limited volume/amount of Hydrogen is burnt in excess chlorine gas to form Hydrogen chloride gas. Hydrogen chloride gas is dissolved in water to form Hydrochloric acid. Hydrochloric acid is used in pickling/washing metal surfaces.

5. As rocket fuel.

Fixed proportions of Hydrogen and Oxygen when ignited explode violently producing a lot of energy/heat. This energy is used to power/propel a rocket to space.

6. In oxy-hydrogen flame for welding.

A cylinder containing Hydrogen when ignited in pure Oxygen from a second cylinder produces a flame that is very hot. It is used to cut metals and welding.

Sample revision questions

1. A colourless liquid was added anhydrous copper (II) sulphate (VI) which turned blue.

(a)Why is it wrong to conclude the liquid was pure water?

Anhydrous copper (II) sulphate (VI) test for presence of water. Purity of water is determined from freezing/melting/boiling point.

(b)Write an equation for the reaction that takes place with anhydrous copper (II) sulphate (VI)

Anhydrous copper (II) sulphate (VI) + Water -> hydrated copper (II) sulphate (VI)

$CuSO_4(s) + 5H2O (l) \rightarrow CuSO_4.5H_2O(s)$

(c)(i)Which other compound would achieve the same results as anhydrous copper (II) sulphate (VI)

Anhydrous cobalt (II) chloride/$CoCl_2.6H_2O$

(ii)Write the equation for the reaction

Anhydrous cobalt (II) chloride + Water -> hydrated cobalt (II) chloride

$CoCl_2 (s) + 6H_2O(l) \rightarrow CoCl_2.6H_2O (s)$

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(d)Complete the equation

(i) Sulphur (VI) oxide + Water -> Sulphuric (VI) acid
(ii) Sulphur (IV) oxide + Water -> Sulphuric (IV) acid
(iii) Carbon (IV) oxide + Water -> Carbonic (IV) acid
(iv) Nitrogen (IV) oxide + Water -> Nitric (V) acid
(v) Phosphorus (V) oxide + Water -> Phosphoric (V) acid
(vi) Sodium oxide + Water -> Sodium hydroxide
(vi) Sodium peroxide + Water -> Sodium hydroxide

2. Metal B reacts with steam. Metal C reacts with cold water. Metal A does not react with water.

(a)Arrange the metals as they should appear in the reactivity series.

B
C
A

(b)A product residue in D which was brown when hot but turned yellow on cooling during the reaction of metal B was formed. Gas E was also evolved. Identify

(i)Metal B Lead/Pb
(ii)Residue D Lead (II) oxide/PbO
(iii)Gas E Hydrogen/H_2

(c)A portion of product residue in D was added dilute nitric (V) acid. Another portion of product residue in D was added dilute sulphuric (VI) acid. State and explain the observations made.

When added dilute nitric (V) acid, D dissolves to form a colourless solution.

Lead (II) Oxide + dilute nitric (V) acid -> Lead (II) nitrate (V) + Water

$PbO\ (s) + 2HNO_3\ (aq) \rightarrow Pb(NO_3)_2\ (\mathbf{aq}) + H_2O\ (l)$

When added dilute sulphuric (VI) acid, D does not dissolve. A white suspension/precipitate was formed. Lead(II)Oxide reacts with sulphuric(VI)acid to form insoluble Lead(II)sulphate(VI) that cover/coat unreacted Lead(II)Oxide, stopping further reaction.

Lead (II) Oxide + dilute sulphuric (VI) acid -> Lead (II) sulphate (VI) + Water

$PbO\ (s) + H_2SO_4\ (aq) \rightarrow PbSO_4\ (\mathbf{s}) + H_2O\ (l)$

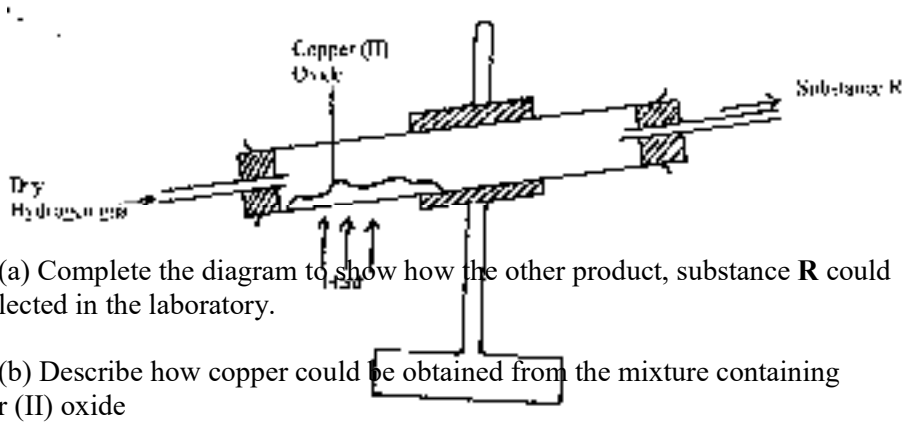
3. (a) Hydrogen can reduce copper (II) Oxide but not alluminium oxide. Explain
(b) When water reacts with potassium metal the hydrogen produced ignites explosively on the surface of water.

(i) What causes this ignition?
(ii) Write an equation to show how this ignition occurs

2. In an experiment, dry hydrogen gas was passed over hot copper (II) oxide in a combustion tube as shown in the diagram below:

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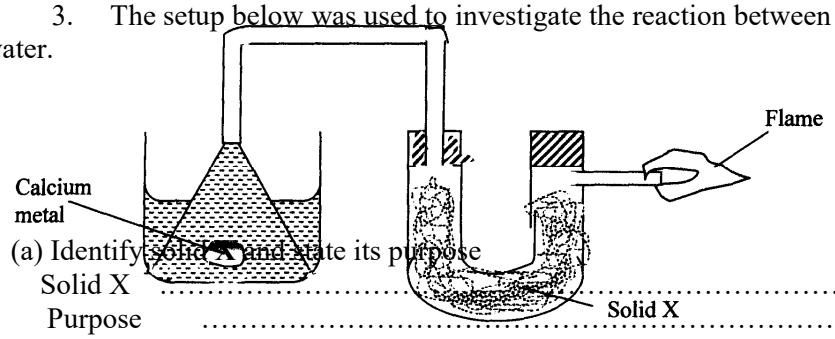
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(a) Complete the diagram to show how the other product, substance **R** could be collected in the laboratory.

(b) Describe how copper could be obtained from the mixture containing copper (II) oxide

3. The setup below was used to investigate the reaction between metals and water.



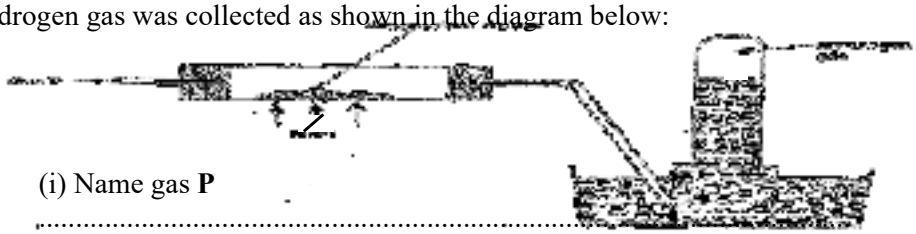
(a) Identify solid **X** and state its purpose

Solid X ……..

Purpose ……..

(b) Write a chemical equation for the reaction that produces the flame.

4. Gas **P** was passed over heated magnesium ribbon and hydrogen gas was collected as shown in the diagram below:



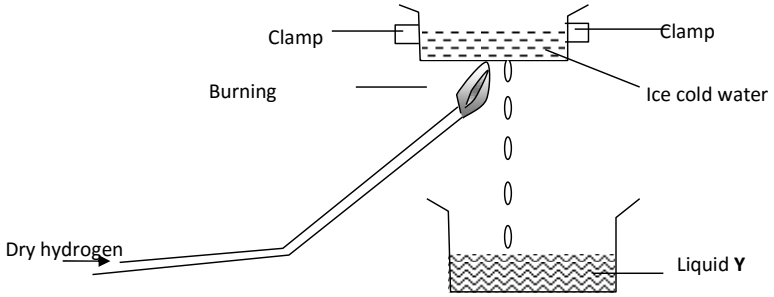
(i) Name gas **P**

……

(ii) Write an equation of the reaction that takes place in the combustion tube

(iii) State **one** precaution necessary at the end of this experiment

5. When hydrogen is burnt and the product cooled, the following results are obtained as shown in the diagram below:



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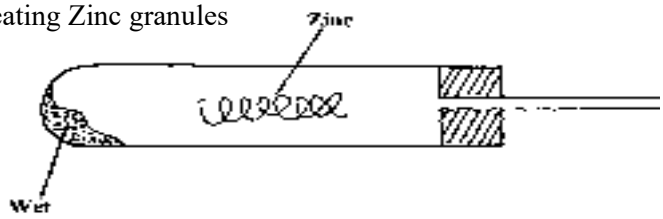
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(a) Write the equation for the formation of liquid **Y**

(b) Give a chemical test for liquid **Y**

Jane set-up the experiment as shown below to collect a gas. The wet sand was heated before

heating Zinc granules

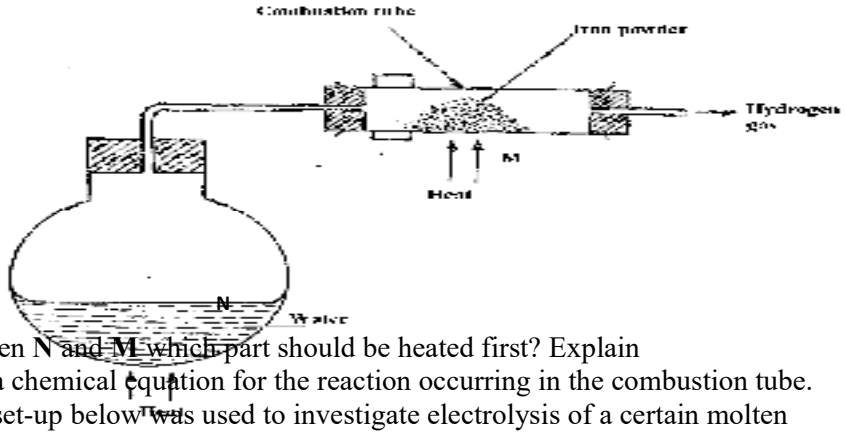


Wet sand

(a) Complete the diagram for the laboratory preparation of the gas

(b) Why was it necessary to heat wet sand before heating Zinc granules?

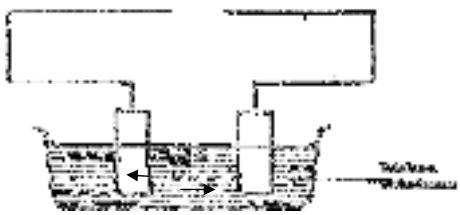
7.



(a) Between N and M which part should be heated first? Explain

(b) Write a chemical equation for the reaction occurring in the combustion tube.

8. The set-up below was used to investigate electrolysis of a certain molten compound;-



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(a) Complete the circuit by drawing the cell in the gap left in the diagram

(b) Write half-cell equation to show what happens at the cathode

(c) Using an arrow show the direction of electron flow in the diagram above

9. Hydrogen can be prepared by reacting zinc with dilute hydrochloric acid.

a) Write an equation for the reaction.

b) Name an appropriate drying agent for hydrogen gas.

c) Explain why copper metal cannot be used to prepare hydrogen gas.

d) Hydrogen burns in oxygen to form an oxide.

(i) Write an equation for the reaction.

(ii) State **two** precautions that must be taken before the combustion begins and at the end of the combustion.

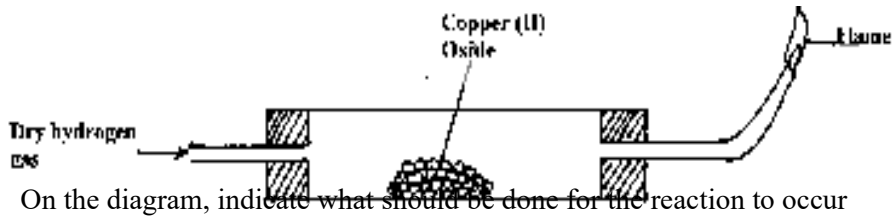
e) Give **two** uses of hydrogen gas.

f) When zinc is heated to redness in a current of steam, hydrogen gas is obtained. Write an equation for the reaction.

g) Element **Q** reacts with dilute acids but not with cold water. Element **R** does not react with dilute acids. Elements **S** displaces element **P** from its oxide. **P** reacts with cold water. Arrange the four elements in order of their reactivity, starting with the most reactive.

h) Explain how hydrogen is used in the manufacture of margarine.

10. a) The set-up below is used to investigate the properties of hydrogen.



i) On the diagram, indicate what should be done for the reaction to occur

ii) Hydrogen gas is allowed to pass through the tube for some time before it is lit. Explain

iii) Write an equation for the reaction that occurs in the combustion tube

iv) When the reaction is complete, hydrogen gas is passed through the apparatus until they cool down. Explain

v) What property of hydrogen is being investigated?

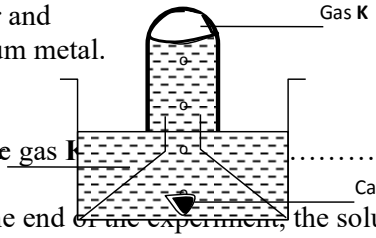
vi) What observation confirms the property stated in **(v)** above?

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vii) Why is zinc oxide not used to investigate this property of hydrogen gas?

11. The set up below was used to collect gas **K,** produced by the reaction between water and calcium metal.



(a) Name gas **K** .....

(b) At the end of the experiment, the solution in the beaker was found to be a weak base. Explain why the solution is a weak base

ACIDS, BASES AND INDICATORS

INTRODUCTION TO ACIDS, BASES AND INDICATORS

1. In a school laboratory:

(i)An acid may be defined as a substance that turns litmus red.

(ii)A base may be defined as a substance that turns litmus blue.

Litmus is lichen found mainly in West Africa. It changes its colour depending on whether the solution it is in, is basic/alkaline or acidic. It is thus able to identify/show whether another substance is an acid, base or neutral.

(iii)An indicator is a substance that shows whether another substance is a base/alkaline,acid or neutral.

2. Common naturally occurring acids include:

Name of acid	**Occurrence**
1.Citric acid	Found in ripe citrus fruits like passion fruit/oranges/lemon
2.Tartaric acid	Found in grapes/baking powder/health salts
3.Lactic acid	Found in sour milk
4.Ethanoic acid	Found in vinegar

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5.Methanoic acid	Present in ants, bees stings
6.Carbonic acid	Used in preservation of fizzy drinks like coke, Lemonade, Fanta
7.Butanoic acid	Present in cheese
8.Tannic acid	Present in tea

3. Most commonly used acids found in a school laboratory are not naturally occurring. They are manufactured. They are called **mineral acids**.
Common mineral acids include:

Name of mineral acid	**Common use**
Hydrochloric acid (HCl)	Used to clean/pickling surface of metals Is found in the stomach of mammals/human beings
Sulphuric(VI) acid (H_2SO_4)	Used as acid in car battery, making battery, making fertilizers
Nitric(V)acid (HNO_3)	Used in making fertilizers and explosives

4. Mineral acids are manufactured to very high concentration. They are **corrosive** (causes painful wounds on contact with the skin) and attack/reacts with garments/clothes/metals.
In a school laboratory, they are mainly used when added a lot of water. This is called **diluting**. Diluting ensures the concentration of the acid is safely low.

5. Bases are opposite of acids. Most bases do not dissolve in water.
Bases which dissolve in water are called **alkalis**.
Common alkalis include:

Name of alkali	**Common uses**
Sodium hydroxide (NaOH)	Making soaps and detergents
Potassium hydroxide(KOH)	Making soaps and detergents
Ammonia solution(NH_4OH)	Making fertilizers, softening hard water

Common bases (which are not alkali) include:

Name of base	Common name
Magnesium oxide/hydroxide	Anti acid to treat indigestion
Calcium oxide	Making cement and neutralizing soil acidity

6. Indicators are useful in identifying substances which look-alike.
An acid-base indicator is a substance used to identify whether another substance is alkaline or acidic.
An acid-base indicator works by changing to different colors in neutral, acidic and alkaline **solutions/dissolved** in water.
Experiment: To prepare simple acid-base indicator

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Procedure

(a)Place some flowers petals in a mortar. Crush them using a pestle. Add a little sand to assist in crushing.

Add about 5cm3 of propanone/ethanol and carefully continue grinding.

Add more 5cm3 of propanone/ethanol and continue until there is enough extract in the mortar.

Filter the extract into a clean 100cm3 beaker.

(b)Place 5cm3 of filtered wood ash, soap solution, ammonia solution, sodium hydroxide, hydrochloric acid, distilled water, sulphuric (VI) acid, sour milk, sodium chloride, toothpaste and calcium hydroxide into separate test tubes.

(c)Put about three drops of the extract in (a)to each test tube in (b). Record the observations made in each case.

Sample observations

Solution mixture	**Colour on adding indicator extract**	**Nature of solution**
wood ash	green	Base/alkaline
soap solution	green	Basic/alkaline
ammonia solution	green	Basic/alkaline
sodium hydroxide	green	Basic/alkaline
hydrochloric acid	Red	Acidic
distilled water	orange	Neutral
sulphuric(VI)acid	Red	Acidic
sour milk	green	Basic/alkaline
sodium chloride	orange	Neutral
Toothpaste	green	Basic/alkaline
calcium hydroxide	green	Basic/alkaline
Lemon juice	Red	Acidic

The plant extract is able to differentiate between solutions by their nature. It is changing to a similar colour for similar solutions.

(i)Since lemon juice is a known acid, then sulphuric (VI) and hydrochloric acids are similar in nature with lemon juice because the indicator shows similar colors. They are acidic in nature.

(ii)Since sodium hydroxide is a known base/alkali, then the green colour of indicator shows an alkaline/basic solution.

(iii) Since pure water is neutral, then the orange colour of indicator shows neutral solutions.

7. In a school laboratory, commercial indicators are used. A commercial indicator is cheap, readily available and easy to store. Common indicators include: Litmus, phenolphthalein, methyl orange, screened methyl orange, bromothymol blue.

Experiment:

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Using commercial indicators to determine acidic, basic/alkaline and neutral solutions

Procedure

Place 5cm3 of the solutions in the table below. Add three drops of litmus solution to each solution.

Repeat with phenolphthalein indicator, methyl orange, screened methyl orange and bromothymol blue.

Sample results

Substance/ Solution	Indicator used				
	Litmus	**Phenolphthalein**	**Methyl orange**	**Screened methyl orange**	**Bromothymol blue**
wood ash	Blue	Pink	Yellow	Orange	Blue
soap solution	Blue	Pink	Yellow	Orange	Blue
ammonia solution	Blue	Pink	Yellow	Orange	Blue
sodium hydroxide	Blue	Pink	Yellow	Orange	Blue
hydrochloric acid	Red	Colourless	Red	Purple	Orange
distilled water	Colourless	Colourless	Red	Orange	Orange
sulphuric(VI)acid	Red	Colourless	Red	Purple	Orange
sour milk	Blue	Pink	Yellow	Orange	Blue
sodium chloride	Colourless	Colourless	Red	Orange	Orange
Toothpaste	Blue	Pink	Yellow	Orange	Blue
calcium hydroxide	Blue	Pink	Yellow	Orange	Blue
Lemon juice	Red	Colourless	Red	Purple	Orange

From the table above, then the colour of indicators in different solution can be summarized.

Indicator	Colour of indicator in		
	Acid	Base/alkali	Neutral
Litmus paper/solution	Red	Blue	Colourless
Methyl orange	Red	Yellow	Red
Screened methyl orange	Purple	Orange	Orange
Phenolphthalein	Colourless	Purple	Colourless

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Bromothymol blue	Orange	Blue	Orange

The universal indicator

The universal indicator is a mixture of other indicator dyes. The indicator uses the pH scale. The pH scale shows the **strength** of bases and acids. The pH scale ranges from 1-14.These numbers are called **pH values**:

(i) pH values 1, 2, 3 shows a substance is **strongly acid**
(ii) pH values 4, 5, 6 shows a substance is a **weakly acid**
(iii) pH value 7 shows a substance is a **neutral**
(iv) pH values 8, 9, 10, 11 shows a substance is a **weak base/alkali.**
(v) pH values 12, 13, 14 shows a substance is a strong **base/alkali**

The pH values are determined from a pH chart. The pH chart is a multicolored paper with each colour corresponding to a pH value.i.e

(i) **red** correspond to pH 1, 2, 3 showing strongly acidic solutions.
(ii)**Orange/** yellow correspond to pH 4, 5, 6 showing weakly acidic solutions.
(iii)**Green** correspond to pH 7 showing neutral solutions.
(iv)**Blue** correspond to pH 8, 9, 10, 11 showing weakly alkaline solutions.
(v)**Purple/dark blue** correspond to pH 12,13,14 showing strong alkalis.

The universal indicator is available as:

(i) Universal indicator **paper/pH paper**
(ii) Universal indicator **solution.**

When determining the pH of a unknown solution using

(i) pH paper then the pH paper is dipped into the unknown solution. It changes/turn to a certain colour. The new colour is marched/compared to its corresponding one on the pH chart to get the pH value.

(ii) universal indicator **solution** then about 3 drops of the universal indicator **solution** is added into about 5cm3 of the unknown solution in a test tube. It changes/turn to a certain colour. The new colour is marched/compared to its corresponding one on the pH chart to get the pH value.

Experiment: To determine the pH value of some solutions

(a)Place 5cm3 of filtered wood ash, soap solution, ammonia solution, sodium hydroxide, hydrochloric acid, distilled water, sulphuric (VI) acid, sour milk, sodium chloride, toothpaste and calcium hydroxide into separate test tubes.

(b)Put about three drops of universal indicator solution or dip a portion of a piece of pH paper into each. Record the observations made in each case.

(c)Compare the colour in each solution with the colors on the pH chart provided. Determine the pH value of each solution.

Sample observations

Solution mixture	**Colour on the pH paper/adding universal indicator**	**pH value**	**Nature of solution**
wood ash	Blue	8	Weakly alkaline

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soap solution	Blue	8	Weakly alkaline
ammonia solution	**green**	8	Weakly alkaline
sodium hydroxide	**Purple**	14	Strongly alkaline
hydrochloric acid	red	1	Strongly acidic
distilled water	green	7	Neutral
sulphuric(VI)acid	**red**	1	Strongly acidic
sour milk	blue	9	Weakly alkaline
sodium chloride	green	7	Neutral
toothpaste	Blue	10	Weakly alkaline
calcium hydroxide	Blue	11	Weakly alkaline
Lemon juice	**Orange**	5	Weakly acidic

Note

1. All the mineral acids Hydrochloric, sulphuric (VI) and nitric (V) acids are strong acids
2. Two alkalis/soluble bases, sodium hydroxide and potassium hydroxide are strong bases/alkali. Ammonia solution is a weak base/alkali. All other bases are weakly alkaline.
3. Pure/deionized water is a neutral solution.
4. Common salt/sodium chloride is a neutral salt.
5. When an acid and an alkali/base are mixed, the final product has pH 7 and is neutral.

Properties of acids

(a)Physical properties of acids

1. Acids have a characteristic sour taste
2. Most acids are colourless liquids
3. Mineral acids are odorless. Organic acids have characteristic smell
4. All acids have pH less than 7
5. All acids turn blue litmus paper red, methyl orange red and phenolphthalein colourless.
6. All acids dissolve in water to form an acidic solution. Most do not dissolve in organic solvents like propanone, kerosene, tetrachloromethane, petrol.

(b)Chemical properties of acids

1. Reaction with metals

All acids react with reactive metals to form a salt and produce /evolve hydrogen gas.

Metal + Acid -> Salt + Hydrogen gas

Experiment: **reaction of metals with mineral acids**.

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(a)Place 5cm3 of dilute hydrochloric acid in a small test tube. Add 1cm length of polished magnesium ribbon. Stopper the test tube using a thump. Light a wooden splint. Place the burning splint on top of the stoppered test tube. Release the thump stopper. Record the observations made.
(b)Repeat the procedure in (a) above using Zinc granules, iron filings, copper turnings, aluminum foil in place of Magnesium ribbon
(c)Repeat the procedure in (a) then (b) using dilute sulphuric (VI) acid in place of dilute hydrochloric acid.

Sample observations

(i) effervescence/bubbles produced/fizzing in all cases except when using copper
(ii) Colourless gas produced in all cases except when using copper
(iii) Gas produced extinguishes a burning wooden splint with an explosion/pop sound.

Explanation

Some metals react with dilute acids, while others do not. Metals which react with acids produce bubbles of hydrogen gas. Hydrogen gas is a colourless gas that extinguishes a burning splint with a pop sound. This shows acids contain hydrogen gas.
This hydrogen is displaced/removed from the acids by some metals like Magnesium, Zinc, aluminium, iron and sodium.
Some other metals like copper, silver, gold; platinum and mercury are not reactive enough to displace/remove the hydrogen from dilute acids.

Chemical equations

1. Magnesium + Hydrochloric acid -> Magnesium chloride + Hydrogen
$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

2. Zinc + Hydrochloric acid -> Zinc chloride + Hydrogen
$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

3. Iron + Hydrochloric acid -> Iron (II) chloride + Hydrogen
$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$

4. Aluminium + Hydrochloric acid -> Aluminium chloride + Hydrogen
$2Al(s) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2(g)$

5. Magnesium + Sulphuric (VI) acid -> Magnesium sulphate (VI) + Hydrogen
$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

6. Zinc + Sulphuric (VI) acid -> Zinc sulphate (VI) + Hydrogen
$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

7. Iron + Sulphuric (VI) acid -> Iron (II) sulphate (VI) + Hydrogen
$Fe(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2(g)$

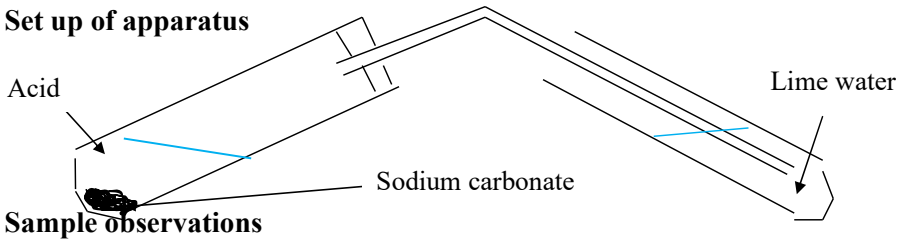
8. Aluminium + Sulphuric (VI) acid -> Aluminium sulphate (VI) + Hydrogen
$2Al(s) + 3H_2SO_4(aq \rightarrow Al_2(SO_4)_3(aq + 3H_2(g)$

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2. **Reaction of metal carbonates and hydrogen carbonates with mineral acids**.
All acids react with carbonates and hydrogen carbonates to form salt, water and produce /evolve carbon (IV) oxide gas.
Metal carbonate + Acid -> Salt + Water+ Carbon(IV)oxide gas
Metal hydrogen carbonate + Acid -> Salt + Water + Carbon (IV) oxide gas
Experiment: **reaction of metal carbonates and hydrogen carbonates with mineral acids.**
(a)Place 5cm3 of dilute hydrochloric acid in a small test tube. Add half spatula full of sodium carbonate. Stopper the test tube using a cork with delivery tube directed into lime water. Record the observations made. Test the gas also with burning splint.
(b)Repeat the procedure in (a) above using Zinc carbonate, Calcium carbonate, copper carbonate, sodium hydrogen carbonate, Potassium hydrogen carbonate in place of Sodium carbonate.
(c)Repeat the procedure in (a) then (b) using dilute sulphuric (VI) acid in place of dilute hydrochloric acid.

Set up of apparatus



Sample observations
(i) effervescence/bubbles produced/fizzing in all cases.
(ii) Colourless gas produced in all cases.
(iii) Gas produced forms a white precipitate with lime water.

Explanation
All metal carbonate/hydrogen carbonate reacts with dilute acids to produce bubbles of carbon (IV) oxide gas. Carbon (IV) oxide gas is a colourless gas that extinguishes a burning splint. When carbon (IV) oxide gas is bubbled in lime water, a white precipitate is formed.

Chemical equations

1. Sodium carbonate +Hydrochloric acid -> Sodium chloride + Carbon (IV) Oxide+ Water

$Na_2CO_3(s) + 2HCl\ (aq) \rightarrow 2NaCl\ (aq) + H_2O(g) + CO_2\ (g)$

2. Calcium carbonate +Hydrochloric acid -> Calcium chloride + Carbon (IV) Oxide+ Water

$CaCO_3(s) + 2HCl\ (aq) \rightarrow CaCl_2\ (aq) + H_2O(g) + CO_2\ (g)$

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3. Magnesium carbonate +Hydrochloric acid ->Magnesium chloride + Carbon (IV) Oxide+ Water

$MgCO_3(s) + 2HCl\ (aq) \rightarrow MgCl_2\ (aq) + H_2O\ (g) + CO_2\ (g)$

4. Copper carbonate +Hydrochloric acid ->Copper (II) chloride + Carbon (IV) Oxide+ Water

$CuCO_3(s) + 2HCl\ (aq) \rightarrow CuCl_2\ (aq) + H_2O\ (g) + CO_2\ (g)$

5. Copper carbonate +Sulphuric (VI) acid ->Copper (II) sulphate (VI) + Carbon (IV) Oxide+ Water

$CuCO_3(s) + H_2SO_4\ (aq) \rightarrow CuSO_4\ (aq) + H_2O\ (g) + CO_2\ (g)$

6. Zinc carbonate +Sulphuric (VI) acid ->Zinc sulphate (VI) + Carbon (IV) Oxide+ Water

$ZnCO_3(s) + H_2SO_4\ (aq) \rightarrow ZnSO_4\ (aq) + H_2O\ (g) + CO_2\ (g)$

7. Sodium hydrogen carbonate +Sulphuric (VI) acid ->Sodium sulphate (VI) + Carbon (IV) Oxide+ Water

$NaHCO_3(s) + H_2SO_4\ (aq) \rightarrow Na_2SO_4\ (aq) + H_2O\ (g) + CO_2\ (g)$

8. Potassium hydrogen carbonate +Sulphuric (VI) acid ->Potassium sulphate (VI) + Carbon (IV) Oxide+ Water

$KHCO_3(s) + H_2SO_4\ (aq) \rightarrow K_2SO_4\ (aq) + H_2O\ (g) + CO_2\ (g)$

9. Potassium hydrogen carbonate +Hydrochloric acid ->Potassium chloride + Carbon (IV) Oxide+ Water

$KHCO_3(s) + HCl\ (aq) \rightarrow KCl\ (aq) + H_2O\ (g) + CO_2\ (g)$

10. Sodium hydrogen carbonate +Hydrochloric acid ->Sodium chloride + Carbon (IV) Oxide+ Water

$NaHCO_3(s) + HCl\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (g) + CO_2\ (g)$

3. **Neutralization by bases/alkalis**

All acids react with bases to form a salt and water only. The reaction of an acid with metal oxides/hydroxides (bases) to salt and water only is called neutralization reaction.

Since no effervescence/bubbling/fizzing take place during neutralization:

(i) The reaction with alkalis requires a suitable indicator. The colour of the indicator changes when all the acid has reacted with the soluble solution of the alkali (metal oxides/ hydroxides).

(ii) Excess of the base is added to ensure all the acid reacts. The excess acid is then filtered off.

Experiment 1: reaction of alkali with mineral acids.

(i)Place about 5cm3 of dilute hydrochloric acid in a boiling tube. Add one drop of phenolphthalein indicator. Using a dropper/teat pipette, add dilute sodium hydroxide dropwise until there is a colour change.

(ii)Repeat the procedure with dilute sulphuric (VI) acid instead of hydrochloric acid.

(iii)Repeat the procedure with potassium hydroxide instead of sodium hydroxide.

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Sample observation:
Colour of phenolphthalein change from colourless to **pink** in all cases.
Explanation
Bases/alkalis neutralize acids. Acids and bases/alkalis are colourless. A suitable indicator like phenolphthalein change colour **to pink,** when all the acid has been neutralized by the bases/alkalis. Phenolphthalein change colour **from pink,** to colourless when all the bases/alkalis has been neutralized by the acid.
Chemical equation

Sodium oxide + Hydrochloric acid -> Sodium chloride + Water
$Na_2O(s)$ + HCl -> NaCl(aq) + $H_2O(l)$

Potassium oxide + Hydrochloric acid -> Potassium chloride + Water
$K_2O(s)$ + HCl -> KCl(aq) + $H_2O(l)$

Sodium hydroxide + Hydrochloric acid -> Sodium chloride + Water
NaOH(s) + HCl -> NaCl(aq) + $H_2O(l)$

Ammonia solution + Hydrochloric acid -> Ammonium chloride + Water
$NH_4OH(s)$ + HCl -> NH_4Cl (aq) + H_2O (l)

Potassium hydroxide + Hydrochloric acid -> Potassium chloride + Water
KOH(s) + HCl -> KCl(aq) + $H_2O(l)$

Sodium hydroxide + sulphuric (VI)acid -> Sodium sulphate(VI) + Water
2NaOH(s) + H_2SO_4 -> Na_2SO_4 (aq) + $2H_2O$ (l)

Potassium hydroxide + sulphuric (VI) acid -> Potassium sulphate (VI) + Water
2KOH(s) + H_2SO_4 -> K_2SO_4 (aq) + $2H_2O$ (l)

Ammonia solution + sulphuric (VI) acid -> Ammonium sulphate (VI) + Water
$2NH_4OH(s)$ + H_2SO_4 -> $(NH4)_2SO_4$ (aq) + $2H_2O$ (l)

Magnesium hydroxide + sulphuric (VI) acid -> Magnesium sulphate (VI) + Water
$Mg(OH)_2(s)$ + H_2SO_4 -> $MgSO_4$ (aq) + $2H_2O(l)$

Magnesium hydroxide + Hydrochloric acid -> Magnesium chloride + Water
$Mg(OH)_2(s)$ + HCl(aq) -> $MgCl_2$ (aq) + $2H_2O(l)$

CHEMISTRY FORM TWO NOTES

INTRODUCTION TO ELECTROLYSIS (ELECTROLYTIC CELL)
1. Electrolysis is defined simply as the **decomposition** of a **compound** by an electric current/**electricity**.
A compound that is decomposed by an electric current is called an electrolyte.
Some electrolytes are **weak** while others are **strong**.
2. Strong electrolytes are those that are fully ionized/dissociated into (many) ions.
Common strong electrolytes include:
(i) All **mineral** acids

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(ii) All strong **alkalis**/sodium hydroxide/potassium hydroxide.
(iii) All soluble **salts**

3. Weak electrolytes are those that are partially/partly ionized/dissociated into (few) ions.

Common weak electrolytes include:
(i) All **organic** acids
(ii) All **bases** except sodium hydroxide/potassium hydroxide.
(iii)**Water**

4. A compound that is **not** decomposed by an electric current is called non-electrolyte.

Non-electrolytes are those compounds /substances that exist as molecules and thus cannot ionize/dissociate into (any) ions.

Common non-electrolytes include:
(i) Most organic solvents (e.g. petrol/paraffin/benzene/methylbenzene/ethanol)
(ii) All hydrocarbons (alkanes /alkenes/alkynes)
(iii)Chemicals of life (e.g. proteins, carbohydrates, lipids, starch, sugar)

5. An electrolytes in **solid** state have **fused** /joined ions and therefore do **not** conduct electricity but the **ions** (cations and anions) are **free** and **mobile** in **molten** and **aqueous** (solution, dissolved in water) state.

6. During electrolysis, the free ions are attracted to the **electrodes**.

An electrode is a rod through which current enter and leave the electrolyte during electrolysis.

An electrode that does not influence/alter the products of electrolysis is called an **inert electrode.**

Common inert electrodes include:
(i)**Platinum**
(ii)**Carbon graphite**

Platinum is not usually used in a school laboratory because it is very **expensive**. Carbon graphite is **easily**/readily and **cheaply** available (from used dry cells).

7. The **positive** electrode is called **Anode**. The anode is the electrode through which **current enter** the electrolyte/**electrons leave** the electrolyte

8. The **negative** electrode is called **Cathode**. The cathode is the electrode through which **current leave** the electrolyte / **electrons enter** the electrolyte

9. During the electrolysis, free **anions** are attracted to the **anode** where they **lose** /**donate** electrons to form **neutral** atoms/molecules. i.e.

$M(l) \rightarrow M^{+}(l) + e$ (for cations from molten electrolytes)

$M(s) \rightarrow M^{+}(aq) + e$ (for cations from electrolytes in aqueous state / solution / dissolved in water)

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The neutral atoms /molecules form the **products** of electrolysis at the anode. This is called **discharge** at anode

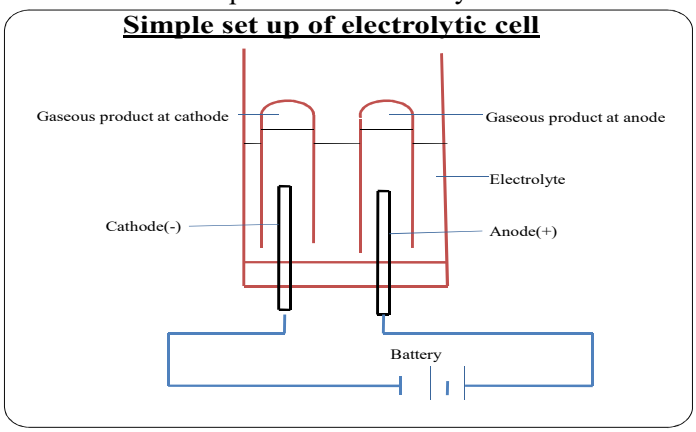
10. During electrolysis, free **cations** are attracted to the **cathode** where they **gain** /**accept/acquire** electrons to form **neutral** atoms/molecules.

X^+ (aq) + 2e -> X(s) (for cations from electrolytes in aqueous state / solution / dissolved in water)

$2X^+$ (l) + 2e -> X (l) (for cations from molten electrolytes)

The neutral atoms /molecules form the **products** of electrolysis at the cathode. This is called **discharge** at cathode.

11. The below set up shows an electrolytic cell.



12. For a compound /salt containing only two ion/binary salt the products of electrolysis in an electrolytic cell can be determined as in the below examples:

a) To determine the products of electrolysis of molten Lead (II) chloride

(i)Decomposition of electrolyte into free ions;

$PbCl_2$ (l) -> Pb^{2+} (l) + $2Cl^-$(l)

(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode (-);

Pb^{2+} (l) + 2e -> Pb (l)

(Cation / Pb^{2+} gains / accepts / acquires electrons to form free **atom)**

(iii)At the anode/positive electrode (+);

$2Cl^-$(l) -> Cl_2 (g) + 2e

(Anion / Cl^- donate/lose electrons to form free **atom** then a gas **molecule)**

(iv)Products of electrolysis therefore are;

I. At the cathode grey beads /solid lead metal.

II.At the anode pale green chlorine gas.

b) To determine the products of electrolysis of molten Zinc bromide

(i)Decomposition of electrolyte into free ions;

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$ZnBr_2$ (l) -> Zn^{2+} (l) + $2Br^-$(l)

(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode (-);

Zn^{2+} (l) + 2e -> Zn (l)

(Cation / Zn^{2+} gains / accepts / acquires electrons to form free **atom)**

(iii)At the anode/positive electrode (+);

$2Br^-$(l) -> Br_2 (g) + 2e

(Anion / Br^- donate/lose electrons to form free **atom** then a liquid **molecule** which change to **gas** on heating)

(iv)Products of electrolysis therefore are;

I. At the cathode grey beads /solid Zinc metal.

II.At the anode **red** bromine **liquid** / **red/brown** bromine **gas.**

c) To determine the products of electrolysis of molten sodium chloride

(i)Decomposition of electrolyte into free ions;

NaCl (l) -> Na^+(l) + Cl^-(l)

(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode (-);

$2Na^+$ (l) + 2e -> Na (l)

(Cation / Na^+ gains / accepts / acquires electrons to form free **atom)**

(iii)At the anode/positive electrode (+);

$2Cl^-$(l) -> Cl_2 (g) + 2e

(Anion / Cl^- donate/lose electrons to form free **atom** then a gas **molecule)**

(iv)Products of electrolysis therefore are;

I. At the cathode grey beads /solid sodium metal.

II.At the anode pale green chlorine gas.

d) To determine the products of electrolysis of molten Aluminum (III) oxide

(i)Decomposition of electrolyte into free ions;

Al_2O_3 (l) -> $2Al^{3+}$ (l) + $3O^{2-}$(l)

(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode (-);

$4Al^{3+}$ (l) + 12e -> 4Al (l)

(Cation / Al^{3+} gains / accepts / acquires electrons to form free **atom)**

(iii)At the anode/positive electrode (+);

$6O^{2-}$(l) -> $3O_2$ (g) + 12e

(Anion /$6O^{2-}$ donate/lose 12 electrons to form free **atom** then three gas **molecule)**

(iv)Products of electrolysis therefore are;

I. At the cathode grey beads /solid aluminum metal.

II.At the anode colourless gas that relights/rekindles glowing splint.

13. In industries electrolysis has the following uses/applications:

(a)Extraction of reactive metals from their ores

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Potassium, sodium, magnesium, and aluminum are extracted from their ores using electrolytic methods.

(b)Purifying copper after extraction from copper pyrites ores

Copper obtained from copper pyrites ores is not pure. After extraction, the copper is refined by electrolyzing copper (II) sulphate (VI) solution using the **impure** copper as **anode** and a thin strip of **pure** copper as **cathode**. Electrode ionization takes place there:

(i)At the cathode; **Cu^{2+} (aq) + 2e -> Cu(s)** (Pure copper deposits on the strip

(ii)At the anode; **Cu(s) ->Cu^{2+} (aq)** + **2e** (impure copper erodes/dissolves)

(c)Electroplating

The label EPNS (**E**lectro **P**lated **N**ickel **S**ilver) on some steel/metallic utensils mean they are plated/coated with silver and/or Nickel to **improve** their **appearance** (**add** their **aesthetic** value) and **prevent**/slow **corrosion**(**rusting** of iron). Electroplating is the process of coating a metal with another metal using an electric current. During electroplating, the **cathode** is made of the metal to be **coated**/impure.

Example:

During the electroplating of a spoon with silver

(i) The spoon/impure is placed as the cathode (negative terminal of battery)

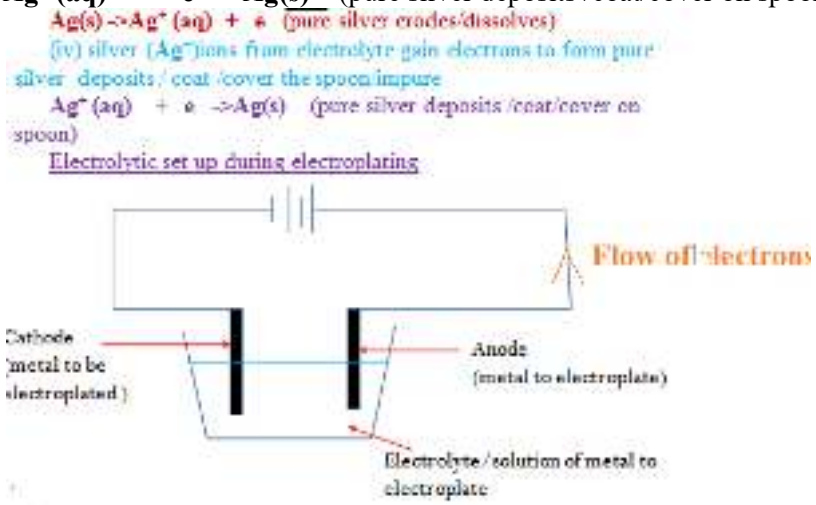
(ii) The pure silver is placed as the anode (positive terminal of battery)

(iii) The pure silver erodes/ionizes/dissociates to release electrons:

Ag(s) ->Ag^{+} (aq) + **e** (impure silver erodes/dissolves)

(iv) silver (**Ag^{+}**)ions from electrolyte gain electrons to form pure silver deposits / coat /cover the spoon/impure

Ag^{+} (aq) + e ->Ag(s) (pure silver deposits /coat/cover on spoon)



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CHEMICAL BONDING AND STRUCTURE

A. CHEMICAL BONDING

A chemical bond is formed when atoms of the same or different elements **share**, **gain**, **donate or delocalize** their **outer** energy level **electrons** to combine during **chemical reactions** in order to be **stable**.

Atoms have equal number of negatively charged electrons in the energy levels and positively charged protons in the nucleus.

Atoms are chemically stable if they have filled outer energy level. An energy level is full if it has duplet (2) or octet (8) state in outer energy level.

Noble gases have duplet /octet. All other atoms try to be like noble gases through chemical reactions and forming molecules.

Only electrons in the outer energy level take part in formation of a chemical bond.

There are three main types of chemical bonds formed by atoms:

(i) Covalent bond
(ii) ionic/electrovalent bond
(iii) Metallic bond

(i)COVALENT BOND

A covalent bond is formed when atoms of the same or different element share some or all the outer energy level electrons to combine during chemical reactions in order to attain duplet or octet.

A shared pair of electrons is attracted by the nucleus (protons) of the two atoms sharing.

Covalent bonds are mainly formed by non-metals to form molecules. A molecule is a group of atoms of the same or different elements held together by a covalent bond. The number of atoms making a molecule is called **atomicity**. Noble gases are **monatomic** because they are stable and thus do not bond with each other or other atoms. Most other gases are **diatomic**

The more the number of electrons shared, the stronger the covalent bond.

A pair of electrons that do not take part in the formation of a covalent bond is called a **lone pair of electrons.**

Mathematically, the number of electrons to be shared by an atom is equal to the number of electrons remaining for the atom to be stable/attain duplet/octet /have maximum electrons in outer energy level.

The following diagrams illustrate the formation of covalent bonds:

a) Hydrogen molecule is made up of two hydrogen atoms in the outer energy level each requiring one electron to have a stable duplet.

To show the formation of covalent bonding in the molecule then the following data/information is required;

Symbol of atom/element taking part in bonding	H	H
Number of protons/electrons	1	1
Electron configuration/structure	1:	1:

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Number of electron in outer energy level	1	1
Number of electrons remaining to be stable/shared	1	1
Number of electrons not shared (lone pairs)	0	0

Diagram method 1

Diagram method 2

H ••x

Note:
After bonding the following **intermolecular** forces exist:
(i) The attraction of the shared electrons by both nucleus /protons of the atoms
(ii) The repulsion of the nucleus of one atom on the other.
(iii) Balance of the attraction and repulsion is maintained inside/intermolecular/within the molecule as follows;

E_1

P_1 P_1

E_1

(iv)Protons (P_1) from nucleus of atom 1 **repel** protons (P_2) from nucleus of atom 2.
(v)Electron (E_1) in the energy levels of atom 1 **repel** electron (E_2) in the energy levels of atom 2.
(vi) Protons (P_1) from nucleus of atom 1 **attract** electron (E_2) in the energy levels of atom 2.
(vii) protons (P_2) from nucleus of atom 2 **attract** electron (E_2) in the energy levels of atom 2.
b) Fluorine, chlorine, bromine and iodine molecules are made up also of two atoms sharing the outer energy level electrons to have a stable octet.

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To show the formation of covalent bonding in the molecule then the following data/information is required:

(i) Fluorine

Symbol of atom/element taking part in bonding	F		F	
Number of protons/electrons	9		9	
Electron configuration/structure		2:7		2:7
Number of electron in outer energy level	7		7	
Number of electrons remaining to be stable/shared		1		1
Number of outer electrons not shared (3-lone pairs)	6		6	

Diagram method 1

Diagram method 2

(ii) Chlorine

Symbol of atom/element taking part in bonding	Cl		Cl	
Number of protons/electrons	17		17	
Electron configuration/structure		2:8:7		2:8:7
Number of electron in outer energy level	7		7	
Number of electrons remaining to be stable/shared		1		1
Number of outer electrons not shared (3-lone pairs)	6		6	

Diagram method 1

Diagram method 2

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(iii) Bromine

Symbol of atom/element taking part in bonding	Br	Br
Number of protons/electrons	35	35
Electron configuration/structure	2:8:18:7	2:8:18:7
Number of electron in outer energy level	7	7
Number of electrons remaining to be stable/shared	1	1
Number of outer electrons not shared (3-lone pairs)	6	6

Diagram method 1

Diagram method 2

(iv) Iodine

Symbol of atom/element taking part in bonding	I	I
Number of protons/electrons	53	53
Electron configuration/structure	2:8:18:18:7	
Number of electron in outer energy level	7	7
Number of electrons remaining to be stable/shared	1	1
Number of outer electrons not shared (3-lone pairs)	6	6

Diagram method 1

Diagram method 2

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c) Oxygen molecule is made up of two atoms sharing each two outer energy level electrons to have a stable octet as shown below;

Symbol of atom/element taking part in bonding	O		O	
Number of protons/electrons	8		8	
Electron configuration/structure		2:6		2:6
Number of electron in outer energy level	6		6	
Number of electrons remaining to be stable/shared		2		2
Number of outer electrons not shared (2-lone pairs)	4		4	

Diagram method 1

Diagram method 2

d) Nitrogen and phosphorus molecule is made up of two atoms sharing each three outer energy level electrons to have a stable octet as shown below;

(i) Nitrogen

Symbol of atom/element taking part in bonding	N		N	
Number of protons/electrons	7		7	
Electron configuration/structure		2:5		2:5
Number of electron in outer energy level	5		5	
Number of electrons remaining to be stable/shared		3		3
Number of outer electrons not shared (3-lone pairs)	2		2	

Diagram method 1

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Diagram method 2

(ii) Phosphorus

Symbol of atom/element taking part in bonding	P		P	
Number of protons/electrons	15		15	
Electron configuration/structure		2:8:5		2:8:5
Number of electron in outer energy level	5		5	
Number of electrons remaining to be stable/shared		3		3
Number of outer electrons not shared (3-lone pairs)	2		2	

Diagram method 1

Diagram method 2

e) Water molecule is made up of hydrogen and oxygen. Hydrogen requires sharing one electron with oxygen to be stable/attain duplet. Oxygen requires to share two electrons to be stable/attain octet. Two hydrogen atoms share with one oxygen atom for both to be stable as shown below;

Symbol of atoms/elements taking part in bonding		O		H
Number of protons/electrons	8		1	
Electron configuration/structure		2:6		1
Number of electron in outer energy level	6		1	
Number of electrons remaining to be stable/shared		2		1
Number of electrons not shared (2-Oxygen lone pairs)		4		0

Diagram method 1

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Diagram method 2

f) Ammonia molecule is made up of Hydrogen and Nitrogen. Hydrogen requires sharing one electron with Nitrogen to be stable/attain duplet. Nitrogen requires to share three electrons to be stable/attain octet. Three hydrogen atoms share with one nitrogen atom for both to be stable as shown below;

Symbol of atoms/elements taking part in bonding	N	H
Number of protons/electrons	7	1
Electron configuration/structure	2:5	1:
Number of electron in outer energy level	5	1
Number of electrons remaining to be stable/shared	3	1
Number of electrons not shared (1-Nitrogen lone pairs)	2	0

Diagram method 1

Diagram method 2

g) Carbon (IV) oxide molecule is made up of carbon and oxygen. Carbon requires to share four electrons with oxygen to be stable/attain octet. Oxygen requires to share two electrons to be stable/attain octet. Two oxygen atoms share with one carbon atom for both to be stable as shown below;

Symbol of atoms/elements taking part in bonding	O	C
Number of protons/electrons	8	6
Electron configuration/structure	2:6	2:4
Number of electron in outer energy level	6	4
Number of electrons remaining to be stable/shared	2	4
2-lone pairs from each Oxygen atom)	2	0

Diagram method 1

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Diagram method 2

h) Methane molecule is made up of hydrogen and carbon. Hydrogen requires sharing one electron with carbon to be stable/attain duplet. Carbon requires sharing four electrons to be stable/attain octet. Four hydrogen atoms share with one carbon atom for both to be stable as shown below;

Symbol of atoms/elements taking part in bonding	C	H
Number of protons/electrons	6	1
Electron configuration/structure	2:4	1
Number of electron in outer energy level	4	1
Number of electrons remaining to be stable/shared	4	1
Number of electrons not shared (No lone pairs)	0	0

Diagram method 1

Diagram method 2

i) Tetrachloromethane molecule is made up of chlorine and carbon. Chlorine requires sharing one electron with carbon to be stable/attain octet. Carbon requires sharing four electrons to be stable/attain octet. Four chlorine atoms share with one carbon atom for both to be stable as shown below;

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Symbol of atoms/elements taking part in bonding		C		Cl
Number of protons/electrons	6		17	
Electron configuration/structure		2:4		2:8:7
Number of electron in outer energy level	4		7	
Number of electrons remaining to be stable/shared		4		1
3-lone pairs from each Chlorine atom (24 electrons)		0		6

Diagram method 1

Diagram method 2

j) Ethane molecule is made up of six hydrogen and two carbon atoms. Hydrogen requires to share one electron with carbon to be stable/attain duplet. Carbon requires to share four electrons to be stable/attain octet. Three hydrogen atoms share with one carbon atom while another three hydrogen atoms share with a different carbon atom. The two carbon atoms bond by sharing a pair of the remaining electrons as shown below;

Symbol of atoms/elements taking part in bonding		C		H
Number of protons/electrons	6		1	
Electron configuration/structure		2:4		1
Number of electron in outer energy level	4		1	
Number of electrons remaining to be stable/shared		4		1
Number of electrons not shared (No lone pairs)	0		0	

Diagram method 1

Diagram method 2

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k) Ethene molecule is made up of four hydrogen and two carbon atoms. Hydrogen requires to share one electron with carbon to be stable/attain duplet. Carbon requires to share four electrons to be stable/attain octet. Two hydrogen atoms share with one carbon atom while another two hydrogen atoms share with a different carbon atom. The two carbon atoms bond by sharing two pairs of the remaining electrons as shown below;

Symbol of atoms/elements taking part in bonding	C		H
Number of protons/electrons	6		1
Electron configuration/structure	2:4		1
Number of electron in outer energy level	4		1
Number of electrons remaining to be stable/shared	4		1
Number of electrons not shared (No lone pairs)	0		0

Diagram method 1

Diagram method 2

l) Ethyne molecule is made up of two hydrogen and two carbon atoms. Hydrogen requires to share one electron with carbon to be stable/attain duplet. Carbon requires to share four electrons to be stable/attain octet. One hydrogen atoms share with one carbon atom while another hydrogen atoms share with a different carbon atom. The two carbon atoms bond by sharing three pairs of the remaining electrons as shown below;

Symbol of atoms/elements taking part in bonding	C		H
Number of protons/electrons	6		1
Electron configuration/structure	2:4		1
Number of electron in outer energy level	4		1
Number of electrons remaining to be stable/shared	4		1
Number of electrons not shared (No lone pairs)	0		0

Diagram method 1

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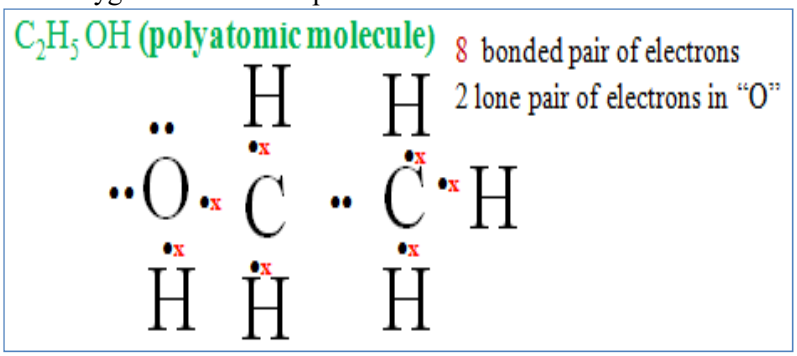
Diagram method 2

j) Ethanol molecule is made up of six hydrogen one Oxygen atom two carbon atoms.

Five Hydrogen atoms share their one electron each with carbon to be stable/attain duplet. One Hydrogen atoms share one electron with Oxygen for both to attain duplet/octet

Each Carbon uses four electrons to share with "O"and "H"attains octet/duplet.

NB: Oxygen has two lone pairs



j) Ethanoic molecule is made up of four hydrogen two Oxygen atom two carbon atoms.

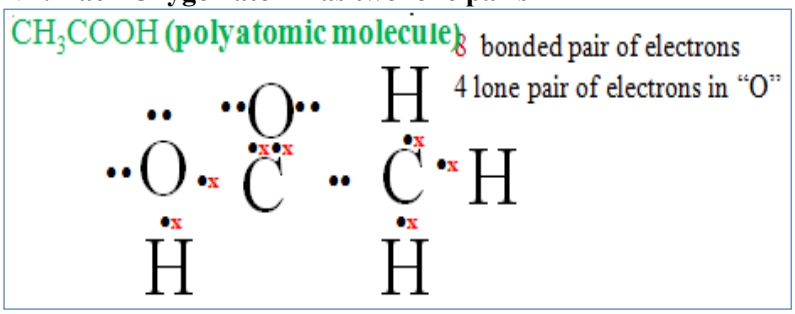
Three Hydrogen atoms share their one electron each with carbon to be stable/attain duplet. One Hydrogen atoms share one electron with Oxygen for both to attain duplet/octet

Each Carbon uses four electrons to share with "O"and "H"attains octet/duplet.

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NB: Each Oxygen atom has two lone pairs



By convention (as a rule), a

(i) **Single** covalent bond made up of two shared (a **pair**) electrons is represented by a dash (---)

(ii) **Double** covalent bond made up of four shared (**two pairs)** electrons is represented by a double dash (==)

(iii) **Triple** covalent bond made up of six shared (**three pairs)** electrons is represented by a triple dash (==)

The representation below show the molecules covered in (a) to (k) above:

a) Hydrogen molecule(H_2) H--H

b) Fluorine molecule(F_2) F--F

c) Chlorine molecule(Cl_2) Cl--Cl

d) Bromine molecule(Br_2) Br--Br

e) Iodine molecule(I_2) I--I

f) Oxygen molecule(O_2) O=O

g) Nitrogen molecule(N_2) N=N

h) Phosphorus molecule(P_2) P=P

i) Water molecule (H_2O) H--O--H

j Ammonia molecule(NH_3) H--N--H

H

k) Carbon (IV) oxide molecule (CO_2) O==C==O

H

l) Methane molecule (CH_4) H--C--H

H

Cl

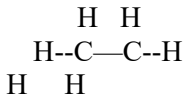
m) Tetrachloromethane molecule (CCl_4 Cl--C--Cl

Cl

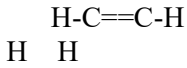
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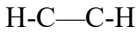
n) Ethane molecule (C_2H_6)



p) Ethene molecule (C_2H_4)



q) Ethyne molecule (C_2H_6)



Dative /coordinate bond

A dative/coordinate bond is a covalent bond formed when a lone pair of electrons is donated then shared to an electron-deficient species/ion/atom.

During dative/coordinate bonding, all the shared pair of electrons are donated by one of the combining/bonding species/ ion/atom.

Like covalent bonding, coordinate /dative bond is mainly formed by non-metals.

Illustration of coordinate /dative bond

a) Ammonium ion (NH_4^+)

The ammonium ion is made up of ammonia (NH_3) molecule and hydrogen (H^+) ion. (H^+) ion has no electrons. NH_3 is made up of covalent bonding from Nitrogen and Hydrogen. One lone pair of electrons is present in Nitrogen atom after the bonding. This lone pair is donated and shared with an electron-deficient H^+ ion

Diagram method 1

Diagram method 2

b) Phosphine ion (PH_4^+)

The Phosphine ion is made up of Phosphine (NH_3) molecule and hydrogen (H^+) ion. (H^+) ion has no electrons. PH_3 is made up of covalent bonding from Phosphorus and Hydrogen. One lone pair of electrons is present in Phosphorus atom. After the bonding this lone pair is donated and shared with the electron-deficient H^+ ion

Diagram method 1

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Diagram method 2

c) Hydroxonium (H_3O^+) ion

The hydroxonium ion is made up of water (H_2O) molecule and hydrogen (H^+) ion. (H^+) ion has no electrons. The H_2O molecule is made up of covalent bonding from Oxygen and Hydrogen. One lone pair of electrons out of the two present in Oxygen atom after the bonding is donated and shared with the electron-deficient H^+ ion

Diagram method 1

Diagram method 2

d) Carbon (II) oxide (CO)

Carbon (II) oxide is made up of carbon and Oxygen atoms sharing each two outer electron and not sharing each two electrons. Oxygen with an extra lone pair of electrons donates and share with the carbon atom for both to be stable.

Diagram method 1

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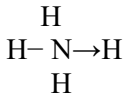
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Diagram method 2

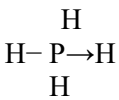
e) Aluminum (III) chloride ($AlCl_3/Al_2Cl_6$)
Aluminum (III) chloride is made up of aluminum and chlorine. One aluminum atom shares its outer electrons with three separate chlorine atoms. All chlorine atoms attain stable octet but aluminum does not. Another molecule of aluminum chloride shares its chlorine lone pair of electrons with the aluminum atom for both to be stable. This type of bond exists only in vapour phase after aluminum chloride sublimes.
Diagram method 1

Diagram method 2

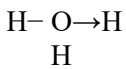
A dative/coordinate bond is by convention represented by an arrow (→) heading from the donor of the shared pair of electrons.
Below is the representation of molecules in the above examples;
a) Ammonium ion.



b) Phosphine ion



c) Hydroxonium ion

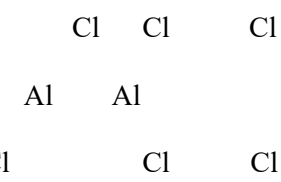


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d) Carbon (II) oxide O→C

d) Aluminum (III) chloride



(ii)**IONIC/ELECTROVALENT BOND**

An ionic/electrovalent bond is **extreme** of a covalent bond.
During ionic/electrovalent bonding there is complete **transfer** of valence electrons to one electronegative atom from an electropositive atom.
All metals are electropositive and easily/readily donate/lose their valence electrons.
All non-metals are electronegative and easily/readily gain/acquire extra electrons.
Ionic/electrovalent bonding therefore mainly involves transfer of electrons from metal/metallic radical to non-metallic radical.
When an electropositive atom donates /loses the valence electrons, it forms a positively charged **cation** to attain stable octet/duplet.
When an electronegative atom gains /acquires extra valence electrons, it forms a negatively charged **anion** to attain stable octet/duplet**.**
The **electrostatic attraction force** between the **stable** positively charged cation and the stable negatively charged anion with **opposite** charges constitute the ionic bond.
Like in covalent/dative/coordinate bonding, only the outer energy level electrons take part in the formation of ionic/electrovalent bond
Like in covalent/dative/coordinate bonding, the **more** electrons taking part / involved in the formation of ionic/electrovalent bond, the **stronger** the ionic /electrovalent bond. **Illustration of ionic /electrovalent bond**

a) Sodium chloride (NaCl)

Sodium chloride (NaCl) is formed when a sodium atom donates its outer valence electrons to chlorine atom for both to attain stable octet:

Symbol of atoms/elements taking part in bonding		Na	Cl
Number of protons/electrons	11		17
Electron configuration/structure		2:8:1	2:8:7
Number of electron in outer energy level	11		7
Number of electrons donated and gained to be stable	1		1
New electron configuration/structure		2:8:	2:8:
Symbol of cation/anion after bonding		Na^+	Cl^-

Diagram

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b) Magnesium chloride ($MgCl_2$)

Magnesium chloride ($MgCl_2$) is formed when a magnesium atom donate its two outer valence electrons to chlorine atoms. Two chlorine atoms are required to gain each one electron. All the ions (cations and anions) attain stable octet:

Symbol of atoms/elements taking part in bonding		Mg	Cl
Number of protons/electrons	11		17
Electron configuration/structure		2:8:2	2:8:7
Number of electron in outer energy level	2		7
Number of electrons donated and gained to be stable	2		1
New electron configuration/structure		2:8:	2:8:
Symbol of cation/anion after bonding		Mg^{2+}	Cl^-

Diagram

c) Lithium oxide (Li_2O)

Lithium oxide (Li_2O) is formed when a Lithium atom donate its outer valence electrons to Oxygen atom. Two Lithium atoms are required to donate/lose each one electron and attain stable duplet. Oxygen atom acquires the two electrons and attain stable octet:

Symbol of atoms/elements taking part in bonding		Li	O
Number of protons/electrons	3		8
Electron configuration/structure		2:1	2:6
Number of electron in outer energy level	1		6
Number of electrons donated and gained to be stable	1		2
New electron configuration/structure		2:	2:8:
Symbol of cation/anion after bonding		Li^+	O^{2-}

Diagram

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d) Aluminum (III) oxide (Al_2O_3)

Aluminum (III) oxide (Al_2O_3) is formed when a Aluminum atom donate its three outer valence electrons to Oxygen atom. Two Aluminum atoms are required to donate/lose each three electron and attain stable octet. Three Oxygen atoms gain/ acquire the six electrons and attain stable octet:

Symbol of atoms/elements taking part in bonding		Al		O
Number of protons/electrons	13		8	
Electron configuration/structure		2:8:3		2:6
Number of electron in outer energy level	3		6	
Number of electrons donated and gained to be stable	3		2	
New electron configuration/structure		2:8:		2:8:
Symbol of cation/anion after bonding		Al^{3+}		O^{2-}

Diagram

e) Calcium oxide (CaO)

Calcium oxide (CaO) is formed when a Calcium atom donate its two outer valence electrons to Oxygen atom. Both attain stable octet:

Symbol of atoms/elements taking part in bonding		Ca		O
Number of protons/electrons	20		8	
Electron configuration/structure		2:8:8:2		2:6
Number of electron in outer energy level	2		6	
Number of electrons donated and gained to be stable	2		2	
New electron configuration/structure		2:8:8:		2:8:
Symbol of cation/anion after bonding		Ca^{2+}		O^{2-}

Diagram

Some compounds can be formed from ionic/electrovalent, covalent and dative/coordinate bonding within their atoms/molecules:

a) Formation of ammonium chloride:

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Ammonium chloride is formed from the reaction of ammonia gas and hydrogen chloride gas. Both ammonia and hydrogen chloride gas are formed from covalent bonding. During the reaction of ammonia and hydrogen chloride gas to form Ammonium chloride;

-ammonia forms a dative/coordinate bond with electron deficient H^+ ion from Hydrogen chloride to form ammonium ion(NH_4^+)ion.

-the chloride ion Cl^- and ammonium ion(NH_4^+)ion bond through ionic / electrovalent bond from the electrostatic attraction between the opposite/unlike charges.

Diagram

b) Dissolution/dissolving of hydrogen chloride:

Hydrogen chloride is formed when hydrogen and chlorine atoms form a covalent bond. Water is formed when hydrogen and Oxygen atoms also form a covalent bond. When hydrogen chloride gas is dissolved in water;

-water molecules forms a dative/coordinate bond with electron deficient H^+ ion from Hydrogen chloride to form hydroxonium ion(H_3O^+)ion.

-the chloride ion Cl^- and hydroxonium ion(H_3O^+)ion bond through ionic / electrovalent bond from the electrostatic attraction between the opposite/unlike charges.

Diagram

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c) Dissolution/dissolving of ammonia gas:

Ammonia gas is formed when hydrogen and Nitrogen atoms form a covalent bond. Water is formed when hydrogen and Oxygen atoms also form a covalent bond. When Ammonia gas is dissolved in water;

-ammonia forms a dative/coordinate bond with electron deficient H^+ ion from a water molecule to form ammonium ion(NH_4^+)ion.

-the hydroxide ion OH^- and ammonium ion(NH_4^+)ion bond through ionic / electrovalent bond from the electrostatic attraction between the opposite/unlike charges.

Diagram

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(iii)**METALLIC BOND**

A metallic bond is formed when metallic atoms **delocalize** their outer electrons in order to be stable.

Metals delocalize their outer electrons to form positively charged cation .

The electrostatic attraction force between the metallic cation and the negatively charged electrons constitute the metallic bond.

The more delocalized electrons the stronger the metallic bond.

Illustration of ionic /electrovalent bond

a) Sodium (Na) is made of one valence electron. The electron is donated to form Na^+ ion. The electron is delocalized /free within many sodium ions.

Symbol of atoms/elements taking part in bonding	Na	Na	Na
Number of protons/electrons	11	11	11
Electron configuration/structure	2:8:1	2:8:1	2:8:1
Number of electron in outer energy level	1	1	1
Number of electrons delocalized/free within	1	1	1
New electron configuration/structure	2:8:	2:8:	2:8:
Symbol of cation after metallic bonding	Na^+	Na^+	Na^+

Diagram

(three)Metallic cations **attract**
(three) free/delocalized electrons

b) Aluminium (Al) is made of three valence electron. The three electrons are donated to form Al^{3+} ion. The electrons are delocalized /free within many aluminium ions.

Symbol of atoms/elements taking part in bonding	Al	Al	Al
Number of protons/electrons	13	13	13
Electron configuration/structure	2:8:3	2:8:3	2:8:3
Number of electron in outer energy level	3	3	3
Number of electrons delocalized/free within	3	3	3
New electron configuration/structure	2:8:	2:8:	2:8:
Symbol of cation after metallic bonding	Al^{3+}	Al^{3+}	Al^{3+}

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Diagram

(three)Metallic cations **attract**
(nine) free/delocalized electrons

c)Calcium (Ca) is made of two valence electron.The two electrons are donated to form Ca^{2+} ion.The electrons are delocalized /free within many Calcium ions.

Symbol of atoms/elements taking part in bonding	Ca	Ca	Ca
Number of protons/electrons	20	20	20
Electron configuration/structure	2:8:8:2	2:8:8:2	2:8:8:2
Number of electron in outer energy level	2	2	2
Number of electrons delocalized/free within	2	2	2
New electron configuration/structure	2:8:8:	2:8:8:	2:8:8:
Symbol of cation after metallic bonding	Ca^{2+}	Ca^{2+}	Ca^{2+}

Diagram

(three)Metallic cations **attract**
(six) free/delocalized electrons

d) Magnesium (Mg) is made of two valence electron. The two electrons are donated to form Mg^{2+}ion.The electrons are delocalized /free within many Magnesium ions.

Symbol of atoms/elements taking part in bonding	Mg		Mg
Number of protons/electrons	12	12	
Electron configuration/structure	2:8:2	2:8:2	
Number of electron in outer energy level	2	2	
Number of electrons delocalized/free within	2		2
New electron configuration/structure	2:8:		2:8:
Symbol of cation after metallic bonding	Mg^{2+}		Mg^{2+}

Diagram

(two)Metallic cations **attract**
(four) free/delocalized electrons

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e)Lithium (Li) is made of one valence electron. The electron is donated to form Li^+ ion. The electron is delocalized /free within many Lithium ions.ie;

Symbol of atoms/elements taking part in bonding		Li	Li	Li	Li
Number of protons/electrons	3	3	3	3	
Electron configuration/structure		2:1	2:1	2:1	2:1
Number of electron in outer energy level	1	1	1	1	
Number of electrons delocalized/free within	1	1	1	1	
New electron configuration/structure		2:1:	2:1:	2:1:	2:1:
Symbol of cation after metallic bonding	Li^+	Li^+	Li^+	Li^+	

Diagram

(four)Metallic cations **attract**
(four) free/delocalized electrons

B.CHEMICAL STRUCTURE

Chemical structure is the pattern/arrangement of atoms **after** they have bonded. There are two main types of chemical structures:

(i) Simple molecular structure
(ii) Giant structures

(i)Simple molecular structure

Simple molecular structure is the pattern formed after atoms of non-metals have **covalently** bonded to form simple molecules.

Molecules are made of atoms joined together by weak intermolecular forces called **Van-der-waals forces.**The Van-der-waals forces hold the **molecules** together while the covalent bonds hold the **atoms** in the molecule.

Illustration of simple molecular structure

a)Hydrogen molecule(H_2)

Hydrogen gas is made up of strong covalent bonds/**intra**molecular forces between each hydrogen atom making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces.

Illustration of simple molecular structure

a)Hydrogen molecule(H_2)

Hydrogen gas is made up of strong covalent bonds/**intra**molecular forces between each hydrogen atom making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces

b)Oxygen molecule(O_2)

Oxygen gas is made up of strong covalent bonds/**intra**molecular forces between each Oxygen atom making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces.

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Strong intramolecular forces/covalent bond

O=O:::: O=O:::: O=O:::: O=O
: : : : : : : : : : : :
O=O:::: O=O:::: O=O:::: O=O

weak intermolecular forces/van-der-waals forces

c) Iodine molecule(I_2)

Iodine solid crystals are made up of strong covalent bonds/**intra**molecular forces between each iodine atom making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces.

Strong intramolecular forces/covalent bond

I--- I:::: I --- I:::: I --- I:::: I --- I
: : : : : : : : : : : :
I --- I:::: I --- I:::: I --- I:::: I --- I

weak intermolecular forces/van-der-waals forces

d)Carbon(IV) oxide molecule(CO_2)

Carbon (IV) oxide gas molecule is made up of strong covalent bonds/**intra**molecular forces between each Carbon and oxygen atoms making the molecule. Each molecule is joined to another by weak Van-der-waals forces/ **inter**molecular forces.

Strong intramolecular forces/covalent bond

O=C=O:::: O=C=O:::: O=C=O
: : : : : :
O=C=O:::: O=C=O:::: O=C=O

weak intermolecular forces/van-der-waals forces

The following are the main characteristic properties of simple molecular structured compounds:

a)**State**

Most simple molecular substances are gases, liquid or liquids or solid that sublimes or has low boiling/melting points at room temperature (25^oC) and pressure (atmospheric pressure).

Examples of simple molecular substances include:

-all gases eg Hydrogen, oxygen, nitrogen, carbon (IV) oxide,
-Petroleum fractions eg Petrol, paraffin, diesel, wax,
-Solid non-metals eg Sulphur, Iodine
-Water

b) **Low melting/boiling points**

Melting is the process of **weakening** the intermolecular/ van-der-waal forces/ of attraction between the molecules that holding the substance/compound.

Note;

(i)Melting and boiling does not involve weakening/breaking the strong intramolecular force/covalent bonds holding the atoms **in** the molecule.

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(ii) Melting and boiling points increase with increase in atomic radius/size of the atoms **making** the molecule as the intermolecular forces / van-der-waal forces of attraction between the molecules increase. e.g.
Iodine has a higher melting/boiling point than chlorine because it has a higher /bigger atomic radius/size than chlorine, making the molecule to have stronger intermolecular force/ van-der-waal forces of attraction between the molecules than chlorine. Iodine is hence a solid and chlorine is a gas.

(c)Insoluble in water/soluble in organic solvents

Polar substances dissolve in polar solvents. Water is a polar solvent .Molecular substances do not thus dissolve in water because they are non-polar. They dissolve in non-polar solvents like methylbenzene, benzene, tetrachloromethane or propanone.

d) Poor conductors of heat and electricity

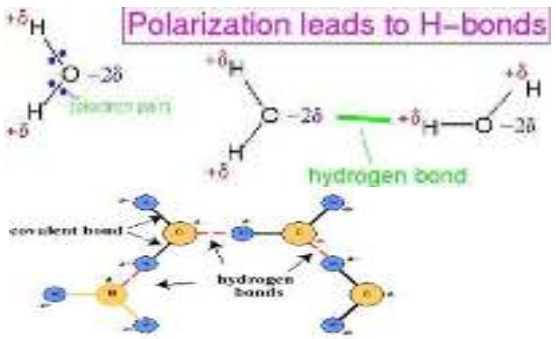
Substances with free mobile **ions** or free mobile/delocalized **electrons** conduct electricity. Molecular substances are poor conductors of heat/electricity because their molecules have no free mobile ions/electrons. This makes them very good **insulators**. **Hydrogen bonds**

A hydrogen bond is an intermolecular force of attraction in which a very electronegative atom attracts hydrogen atom of another molecule.
The most electronegative elements are Fluorine, Oxygen and Nitrogen .Molecular compounds made up of these elements usually have hydrogen bonds.
Hydrogen bonds are **stronger** than van-der-waals forces but **weaker** than covalent bonds. Molecular compounds with hydrogen bonds thus have higher melting/boiling points than those with van-der-waals forces.

Illustration of Hydrogen bonding

a)Water molecule

During formation of covalent bond, the oxygen atom attract/pull the shared electrons more to itself than Hydrogen creating partial negative charges(δ^-)in Oxygen and partial positive charges(δ^+)in Hydrogen.
Two molecules **attract** each other at the partial charges through Hydrogen bonding.



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The hydrogen bonding in water makes it;

(i)a **liquid** with higher boiling and melting point than simple molecular substances with higher molecular mass. e.g. Hydrogen sulphide as in the table below;

Influence of H-bond in water (H_2O) in comparison to H_2S

Substance	Water/ H_2O	Hydrogen sulphide/ H_2S
Relative molecular mass	18	34
Melting point(°C)	0	-85
Boiling point(°C)	100	-60

(ii)have higher **volume** in solid (ice) than liquid (water) and thus ice is less dense than water. Ice therefore floats above liquid water.

b)Ethanol molecule

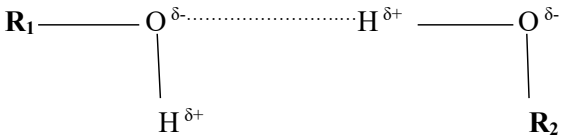
Like in water, the oxygen atom attracts/pulls the shared electrons in the covalent bond more to itself than Hydrogen.

This creates a partial negative charge ($^{\delta-}$) on oxygen and partial positive charge($^{\delta+}$) on hydrogen.

Two ethanol molecules attract each other at the partial charges through Hydrogen bonding forming a **dimmer**.

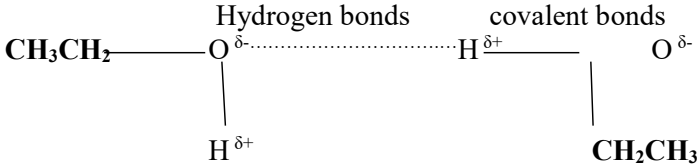
A dimmer is a molecule formed when two molecules join together as below:

Hydrogen bonds covalent bonds



R_1 and R_2 are extensions of the molecule.

For ethanol it is made up of CH_3CH_2 – to make the structure:



---

b)Ethanoic acid molecule

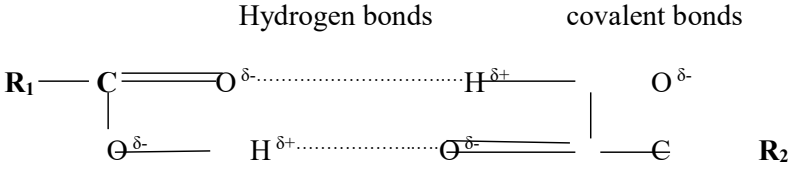
Like in water and ethanol above, the oxygen atom attracts/pulls the shared electrons in the covalent bond in ethanoic acid more to itself than Hydrogen.

This creates a partial negative charge ($^{\delta-}$)on oxygen and partial positive charge($^{\delta+}$) on hydrogen.

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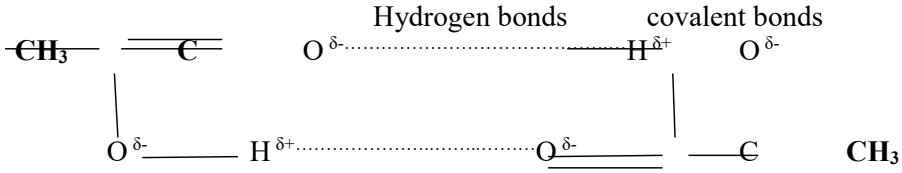
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Two ethanoic acid molecules attract each other at the partial charges through Hydrogen-bonding forming a **dimer.**



R_1 and $_2$ are extensions of the molecule.

For ethanoic acid the extension is made up of CH_3 – to make the structure;



Ethanoic acid like ethanol exists as a dimer.

Ethanoic acid has a **higher** melting/boiling point than ethanol .This is because ethanoic acid has **two/more** hydrogen bond than ethanol.

d) Proteins and sugars in living things also have multiple/**complex** hydrogen bonds in their structures.

(ii) Giant structure

This is the pattern formed after substances /atoms /ions bond to form a long chain network.

Giant structures therefore extend in all directions to form a pattern that continues **repeating** itself.

There are **three** main giant structures.

a) giant covalent/atomic structure
b)giant ionic structure
c)giant metallic structure

a) giant covalent/atomic structure

Giant covalent/atomic structure is the pattern formed after atoms have covalently bonded to form long chain pattern consisting of indefinite number of atoms covalently bonded together.

The strong covalent bonds hold all the atoms together to form a very well packed structure. Examples of substances with giant covalent/atomic structure include:

(i) carbon-**diamond**
(ii) carbon-**graphite**
(iii)silicon

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(iv) silicon(IV) oxide/sand

Carbon-graphite and carbon-diamond are **allotropes** of carbon.

Allotropy is the existence of an element in more than one stable physical **form** at the same temperature and pressure.

Allotropes are atoms of the same element existing in more than one stable physical form at the same temperature and pressure.

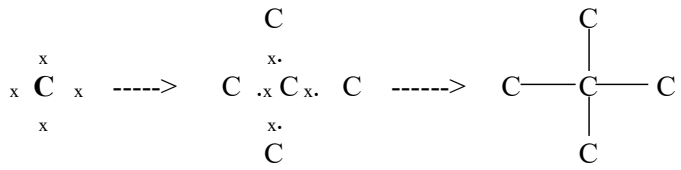
Other elements that exhibit/show allotropy include;

-Sulphur as **monoclinic** sulphur and **rhombic** sulphur

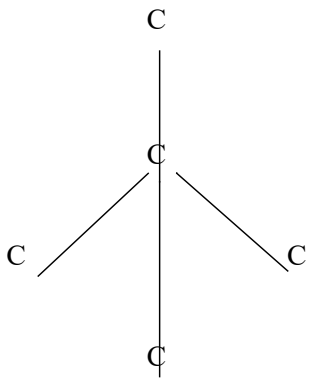
-Phosphorus as **white** phosphorus and **red** phosphorus

The structure of carbon-diamond

Carbon has four valence electrons. The four valence electrons are used to form covalent bonds. During the formation of diamond, one carbon atom covalently bond with four other carbon atoms.

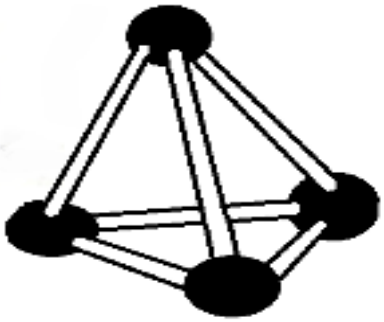


After the bonding, the atoms rearrange to form a regular **tetrahedral** in which **one** carbon is in the **centre** while **four** are at the **apex**/corners.

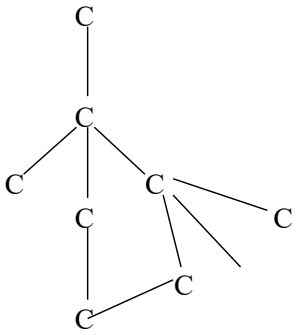


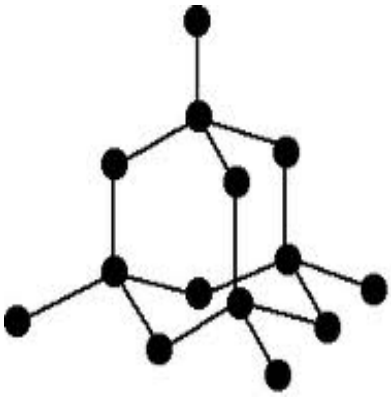
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This pattern repeats itself to form a long chain number of atoms covalently bonded together indefinitely. The pattern is therefore called **giant tetrahedral structure.** It extends in all directions where one atom of carbon is always a centre of four others at the apex/corner of a regular tetrahedral.





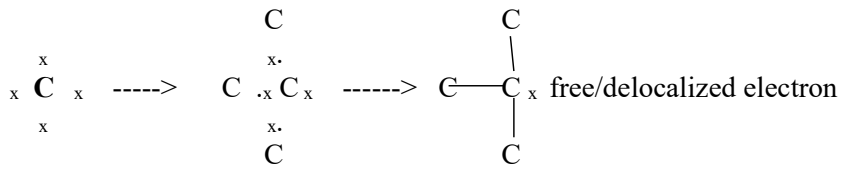
The giant tetrahedral structure of carbon-diamond is very well/closely packed and joined/bonded together by strong covalent bond.

This makes carbon-diamond to have the following properties:

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a) High melting/boiling point.
The giant tetrahedral structure is very well packed and joined together by strong covalent bonds.
This requires a lot of energy/heat to weaken for the element to melt and break for the element to boil.
b) High density.
Carbon diamond is the **hardest** known **natural** substance.
This is because the giant tetrahedral structure is a very well packed pattern/structure and joined together by strong covalent bonds.
This makes Carbon diamond be used to make **drill** for drilling boreholes/oil wells.
The giant tetrahedral structure of carbon diamond is a very closely packed pattern /structure such that heat transfer by **conduction** is possible. This makes carbon diamond a **good thermal** conductor.
c) Poor conductor of electricity.
Carbon-diamond has **no free/delocalized electrons** within its structure and thus do not conduct electricity.
d) Insoluble in water.
Carbon-diamond is insoluble in water because it is non-polar and do not bond with water molecules.
e) Is abrasive/Rough.
The edges of the closely well packed pattern/structure of Carbon-diamond make its surface rough/abrasive and thus able to smoothen /cut metals and glass.
f) Have characteristic luster.
Carbon-diamond has a high **optical dispersion** and thus able to disperse light to different colours .This makes Carbon-diamond one of the most popular **gemstone** for making **jewellery**.
The structure of carbon-graphite
During the formation of graphite, one carbon atom covalently bond with **three** other carbon atoms leaving **one** free/delocalized electron.

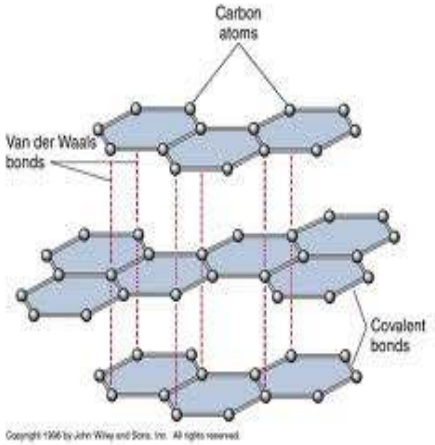


After the bonding, the atoms rearrange and join together to form a regular **hexagon** in which **six** carbon atoms are at the **apex**/corners**.**
The regular hexagon is joined to another in layers on the **same surface** by van-der-waals forces.

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Each layer extends to form a plane in all directions.
The fourth valence electron that does not form covalent bonding is free/mobile /delocalized within the layers.
This structure/pattern is called **giant hexagonal planar** structure.



The giant hexagonal planar structure of carbon-graphite is closely packed and joined/bonded together by strong covalent bonds. This makes carbon-graphite to have the following properties:

a) High melting/boiling point.

The giant hexagonal planar structure of carbon-graphite is well packed and joined together by strong covalent bonds.
This requires a lot of energy/heat to **weaken** for the element to melt and **break** for the element to boil.

b) Good conductor of electricity.

Carbon-graphite has **free/delocalized 4th valence electrons** within its structure and thus conducts electricity.

c) Insoluble in water.

Carbon-graphite is insoluble in water because it is **non-polar** and do not bond with water molecules.

d) Soft.

Layers of giant hexagonal planar structure of carbon graphite are held together by van-der-waals forces.
The van-der-waals forces easily break when pressed and reform back on releasing/reducing pressure/force thus making graphite soft.

e) Smooth and slippery.

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When pressed at an **angle** the van-der-waals forces easily break and slide over each other making graphite soft and slippery.
It is thus used as a **dry** lubricant instead of oil.

f)Some uses of carbon-graphite.

1. As a dry lubricant- carbon graphite is smooth and slippery and thus better lubricant than oil.Oil **heat up** when reducing friction.
2. Making Lead-pencils- When pressed at an **angle** on paper the van-der-waals forces easily break and slide smoothly over contrasting background producing its characteristic black background.
3. As moderator in nuclear reactors to reduce the rate of decay/disintegration of radioactive nuclides/atoms/isotopes.
4. As electrode in dry/wet cells/battery- carbon graphite is inert and good conductor of electricity. Current is thus able to move from one electrode/terminal to the other in dry and wet cells/batteries. Carbon graphite is also very **cheap**.

b) giant ionic structure

Giant ionic structure is the pattern formed after **ions** have bonded through ionic/electrovalent bonding to form a long chain consisting of **indefinite** number of ions.
The strong ionic/electrovalent bond holds all the **cations** and **anions** together to form a very well packed structure.
Substances with giant ionic structure are mainly crystals of salts e.g. sodium chloride, Magnesium chloride, Sodium iodide, Potassium chloride, copper (II) sulphate(VI).

The structure of sodium chloride

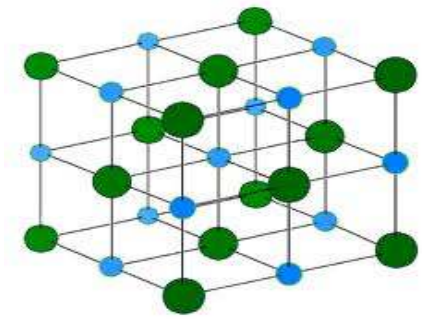
Sodium chloride is made up of sodium (Na^+) and chloride (Cl^-)ions.
Sodium (Na^+) ion is formed when a sodium atom donate /loose/donate an electron.
Chloride (Cl^-) ion is formed when a chlorine atom gain /acquire an extra electron from sodium atom.
Many Na^+ and Cl^- ions then rearrange such that **one** Na^+ ion is surrounded by **six** Cl^- ions and one Cl^- ion is surrounded by six Na^+ ions.
The pattern formed is a **giant cubic structure** where Cl^- ion is sand witched between Na^+ ions and the same to Na^+ ions.
This pattern forms a **crystal**.
A crystal is a solid form of a substance in which particles are arranged in a definite pattern regularly repeated in three dimensions.

The structure of sodium chloride

The giant cubic structure/crystal of sodium chloride is as below;

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The giant cubic structure/crystal of sodium chloride is very well packed and joined by strong ionic/electrovalent bonds. This makes sodium chloride and many ionic compounds to have the following properties:

a) Have high melting /boiling points.

The giant cubic lattice structure of sodium chloride is very **closely** packed into a crystal that requires a lot of energy/heat to **weaken** and melt/boil. This applies to all crystalline ionic compounds.

b) Are good conductors of electricity in molten and aqueous state but **poor conductor of electricity in solid.**

Ionic compounds have **fused** ions in solid crystalline state.

On heating and dissolving in water, the crystal is broken into free mobile ions (Na^+ and Cl^- ions).

The free mobile **ions** are responsible for **conduct**ing electricity in ionic compounds in molten and aqueous states.

c)Soluble in water

Ionic compounds are **polar** and dissolve in **polar** water molecules.

On dissolving, the crystal breaks to free the fused ions which are then surrounded by water molecules.

b) giant metallic structure

This is the pattern formed after metallic atoms have bonded through metallic bond.

The pattern formed is one where the metallic cations rearrange to form a cubic structure.

The cubic structure is bound together by the free delocalized electrons that move freely within.

The **more** delocalized electrons, the **stronger** the metallic bond.

The structure of sodium and aluminium.

Sodium has one valence electrons.

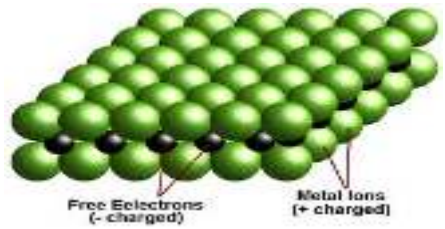
Aluminium has three valence electrons.

After delocalizing the valence electrons ,the metal cations (Na^+ and Al^{3+}) rearrange to the apex /corners of a regular cube that extend in all directions.

The delocalized electrons remain free and mobile as shown below:

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The giant cubic structure makes metals to have the following properties:

a) Have high melting/boiling point

The giant cubic structure is very well packed and joined/bonded together by the free delocalized electrons.

The more delocalized electrons the higher the melting/boiling point.

The larger/bigger the metallic cation, the weaker the packing of the cations and thus the lower the melting/boiling point. e.g.

(i) Sodium and potassium have both one valence delocalized electron.

Atomic radius of potassium is larger/bigger than that of sodium and hence less well packed in its metallic structure.

Sodium has therefore a higher melting/boiling point than potassium.

(ii) Sodium has one delocalized electron.

Aluminium has three delocalized electrons.

Atomic radius of sodium is larger/bigger than that of aluminium and hence less well packed in its metallic structure.

Aluminium has therefore a higher melting/boiling point than sodium because of the smaller well packed metallic (Al^{3+}) ions and bonded/joined by more/three delocalized electrons.

The table below shows the **comparative** melting/boiling points of some metals:

Metal	Electronic structure	Atomic radius(nM)	Melting point(°C)	Boiling point(°C)
Sodium	2:8:1	0.155	98	890
Potassium	2:8:8:1	0.203	64	774
Magnesium	2:8:2	0.136	651	1110
Aluminium	2:8:3	0.125	1083	2382

b) Good electrical and thermal conductor/electricity.

All metals are good conductors of heat and electricity including Mercury which is a **liquid**.

The mobile delocalized electrons are free within the giant metallic structure to move from one end to the other transmitting heat/electric current.

The more delocalized electrons the better the thermal/electrical conductivity.

High temperatures/heating lowers the thermal/electrical conductivity of metals because the delocalized electrons vibrate and move randomly hindering transfer of heat

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From the table above:

Compare the electrical conductivity of:

(i)**Magnesium and sodium**

Magnesium is a better conductor than sodium.

Magnesium has more/two delocalized electrons than sodium. The more delocalized electrons the better the electrical conductor.

(ii)**Potassium and sodium**

Potassium is a better conductor than sodium.

Potassium has bigger/larger atomic radius than sodium. The delocalized electrons are less attracted to the nucleus of the atom and thus more free /mobile and thus better the electrical conductor.

c) Insoluble in water

All metals are insoluble in water because they are non polar and thus do not bond with water.

Metals higher in the reactivity/electrochemical series like; Potassium, sodium, Lithium and Calcium **reacts** with **cold** water producing hydrogen gas and forming an **alkaline** solution of their **hydroxides**.ie

$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

$2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$

$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

Heavy metal like Magnesium, Aluminium, Iron, Zinc and Lead **react** with **steam/water vapour** to produce hydrogen gas and form the corresponding **oxide**.

$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$

$Fe(s) + H_2O(g) \rightarrow FeO(s) + H_2(g)$

$Zn(s) + H_2O(g) \rightarrow ZnO(s) + H_2(g)$

$Pb(s) + H_2O(g) \rightarrow PbO(s) + H_2(g)$

$2Al(s) + 3H_2O(g) \rightarrow Al_2O_3(s) + 3H_2(g)$

Metals **lower** in the reactivity/electrochemical series than hydrogen like; copper, Mercury, Gold Silver and Platinum **do not** react with water/vapour.

d) Shiny metallic-lustre

All metals have a shiny **grey** metallic luster except copper which is **brown.**

When exposed to sunlight, the delocalized electrons gain energy, they vibrate on the metal surface scattering light to appear shiny.

With time, most metals corrode and are covered by a layer of the metal oxide. The delocalized electrons are unable to gain and scatter light and the metal surface tarnishes/become dull.

e) Ductile and malleable

All metals are malleable (can be made into thin **sheet**) and ductile (can be made into **wire**.

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When beaten/hit/pressed **lengthwise** the metallic cations extend and is bound /bonded by the free/mobile electrons to form a sheet.
When beaten/hit/pressed **lengthwise and bredthwise** the metallic cations extend and is bound /bonded by the free/mobile electrons to form a wire/thin strip.

f) Have high tensile strength

Metals are not brittle. The free delocalized electrons bind the metal together when it is bent /coiled at any angle.
The metal thus withstand stress/coiling

g) Form alloys

An alloy is a uniform mixture of two or more metals.
Some metals have spaces between their metallic cations which can be occupied by another metal cation with smaller atomic radius.
Common alloys include:

Brass(Zinc and Copper alloy)
Bronze(Copper and Tin alloy)
German silver

Summary of Bonding and structure

	Simple molecular structure	**Giant covalent /atomic structure**	**Giant ionic structure**	**Giant metallic structure**
(i)Examples	I_2,S_8,HCl,O_2,CH_4	Graphite,diamond Si,SiO_2	NaCl, KCl, CaO,$CuSO_4$	Na,Fe,Cr,Hg,K
Constituent particles making structure	molecules	Atoms (of non-metals)	Ions (cation and anions)	Atoms (of metals)
Type of substance	Non-metal element/non-metal molecule/non-metal compound(electronegative elements)	Group IV non-metals and some of their oxides	Metal-non metal compounds(compounds of electropositive and electronegative compounds)	Metallic elements (with low electonegativity and high electropositivity)
	-Strong covalent	Atoms are	Electrostatic	

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Bonding in solid state	bonds hold atoms together within separate molecules (intramolecular forces) -Weak van-der-waals forces hold separate molecules together (intermolecular forces)	linked through the whole structure by very strong covalent bonds.	attraction of cations and anions link the whole structure through strong ionic bond.	Electrostatic attraction of outer mobile electrons for positive nuclei binds atoms together though metallic bond
Properties (i) Volatility	-Highly volatile with low melting/boiling point -Low latent heat of fusion/vaporization	-Non volatile with very high melting/boiling points -Low latent heat of fusion / vaporization	-Non volatile with very high melting/boiling points -Low latent heat of fusion / vaporization	-Non volatile with very high melting/boiling points -Low latent heat of fusion / vaporization
(ii) State at room temperature /pressure	Usually gases,volatile liquids or solids that sublimes	solids	solids	Solids except Mercury(liquid)
(iii) Hardness	Soft and brittle(low tensile strength)	Hard and brittle(low tensile strength)	Hard and brittle(low tensile strength)	Hard, malleable, ductile and have high tensile strength

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(iv) Thermal /electrical conductivity	Poor thermal and electrical conductor when solid ,liquid or aqueous solutions but some dissolve and react to form electrolytes e.g. Hydrogen chloride and ammonia gases.	Poor thermal and electrical conductor when solid ,liquid or aqueous solutions but -Carbon-graphite is a good electrical conductor while -Carbon-diamond is a good thermal conductor.	Poor thermal and electrical conductor when solid. Good thermal and electrical conductor in liquid/molten and aqueous states when the ions are not fused	Good thermal and electrical conductor in solid and liquid/molten states due to the free mobile /delocalized electrons
(v) Solubility	Insoluble in polar solvents e.g. Water Soluble in non-polar solvents e.g. tetrachloromethane , benzene, methylbenzene	Insoluble in all solvents	Soluble in polar solvents e.g. Water Insoluble in non-polar solvents e.g. tetrachloromethane, benzene, methylbenzene	Insoluble in polar/non-polar colvents. -Some react with polar solvents -Some metal dissolve in other metals to form alloys e.g. Brass is formed when Zinc dissolve in copper.

C. PERIODICITY OF BONDING AND STRUCTURE

The periodic table does not classify elements as metals and non-metals. The table arranges them in terms of atomic numbers.

However, based on structure and bonding of the elements in the periodic table;

(i)-the top right hand corner of about twenty elements is non-metals

(ii)-left of each non-metal is an element which shows characteristics of both metal and non-metal.

These elements are called **semi-metals/metalloids**. They include Boron, silicon, Germanium, Arsenic, and Terullium

(iii)-all other elements in the periodic table are metal.

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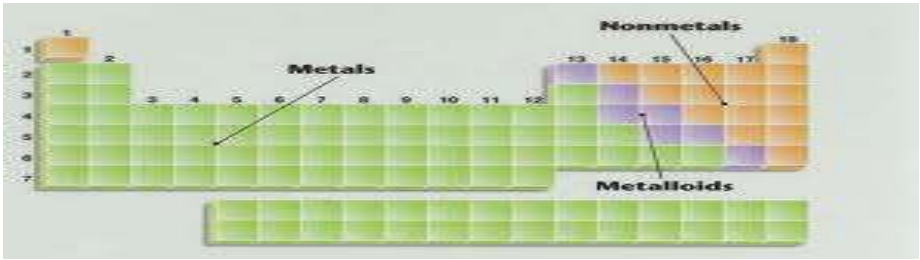
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(iv)-Hydrogen is a non-metal with metallic characteristic/property of donating/losing outer electron to form cation/H^+ ion.
(v) –bromine is the only known natural liquid non-metal element at room temperature and pressure.
(vi) –mercury is only known natural liquid metal element at room temperature and pressure.
(vii) Carbon-graphite is a semi metals/metalloids. Carbon-diamond is a pure non-metal yet both are allotropes of carbon (same element)

a) Sketch of the periodic table showing metals ,metalloid and non-metals

Metals **Metalloids** *Non-metals*

H								*He*
Li	Be		**B**	**C**	*N*	*O*	*F*	*Ne*
Na	Mg		Al	**Si**	*P*	*S*	*Cl*	*Ar*
K	Ca	Transition metals	Ga	Ge	**As**	*Se*	*Br*	*Kr*
Rb	Sr		In	Sn	Sb	**Te**	*I*	*Xe*
Cs	Ba		Tl	Pb	Bi	Po	**At**	*Rn*
Fr	Ra							



b)Periodicity in the physical properties of elements across period 2 and 3

Study table I and II below:

Table I(period 2)

Property	**Li**	**Be**	**B**	**C**	**N**	**O**	**F**	**Ne**
Melting point(°C)	180	1280	2030	3700 (graphite) 3550 (diamond)	-210	-219	-220	-250
Boiling	1330	2480	3930	Graphite	-200	-180	-190	-245

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point(°C)				sublimes 4830 (diamond)				
Density at room temperature (gcm^{-3})	0.50	1.85	2.55	2.25 (graphite) 3.53 (diamond)	0.81	0.14	0.11	0.021
Type of element	Metal	Metal	Metal	Metalloid	Non-metal	Non-metal	Non-metal	Non-metal
Chemical structure	Giant metallic	Giant metallic	Giant atomic/ covalent	Giant atomic/ covalent	Simple molecula or molecule/ N_2	Simple molecula or molecules /O_2	Simple molecula or molecule/F_2	Simple molecula or molecule/Ne
State at room temperature	Solid	Solid	Solid	Solid	gas	gas	gas	gas
Electron structure	2:1	2:2	2:3	2:4	2:5	2:6	2:7	2:8
Valency	1	2	3	4	3	2	1	-
Formular of ion	Li^+	Be^{2+}	B^{3+}	-	N^{3-}	O^{2-}	F^-	-

Table II (period 3)

Property	**Na**	**Mg**	**Al**	**Si**	**P(white)**	**S(Rhombic)**	**Cl**	**Ar**

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Melting point(°C)	98	650	660	1410	44	114	-101	-189
Boiling point(°C)	890	1120	2450	2680	280	445	-34	-186
Density at room temperature (gcm^{-3})	0.97	1.74	2.70	2.33 (graphite) 3.53 (diamond)	1.82	2.07	0.157	0.011
Type of element	Metal	Metal	Metal	Metalloid	Non-metal	Non-metal	Non-metal	Non-metal
Chemical structure	Giant metallic	Giant metallic	Giant metallic	Giant atomic/ covalent	Simple molecula or molecule/ P_4	Simple molecula or molecules /S_8	Simple molecula or molecule/Cl_2	Simple molecula or molecule/Ar
State at room temperature	Solid	Solid	Solid	Solid	Solid	Solid	gas	gas
Electron structure	2:8:1	2:8:2	2:8:3	2:8:4	2:8:5	2:8:6	2:8:7	2:8:8
Valency	1	2	3	4	3	2	1	-
Formular of ion	Na^+	Mg^{2+}	Al^{3+}	-	P^{3-}	S^{2-}	Cl^-	-

From table I and II above:

1. Explain the trend in atomic radius along /across a period in the periodic table

Observation

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Atomic radius of elements in the same period decrease successively across/along a period from left to right.

Explanation

Across/along the period from left to right there is an increase in nuclear charge from additional number of protons and still additional number of electrons entering the same energy level.

Increase in nuclear charge increases the effective nuclear attraction on the outer energy level pulling it closer to the nucleus successively across the period .e.g.

(i)From the table 1and 2 above, atomic radius of Sodium (0.157nM) is higher than that of Magnesium(0.137nM). This is because Magnesium has more effective nuclear attraction on the outer energy level than Sodium hence pulls outer energy level more nearer to its nucleus than sodium.

(ii)The rate of decrease in the atomic radius become smaller as the atom become heavier e.g. Atomic radius of Magnesium from sodium falls by(0.157nM-0.137nM) =**0.02**

Atomic radius of Chlorine from sulphur falls by(0.104nM- 0.099nM) =**0.005**

This is because gaining/adding one more proton to 11 already present cause greater proportional change in nuclear attraction power to magnesium than gaining/adding one more proton to 16 already present in sulphur to chlorine.

(iii)Period 3 elements have more energy levels than Period 2 elements. They have therefore bigger/larger atomic radius/size than corresponding period 2 elements in the same group.

2. Explain the trend in ionic radius along/across a period in the periodic table

Observation

Ionic radius of elements in the same period decrease successively across/along a period from left to right for the first three elements then increase drastically then slowly successively decrease.

Explanation

Across/along the period from left to right elements change form electron donors/losers (**reducing** agents) to electron acceptors (**oxidizing** agents).

(i)An atom form stable ion by either gaining/acquiring/ accepting extra electron or donating/losing outer electrons.

(ii)Metals form stable ions by donating/losing **all** the outer energy level electrons and thus **also** the outer energy level .i.e.

-Sodium ion has one less energy level than sodium atom. The ion is formed by sodium atom donating/losing (all) the outer energy level electron and thus also the outer energy level making the ion to have smaller ionic radius than atom.

(iii)Ionic radius therefore decrease across/along the period from Lithium to Boron in period 2 and from Sodium to Aluminium in period 3.This is because the number of electrons donated/lost causes increased effective nuclear attraction on remaining electrons /energy levels.

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(iv) Non-metals form stable ion by gaining/acquiring/accepting extra electron in the outer energy level. The extra electron/s increases the repulsion among electrons and reduces the effective nuclear attraction on outer energy level. The outer energy level therefore expand/enlarge/increase in order to accommodate the extra repelled electrons .The more electrons gained/accepted/acquired the more repulsion and the more expansion to accommodate them and hence bigger/larger atomic radius. e.g.
-Nitrogen ion has three electrons more than Nitrogen atom. The outer energy level expand/enlarge/increase to accommodate the extra repelled electrons. Nitrogen atom thus has smaller atomic radius than the ionic radius of nitrogen ion.
(v) Ionic radius decrease from group IV onwards from left to right. This because the number of electrons gained to form ion decrease across/along the period from left to right. e.g. Nitrogen ion has bigger/larger ionic radius than Oxygen.

3. Explain the trend in melting and boiling point of elements in a period in the periodic table.

Observation

The melting and boiling point of elements rise up to the elements in Group IV(Carbon/Silicon) along/across the period then continuously falls.

Explanation

Melting/boiling points depend on the packing of the structure making the element and the strength of the bond holding the atoms/molecules together.
Across/along the period (2 and 3) the structure changes from giant metallic, giant atomic/covalent to simple molecular.
(i)For metals, the number of delocalized electrons increases across/along the period and hence stronger metallic bond/structure thus requiring a lot of heat/energy to weaken.
The strength of a metallic bond also depends on the atomic radius/size. The melting /boiling point decrease as the atomic radius/size of metals increase due to decreased packing of larger atoms. e.g.
-The melting /boiling point of Lithium is lower than that of Beryllium because Beryllium has two/more delocalized electrons and hence stronger metallic structure/bond.
- The melting /boiling point of Lithium is higher than that of Sodium because the atomic radius/size Lithium is smaller and hence better packed and hence forms stronger metallic structure/bond.
(ii)Carbon-graphite/carbon-diamond in period 2 and Silicon in period 3 form very well packed giant atomic/covalent structures held together by strong covalent bonds. These elements have therefore very high melting/boiling points.
Both Carbon-graphite/ carbon-diamond have smaller atomic radius/size than Silicon in period 3 and thus higher melting/boiling points due to better/closer packing of smaller atoms in their well packed giant atomic/covalent structures.

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(ii) Non-metals from group V along/across the period form simple molecules joined by weak intermolecular /van-der-waals force. The weak intermolecular /van-der-waals force require little energy/heat to weaken leading to low melting/boiling points. The strength of the intermolecular /van-der-waals forces decrease with decrease in atomic radius/ size lowering the melting/boiling points along/across the period (and raising the melting/boiling points down the group).e.g.
-The melting /boiling point of Nitrogen is higher than that of Oxygen. This is because the atomic radius/ size of Nitrogen is higher than that of Oxygen and hence stronger intermolecular /van-der-waals forces between Nitrogen molecules.
-The melting /boiling point of Chlorine is higher than that of Fluorine. This is because the atomic radius/ size of Chlorine is higher than that of Fluorine and hence stronger intermolecular /van-der-waals forces between Chlorine molecules.
(iii) Rhombic sulphur exists as a puckered ring of S_8atoms which are well packed. Before melting the ring break and join to very long chains that entangle each other causing the **unusually** high melting/boiling point of Rhombic sulphur.
(iv) Both sulphur and phosphorus exists as allotropes.
Sulphur exists as **Rhombic**-sulphur and **monoclinic**-sulphur. Rhombic-sulphur is the stable form of sulphur at room temperature and pressure.
Phosphorus exists as **white**-phosphorus and **red**-phosphorus.
White-phosphorus is the stable form of Phosphorus at room temperature and pressure.

4. State and explain the trend in density of elements in a period in the periodic table.

Observation: Density increase upto the elements in group IV then falls across/along the period successively
Explanation:
Density is the mass per unit volume occupied by matter/particles/atoms/molecules of element.
(i)For metals ,the stronger metallic bond and the more delocalized electrons ensure a very well packed giant metallic structure that occupy less volume and thus higher density.The more the number of delocalized electrons along/across the period, the higher the density. e.g.
(i)Aluminium has a higher density than sodium. This is because aluminium has more /three delocalized electrons than /one sodium thus forms a very well packed giant metallic structure that occupy less volume per given mass/density.
(ii)Carbon-graphite ,carbon-diamond and silicon in group IV form a well packed giant atomic/covalent structure that is continuously joined by strong covalent bonds hence occupy less volume per given mass/density.

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Carbon-graphite form a less well packed giant hexagonal planar structure joined by Van-der-waals forces. Its density ($2.25gcm^{-3}$) is therefore less than that of Carbon-diamond($3.53gcm^{-3}$) and silicon($2.33gcm^{-3}$).Both diamond and silicon have giant tetrahedral structure that is better packed. Carbon-diamond has smaller atomic radius/size than silicon. Its density is thus higher because of better packing and subsequently higher density. Carbon-diamond is the hardest known natural substance by having the highest density.

(iii) For non-metals, the strength of the intermolecular /van-der-waals forces decreases with decrease in atomic radius/size along/across the period. This decreases the mass occupied by given volume of atoms in a molecule from group VI onwards. e.g.

Phosphorus has a higher atomic radius/size than chlorine and Argon and thus stronger intermolecular/van-der-waals forces that ensure a given mass of phosphorus occupy less volume than chlorine and neon.

5. State and explain the trend in thermal/electrical conductivity of elements in a period in the periodic table.

Observation:

Increase along/across the period from group I, II, and III then decrease in Group IV to drastically decrease in group V to VIII (O).

Explanation

(i)Metals have free delocalized electrons that are responsible for thermal/electrical conductivity.Thermal/electrical conductivity increase with increase in number of delocalized electrons. The thermal conductivity decrease with increase in temperature/heating. e.g. Aluminium with three delocalized electrons from each atom in its metallic structure has the highest electrical /thermal conductivity in period 3.

(ii)Carbon-graphite has also free 4th valency electrons that are delocalized within its layers of giant hexagonal planar structure. They are responsible for the electrical conductivity of graphite.

(iii)Silicon and carbon diamond do not conduct electricity but conducts heat. With each atom too close to each other in their very well packed giant tetrahedral structure, heat transfer /radiate between the atoms. The thermal conductivity increase with increase in temperature/heating.

(iv) All other non-metals are poor /non-conductor of heat and electricity. They are made of molecules with no free /mobile delocalized electrons in their structure.

Periodicity of the oxides of elements along/across period 3

The table below summarizes some properties of the oxides of elements in period 3 of the periodic table.

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Formular of oxide/ Property	**Na_2O**	**MgO**	**Al_2O_3**	**SiO_2**	**P_2O_5 P_4O_6**	**SO_2 SO_3**	**Cl_2O_7 Cl_2O**
Melting point(°C)	1193	3075	2045	1728	563	-76	-60
Boiling point(°C)	1278	3601	2980	2231	301	-10	-9
Bond type	Ionic	Ionic	Ionic	Covalent	Covalent	Covalent	Covalent
Chemical structure	Giant ionic structure	Giant ionic structure	Giant ionic structure	Giant atomic/ covalent	Simple molecula or molecule	Simple molecula or molecules	Simple molecula or molecule
State at room temperature	Solid	Solid	Solid	Solid	Solid	gas	Gas (**Cl_2O_7** is a liquid)
Nature of Oxide	Basic/ alkaline	Basic/ alkaline	Amphotellic oxide	2:8:4	2:8:5	2:8:6	2:8:7
Reaction with water	React to form NaOH /alkaline solution	React to form $MgOH)_2$ /weakly alkaline solution	Don't react with water.	Don't react with water.	React to form H_2PO_4 /weakly acidic solution	-SO_2 react to form H_2SO_3 . H_2SO_3 is quickly oxidized to H_2SO_4 -SO_2 react to form H_2SO_4/ strongly	**-Cl_2O_7 reacts to form $HClO_4$** /weakly acidic solution

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						acidic	
Reaction with dilute acids	Reacts to form salt and water	Reacts to form salt and water	Reacts to form salt and water	No reaction	No reaction	No reaction	No reaction

1. All the oxides of elements in period 3 except those of sulphur and chlorine are **solids** at room temperature and pressure.
2. Across/along the period, bonding of the oxides changes from **ionic** in sodium oxide magnesium oxide and aluminium oxide (show both ionic and covalent properties) to **covalent** in the rest of the oxides.
3. Across/along the period, the structure of the oxides changes from giant **ionic** structure in sodium oxide, magnesium oxide and aluminium oxide to giant **atomic**/covalent structure in silicon (IV) oxide. The rest of the oxides form simple **molecules**/**molecular** structure.
4. Sodium oxide and magnesium oxide are **basic /alkaline** in nature. Aluminium oxide is **amphotellic** in nature (shows both acidic and basic characteristics). The rest of the oxides are **acidic** in nature.
5. Ionic compounds/oxides have very **high** melting/boiling points because of the strong **electrostatic attraction** joining the giant ionic crystal lattice.
The melting/boiling points increase from sodium oxide to aluminium oxide as the number of electrons involved in bonding increase, increasing the strength of the ionic bond/structure.
6. Silicon (IV) oxide is made of a well packed giant atomic/covalent structure joined by strong covalent bonds.
This results in a solid with very **high** melting/boiling point.
7.Phosphorus (V) oxide, sulphur(IV) oxide/ sulphur (VI) oxide and dichloride heptoxide exist as simple molecules/molecular structure joined by weak van-der-waals/intermolecular forces.
This results in them existing as **low** melting /boiling point solids/gases.
8. Ionic oxide conducts electricity in molten and aqueous states but not in solid.
In solid state the ions are fused/fixed but on heating to molten state and when dissolved in water, the ions are free / mobile.
Sodium oxide, magnesium oxide and aluminium oxide are therefore good conductors in molten and aqueous states.
9. Covalent bonded oxides do not conduct electricity in solid, molten or in aqueous states.

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This is because they do not have free / mobile ion. Phosphorus (V) oxide, sulphur (IV) oxide/ sulphur (VI) oxide and dichloride heptoxide are thus non-conductors/insulators.

10. Silicon (IV) oxide is a poor/weak conductor of heat in solid state. This is because it has very closely packed structure for heat to radiate conduct along its structure.

11. Electopositivity decrease across the period while electronegativity increase across the period. The oxides thus become less ionic and more covalent along/across the period.

12.The steady change from giant ionic structure to giant atomic/ covalent structure then simple molecular structure lead to profound differences in the reaction of the oxides with water,acids and alkalis/bases:

(i) Reaction with water

a) Ionic oxides react with water to form alkaline solutions e.g.;

I.Sodium oxide reacts/dissolves in water forming an alkaline solution of sodium hydroxide.

Chemical equation: $Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$

II. Magnesium oxide slightly/ slowly reacts/dissolves in water forming an alkaline solution of magnesium hydroxide

Chemical equation: $MgO(s) + 2H_2O(l) \rightarrow Mg(OH)_2(aq)$

III. Aluminium oxide does reacts/dissolves in water.

b) Non-metallic oxides are acidic. They react with water to form weakly acidic solutions:

I. Phosphorus (V) oxide readily reacts/dissolves in water forming a weak acidic solution of phosphoric (V) acid.

Chemical equation: $P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(aq)$

Chemical equation: $P_2O_5(s) + 3H_2O(l) \rightarrow 2H_3PO_4(aq)$

II. Sulphur (IV) oxide readily reacts/dissolves in water forming a weak acidic solution of sulphuric (IV) acid.

Chemical equation: $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

Sulphur (VI) oxide quickly fumes in water to form concentrated sulphuric (VI) acid which is a strong acid.

Chemical equation: $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

III. Dichlorine oxide reacts with water to form weak acidic solution of chloric(I) acid/hypochlorous acid.

Chemical equation: $Cl_2O(g) + H_2O(l) \rightarrow 2HClO(aq)$

IV. Dichlorine heptoxide reacts with water to form weak acidic solution of chloric(VII) acid.

Chemical equation: $Cl_2O_7(l) + H_2O(l) \rightarrow 2HClO_4(aq)$

c) Silicon (IV) oxide **does not** react with water.

It reacts with hot concentrated alkalis forming silicate (IV) salts. e.g.

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Silicon (IV) oxide react with hot concentrated sodium hydroxide to form sodium silicate (IV) salt.

Chemical equation: SiO_2 (s) + 2NaOH (aq) -> Na_2SiO_3 (aq) + H_2O (l)

(ii) Reaction with dilute acids

a) Ionic oxides react with dilute acids to form salt and water only. This is a **neutralization** reaction. e.g.

Chemical equation: Na_2O(s) + H_2SO_4 (aq) -> Na_2SO_4 (aq) + H_2O(l)

Chemical equation: MgO(s) + $2HNO_3$(aq) -> Mg $(NO_3)_2$ (aq) + H_2O(l)

Chemical equation: Al_2O_3 (s) + 6HCl(aq) -> $2AlCl_3$ (aq) + $3H_2O$(l)

Aluminium oxide is amphotellic and reacts with **hot concentrated strong** alkalis sodium/potassium hydroxides to form **complex** sodium aluminate(III) and potassium aluminate(III) salt.

Chemical equation: Al_2O_3 (s) + 2NaOH (aq) + 3H2O(l) -> 2 $NaAl(OH)_4$ (aq)

Chemical equation: Al_2O_3 (s) + 2KOH(aq) + $3H_2O$(l) -> 2 $KAl(OH)_4$ (aq)

b) Acidic oxides **do not** react with dilute acids.

c) Periodicity of the Chlorides of elements along/across period 3

The table below summarizes some properties of the chlorides of elements in period 3 of the periodic table.

Formular of chloride/ Property	**NaCl**	**$MgCl_2$**	**$AlCl_3$**	**$SiCl_4$**	**PCl_5 PCl_3**	**SCl_2 S_2Cl_2**	**Cl_2**
Melting point(°C)	801	714	Sublimes at 180 °C	-70	PCl_5 Sublimes at -94 °C	-78	-101
Boiling point(°C)	1465	1418	423(as **Al_2Cl_6** vapour	57	74(as **P_2Cl_6** Vapour 164 (as **PCl_5)**	decomposes at 59 °C	-34
Bond type	Ionic	Ionic	Ionic/ Covalent/	Covalent	Covalent	Covalent	Covalent

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			dative				
Chemical structure	Giant ionic structure	Giant ionic structure	Molecular/ dimerizes	Simple molecula or molecule	Simple molecula or molecule	Simple molecula or molecules	Simple molecula or molecule
State at room temperature	Solid	Solid	Solid	liquid	Liquid **PCl_5 is solid**	liquid	Gas
Nature of Chloride	Neutral	Neutral	Strongly acidic	Strongly acidic	Strongly acidic	Strongly acidic	Strongly acidic
pH of solution	7.0	7.0	3.0	3.0	3.0	3.0	3.0
Reaction with water	Dissolve	Dissolve	- Hydrolysed by water -Acidic hydrogen chloride fumes produced	- Hydrolysed by water -Acidic hydrogen chloride fumes produced	Hydrolysed by water -Acidic hydrogen chloride fumes produced	Hydrolysed by water -Acidic hydrogen chloride fumes produced	Forms HCl and HClO
Electrical conductivity in molten/aqueous state	good	good	poor	nil	nil	nil	nil

1. Sodium Chloride, Magnesium chloride and aluminium chloride are **solids** at room temperature and pressure.
Silicon(IV) chloride, phosphorus(III)chloride and disulphur dichloride are **liquids.** Phosphorus(V)chloride is a **solid**. Both chlorine and sulphur chloride are **gases.**
2. Across/along the period bonding changes from **ionic** in Sodium Chloride and Magnesium chloride to **covalent** in the rest of the chlorides.

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3. Anhydrous aluminium chloride is also a molecular compound .Each aluminium atom is covalently bonded to three chlorine atoms.
In vapour/gaseous phase/state two molecules dimerizes to Al_2O_6 molecule through coordinate/dative bonding.
4. Across/along the period the structure changes from giant **ionic** in Sodium Chloride and Magnesium chloride to **simple molecules/molecular structure** in the rest of the chlorides.
5. Ionic chlorides have very high melting /boiling points because of the strong ionic bond/electrostatic attraction between the ions in their crystal lattice.The rest of the chlorides have low melting /boiling points because of the weak van-der-waal /intermolecular forces.
6. Sodium Chloride and Magnesium chloride in molten and aqueous state have free/mobile ions and thus good electrical conductors. Aluminium chloride is a poor conductor. The rest of the chlorides do not conduct because they have no free/mobile **ions**.
7. Ionic chloride form **neutral** solutions with pH =7. These chlorides **ionize**/**dissociate** completely into free cations and anions.i.e;
Sodium Chloride and Magnesium chloride have pH=7 because they are fully/completely ionized/dissociated into free ions.

Chemical equation $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$

Chemical equation $MgCl_2(s) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq)$

8 Across/along the period from aluminium chloride, **hydrolysis** of the chloride takes place when reacting/dissolved in water.
Hydrolysis is the reaction of a compound when dissolved in water.
a) Aluminium chloride is hydrolyzed by water to form aluminium **hydroxide** and fumes of hydrogen chloride gas. Hydrogen chloride gas dissolves in water to acidic hydrochloric acid. Hydrochloric acid is a strong acid with low pH and thus the mixture is strongly acidic.

Chemical equation $AlCl_3(s) + 3H_2O(l) \rightarrow Al(OH)_3(s) + 3HCl(g)$

b)Silicon(IV) chloride is hydrolyzed by water to form silicon(IV)**oxide** and fumes of hydrogen chloride gas. Hydrogen chloride gas dissolves in water to acidic hydrochloric acid. Hydrochloric acid is a strong acid with low pH and thus the mixture is strongly acidic.

Chemical equation $SiCl_4(l) + 2H_2O(l) \rightarrow SiO_2(s) + 4HCl(g)$

This reaction is highly exothermic producing /evolving a lot of **heat** that cause a rise in the temperature of the mixture.

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c) Both phosphoric (V) chloride and phosphoric (III) chloride are hydrolyzed by water to form phosphoric (V) **acid** and phosphoric (III) **acid** respectively. Fumes of hydrogen chloride gas are produced. Hydrogen chloride gas dissolves in water to acidic hydrochloric acid. Hydrochloric acid is a strong acid with low pH and thus the mixture is strongly acidic.

Chemical equation PCl_5 (s) + $4H_2O(l)$-> $H_3PO_4(aq)$ + 5HCl(g)

Chemical equation PCl_3 (s) + $3H_2O(l)$-> $H_3PO_4(aq)$ + 3HCl(g)

This reaction is also highly exothermic producing /evolving a lot of **heat** that cause a rise in the temperature of the mixture.

d) Disulphur dichloride similarly hydrolyzes in water to form **yellow** deposits of sulphur and produce a mixture of **sulphur (IV) oxide** and **hydrogen chloride** gas. Hydrogen chloride gas dissolves in water to acidic hydrochloric acid. Hydrochloric acid is a strong acid with low pH and thus the mixture is strongly acidic.

Chemical equation $2S_2Cl_2$ (l) + $2H_2O(l)$-> 3S(s) + $SO_2(g)$ + 4HCl(g)

D. COMPREHENSIVE REVISION QUESTIONS

1.The grid below represents periodic table. Study it and answer the questions that follow. The letters do not represent the actual symbols of the elements.

								A
B				G		H	E	C
	J		I	L				
D	***N***						M	

(a) (I) Indicate on the grid the position of an element represented by letter N whose electronic configuration of a divalent cation is 2:8:8 . (1 mark)

(II) Name the bond formed between D and H react. Explain your answer.(2 marks)

Ionic/electrovalent

D is electropositive thus donates two electrons to electronegative H

(III) Write an equation for the reaction between B and water. (1 mark)

Chemical equation **2B (s) + $2H_2O(l)$ -> 2BOH(aq) + H_2 (g)**

(IV) How do the atomic radii of I and L compare. Explain. (2 marks)

(V) In terms of structure and bonding explain why the oxide of G has lower melting point than oxide of L. (2 marks)

(b) Study the information given below and answer the questions that follow.

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Formula of compound N	NaCl	$MgCl_2$	Al_2Cl_6	$SiCl_4$	PCl_3	SCl_2
B.P(^{0}C)	1470	1420	Sublimes	60	75	60
M.P(^{0}C)	800	710	At 800^{0}C	-70	90	-80

(I)Why is the formula of aluminium chloride given as Al_2Cl_6 and not $AlCl_3$? (1 mark)

(II) Give two chlorides that are liquid at room temperature. Give a reason for the answer. (2 marks)

(III) Give a reason why Al_2Cl_6 has a lower melting point than $MgCl_2$ although both Al and Mg are metals. (1 mark)

(IV) Which of the chlorides would remain in liquid state for the highest temperature range explain why? (2 mark)

(Kakamega)

2. a) Study the information given below and answer the questions that follow.

Element	Atomic radius (nm)	Ionic radius (nm)	Formula of oxide	Melting point of oxide ('C)
P	0.364	0.421	A_2O	-119
Q	0.830	0.711	BO_2	837
R	0.592	0.485	E_2O_3	1466
S	0.381	0.446	G_2O_5	242
T	0.762	0.676	JO	1054

(i) Which elements are non-metals? Give a reason. (2mks)

(ii) Explain why the melting point of the oxide of R is higher than that of the oxide of S. (2mks)

(iii) Give **two** elements that would react vigorously with each other. Explain your answer. (2mks)

b) Study the information in the table below and answer the questions that follow (The letters do not represent the actual symbols of the elements)

		Ionization Energy_kJ/Mole	
Element	Electronic configuration	1st ionization energy	2nd ionization energy
A	2.2	900	1800
B	2.8.2	736	1450
C	2.8.8.2	590	1150

(i) What chemical family do the elements A, B and C belong? (1mk)

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(ii) What is meant by the term ionization energy?
(1mk)

iii) The 2nd ionization energy is higher that the 1st ionization energy of each. Explain
(1mk)

(iv)When a piece of element C is placed in cold water, it sinks to the bottom and an effervescence of a colourless gas that burns explosively is produced. Use a simple diagram to illustrate how this gas can be collected during this experiment.
(3mks)

3. The grid below represents part of the periodic table. The letters do not represent the actual symbols.

								A
B			X	G		Z	E	V
	J		I	L		T		
D	***N***						M	

a) Select the most reactive non-metal.
(1mk)

b) Write the formula of the compound consisting of
I.D and Z only.
(2mk)

II. X and Z

c) Select an element that can form an ion of change +2
(1mk)

d) Which element has the least ionization energy? Explain
(2mks)

e) Suggest with reasons a likely pH value of an aqueous solution of the chlorine of:(3mks)
T.
B
X

f) To which chemical family do the following elements belong? (2mk)
J
V

g) An element K has relative atomic mass of 40.2.It has two isotopes of masses 39 and 42. Calculate the relative abundance of each isotope. (3mks)

4.The grid below shows part of the periodic table study it and answer the questions that follow. The letters do not represent the true symbols.

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						A		
	B		C		D		E	
F	G							
							H	

\(a) Which element forms ions with charge of 2-. Explain (2mks)
(b) What is the nature of the oxide formed by C. (1mk)
(c) How does the reactivity of H compare with that of E. Explain? (2mks)
(d)Write down a balanced equation between B and Chlorine. (1mk)
(e) Explain how the atomic radii of F and G compare. (1mk)
(f) If the oxides of F and D are separately dissolved in water, state and explain the effects of their aqueous solutions on litmus. (3mks)

5. (a) The grid below show part of the periodic table.(The letter do not represent the actual symbols).Use it to answer the questions that follow.

T								Q
				S		R	K	
A	J		Y		U		L	
W							M	B
	C						N	
P								

(i)Select the most reactive non-metal. (1mk)
(ii)Select an element that forms a divalent cation. (1mk)
(iii)Element **Z** has atomic number 14.Show its position in the grid. (1mk)
(iv)How do the atomic radii of **U** and **J** compare?
(2mks)
(v)How do electrical conductivity of **A** and **Y** compare?
(2mks)
(vi)How does the boiling point of elements **K, L** and **M** vary? Explain (2mks

(b) The table below gives information on four elements by letters K, L, M and N. Study it and answer the questions that follow. The letters do not represent the actual symbols of the elements.

Element	Electron arrangement	Atomic radius	Ionic radius
K	2:8:2	0.136	0.065

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L	**2:8:7**	**0.099**	**0.181**
M	**2:8:8:1**	**0.203**	**0.133**
N	**2:8:8:2**	**0.174**	**0.099**

(a) Which two elements have similar properties? Explain. (2mks)

(b) Which element is a non-metal? Explain. (1mk)

(c) Which one of the elements is the strongest reducing agent. (1mk)

6. The grid given below represents part of the periodic table study it and answer the questions that follow. (The letters do not represent the actual symbols of the elements.)

									A
					B				
	C			**D**			**E**		
	F								

(i) What name is given to the group of elements to which C and F belong? (1mk)

(ii) Which letter represents the element that is the least reactive? (1mk)

(iii) What type of bond is formed when B and E react? Explain (2mks)

(iv)Write formula of the compound formed where elements D and oxygen gas react. (1mk)

(v) On the grid indicate the a tick (√) the position of element G which is in the third period of the periodic table and forms G^{3-} ions. (1mk)

(b) Study the information in the table below and answer the questions that follow. (The letter do not represents the actual symbols of the substance).

Substance	Melting point ^oC	Boiling point ^oC	Solubility in water	Density at room. Temp/g/cm^3
H	-117	78.5	Very soluble	0.8
J	-78	-33	Very soluble	0.77×1^{-3}
K	-23	77	Insoluble	1.6
L	- 219	-183	Slightly Soluable	1.33×10^{-3}

I.(i) Which substance would dissolve in water and could be separated from the solution by fractional distillation. (1mk)

(ii) Which substances is a liquid at room temperature and when mixed with water two layers would be formed? (1mk)

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II. Which letter represents a substance that is a gas at room temperature and which can be collected;
(i) Over water? (1mk)
(ii) By downward displacement of air? Density of air at room temperature = 1.29 x 10^{-3} g/C (1mk)

A.ATOMIC STRUCTURE

The atom is the smallest particle of an element that take part in a chemical reaction. The atom is made up of three subatomic particles:

(i)Protons
(ii)Electrons
(iii)Neutrons

(i)Protons

1. The proton is positively charged
2. Is found in the centre of an atom called nucleus
3. It has a relative mass 1
4. The number of protons in a atom of an element is its Atomic number

(ii)Electrons

1. The Electrons is negatively charged
2. Is found in fixed regions surrounding the centre of an atom called energy levels/orbitals.
3. It has a relative mass $^{1}/_{1840}$
4. The number of protons and electrons in a atom of an element is always equal

(iii)Neutrons

1. The Neutron is neither positively nor negatively charged thus neutral.
2. Like protons it is found in the centre of an atom called nucleus
3. It has a relative mass 1
4. The number of protons and neutrons in a atom of an element is its Mass number

Diagram showing the relative positions of protons, electrons and neutrons in an atom of an element

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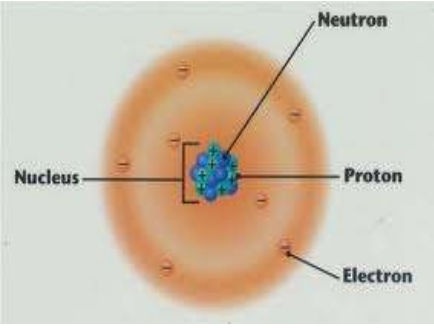
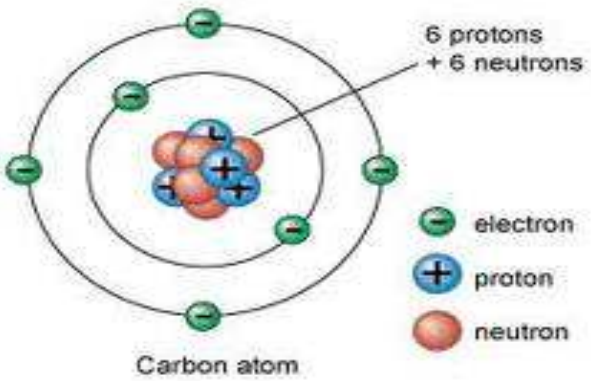


Diagram showing the relative positions of protons, electrons and neutrons in an atom of Carbon



The table below show atomic structure of the 1st twenty elements.

Element	Symbol	Protons	Electrons	Neutrons	Atomic number	Mass number
Hydrogen	**H**	**1**	**1**	**0**	**1**	**1**
Helium	**He**	**2**	**2**	**2**	**2**	**4**
Lithium	**Li**	**3**	**3**	**4**	**3**	**7**
Beryllium	**Be**	**4**	**4**	**5**	**4**	**9**
Boron	**B**	**5**	**5**	**6**	**5**	**11**
Carbon	**C**	**6**	**6**	**6**	**6**	**12**
Nitrogen	**N**	**7**	**7**	**7**	**7**	**14**
Oxygen	**O**	**8**	**8**	**8**	**8**	**16**
Fluorine	**F**	**9**	**9**	**10**	**9**	**19**
Neon	**Ne**	**10**	**10**	**10**	**10**	**20**

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Sodium	**Na**	**11**	**11**	**12**	**11**	**23**
Magnesium	**Mg**	**12**	**12**	**12**	**12**	**24**
Aluminium	**Al**	**13**	**13**	**14**	**13**	**27**
Silicon	**Si**	**14**	**14**	**14**	**14**	**28**
Phosphorus	**P**	**15**	**15**	**16**	**15**	**31**
Sulphur	**S**	**16**	**16**	**16**	**16**	**32**
Chlorine	**Cl**	**17**	**17**	**18**	**17**	**35**
Argon	**Ar**	**18**	**18**	**22**	**18**	**40**
Potassium	**K**	**19**	**19**	**20**	**19**	**39**
Calcium	**Ca**	**20**	**20**	**20**	**20**	**40**

Most atoms of elements exist as **isotopes.**

Isotopes are atoms of the same **element**, having the same number of **protons**/atomic number but **different** number of **neutrons**/mass number.

By convention, isotopes are written with the mass number as superscript and the atomic number as subscript to the left of the chemical symbol of the element. i.e.

mass number → $^{m}_{n}X$ ← symbol of element

atomic number →

Below is the conventional method of writing the 1st twenty elements showing the mass numbers and atomic numbers:

$^{1}_{1}H$ $^{4}_{2}He$ $^{7}_{3}Li$ $^{9}_{4}Be$ $^{11}_{5}B$ $^{12}_{6}C$

$^{14}_{7}N$ $^{16}_{8}O$ $^{19}_{9}F$ $^{20}_{10}Ne$ $^{23}_{11}Na$ $^{24}_{12}Mg$

$^{27}_{13}Al$ $^{28}_{14}Si$ $^{31}_{15}P$ $^{32}_{16}S$ $^{35}_{17}Cl$ $^{40}_{18}Ar$

$^{39}_{19}K$ $^{40}_{20}C$

The table below shows some common **natural** isotopes of some elements

Element	Isotopes	Protons	Electrons	Neutrons	Atomic number	Mass number
Hydrogen	$^{1}_{1}H$ $^{2}_{1}H$(deuterium) $^{3}_{1}H$(Tritium)	1 1 1	1 1 1	0 2 3	1 1 1	1 2 3
Chlorine	$^{35}_{17}Cl$ $^{37}_{17}Cl$	17 17	17 17	18 20	17 17	35 37

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Potassium	$^{39}_{19}K$ $^{40}_{19}K$ $^{41}_{19}K$	19 19 19	19 19 19	20 21 22	19 19 19	39 40 41
Oxygen	$^{16}_{8}O$ $^{18}_{8}O$	8 8	8 8	8 10	8 8	16 18
Uranium	$^{235}_{92}U$ $^{238}_{92}U$	92 92	92 92	143 146	92 92	235 238
Neon	$^{22}_{10}Ne$ $^{20}_{10}Ne$ $^{21}_{10}Ne$	10 10 10	10 10 10	12 10 11	10 10 10	22 20 21

The mass of an average atom is very small (10^{-22} g).Masses of atoms are therefore expressed in relation to a chosen element.

The atom recommended is ^{12}C isotope whose mass is arbitrarily assigned as 12.000 atomic mass units(a.m.u) .

All other atoms are compared to the mass of ^{12}C isotope to give the **relative at** The relative atomic mass(RAM) is therefore defined as "**the mass of average atom of an element compared to $^{1}/_{12}$ an atom of ^{12}C isotope whose mass is arbitrarily fixed as 12.000 atomic mass units(a.m.u)** " i.e;

$$RAM = \frac{\text{mass of atom of an element}}{^{1}/_{12} \text{ of one atom of } ^{12}C \text{ isotope}}$$

Accurate relative atomic masses (RAM) are got from the **mass spectrometer.** Mass spectrometer determines the **isotopes** of the element and their relative **abundance/availability.**

Using the relative abundances/availability of the isotopes, the relative atomic mass (RAM) can be determined /calculated as in the below examples.

a) Chlorine occurs as 75% $^{35}_{17}Cl$ and 25% $^{37}_{17}Cl$ isotopes. Calculate the relative atomic mass of Chlorine.

Working

100 atoms of chlorine contains 75 atoms of $^{35}_{17}Cl$ isotopes

100 atoms of chlorine contains 75 atoms of $^{37}_{17}Cl$ isotopes

Therefore;

RAM of chlorine = (75/100 x 35) + 25/100 x 37 = **35.5**

Note that:

Relative atomic mass has no units

More atoms of chlorine exist as $^{35}_{17}Cl$(75%) than as $^{37}_{17}Cl$(25%) therefore RAM is nearer to the more abundant isotope.

b) Calculate the relative atomic mass of potassium given that it exist as;

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93.1% $^{39}_{19}K$, 0.01% $^{40}_{19}K$, 6.89% $^{41}_{19}K$,

Working

100 atoms of potassium contains 93.1 atoms of $^{39}_{19}K$ isotopes
100 atoms of potassium contains 0.01 atoms of $^{40}_{19}K$ isotopes
100 atoms of potassium contains 6.89 atoms of $^{41}_{19}K$ isotopes
Therefore;
RAM of potassium = (93.1/100 x39) + (0.01/100 x 40) +(6.89 /100 x 39)
Note that:
Relative atomic mass has no units
More atoms of potassium exist as $^{39}_{19}K$ (93.1%) therefore RAM is nearer to the more abundant $^{39}_{19}K$ isotope.

c) Calculate the relative atomic mass of Neon given that it exist as;
90.92% $^{20}_{10}Ne$, 0.26% $^{21}_{10}Ne$, 8.82% $^{22}_{10}Ne$,

Working

100 atoms of Neon contains 90.92 atoms of $^{20}_{10}Ne$ isotopes
100 atoms of Neon contains 0.26 atoms of $^{21}_{10}Ne$ isotopes
100 atoms of Neon contains 8.82 atoms of $^{22}_{10}Ne$ isotopes Therefore;
RAM of Neon = (90.92/100 x20) + (0.26/100 x 21) +(8.82 /100 x 22)

Note that:
Relative atomic mass has no units
More atoms of Neon exist as $^{20}_{10}Ne$ (90.92%) therefore RAM is nearer to the more abundant $^{20}_{10}Ne$ isotope.

d) Calculate the relative atomic mass of Argon given that it exist as;
90.92% $^{20}_{10}Ne$, 0.26% $^{21}_{10}Ne$, 8.82% $^{22}_{10}Ne$,

NB

The relative atomic mass is a measure of the masses of atoms. The higher the relative atomic mass, the heavier the atom.
Electrons are found in **energy levels/orbital.**
An energy level is a **fixed region** around/surrounding the nucleus of an atom occupied by electrons of the **same (potential) energy.**
By convention energy levels are named 1,2,3… **outwards** from the region **nearest** to nucleus.
Each energy level is occupied by a fixed number of electrons:

The **1st** energy level is occupied by a maximum of **two** electrons
The **2nd** energy level is occupied by a maximum of **eight** electrons
The **3rd** energy level is occupied by a maximum of **eight** electrons(or **eighteen** electrons if available)

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The **4th** energy level is occupied by a maximum of **eight** electrons(or **eighteen or thirty two** electrons if available)

This arrangement of electrons in an atom is called **electron configuration / structure**.

By convention the electron configuration / structure of an atom of an element can be shown in form of a diagram using either cross(**x**) or dot(●) to

Practice examples drawing electronic configurations

a) $^{1}_{1}H$ has - in nucleus **1** proton and **0** neutrons

- 1 electron in the 1st energy levels thus:

Nucleus

Energy levels

Electrons (represented by cross(x)

Electronic structure of Hydrogen is thus: **1:**

b) $^{4}_{2}He$ has - in nucleus **2** proton and **2** neutrons

- 2 electron in the 1st energy levels thus:

Nucleus

Energy levels

Electrons (represented by cross(x)

Electronic structure of Helium is thus: **2:**

c) $^{7}_{3}Li$ has - in nucleus **3** proton and **4** neutrons

- **2** electron in the 1st energy levels

-**1** electron in the 2nd energy levels thus

Nucleus

Energy levels

Electrons (represented by cross(x)

Electronic structure of Lithium is thus: **2:1**

d) $^{9}_{4}Be$ has - in nucleus **4** proton and **5** neutrons

- **2** electron in the 1st energy levels

-**2** electron in the 2nd energy levels thus

Nucleus

Energy levels

Electrons (represented by cross(x)

Electronic structure of Beryllium is thus: **2:2**

e) $^{11}_{5}B$ has - in nucleus **5** proton and **6** neutrons

- **2** electron in the 1st energy levels

-**3** electron in the 2nd energy levels thus

Nucleus

Commented [s1]:

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Energy levels
Electrons (represented by cross(x)
Electronic structure of Boron is thus: **2:3**

f) $^{12}_{6}C$ has - in nucleus **6** proton and **6** neutrons
- **2** electron in the 1st energy levels
-**4** electron in the 2nd energy levels thus

Nucleus
Energy levels
Electrons (represented by cross(x)
Electronic structure of Carbon is thus: **2:4**

g) $^{14}_{7}N$ has - in nucleus **7** proton and **7** neutrons
- **2** electron in the 1st energy levels
-**5** electron in the 2nd energy levels thus

Nucleus
Energy levels
Electrons (represented by cross(x)
Electronic structure of Nitrogen is thus: **2:5**

h) $^{16}_{8}O$ has - in nucleus **8** proton and **8** neutrons
- **2** electron in the 1st energy levels
-**6** electron in the 2nd energy levels thus

Nucleus
Energy levels
Electrons (represented by cross(x)
Electronic structure of Oxygen is thus: **2:6**

i) $^{19}_{9}F$ has - in nucleus **9** proton and **10** neutrons
- **2** electron in the 1st energy levels
-**7** electron in the 2nd energy levels thus

Nucleus
Energy levels
Electrons (represented by cross(x)
Electronic structure of Fluorine is thus: **2:7**

i) $^{20}_{10}Ne$ has - in nucleus **10** proton and **10** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels thus

Nucleus
Energy levels
Electrons (represented by cross(x)

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Electronic structure of Neon is thus: **2:8**

j) $^{23}_{11}$Na has - in nucleus **11** proton and **12** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**1** electron in the 3rd energy levels thus

Nucleus
Energy levels
Electrons (represented by dot (.)
Electronic structure of Sodium is thus: **2:8:1**

k) $^{24}_{12}$Mg has - in nucleus **12** proton and **12** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**2** electron in the 3rd energy levels thus

Nucleus
Energy levels
Electrons (represented by dot (.)
Electronic structure of Magnesium is thus: **2:8:2**

l) $^{27}_{13}$Al has - in nucleus **13** proton and **14** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**3** electron in the 3rd energy levels thus

Nucleus
Energy levels
Electrons (represented by dot (.)
Electronic structure of Aluminium is thus: **2:8:3**

m) $^{28}_{14}$Si has - in nucleus **14** proton and **14** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**4** electron in the 3rd energy levels thus

Nucleus
Energy levels
Electrons (represented by dot (.)
Electronic structure of Silicon is thus: **2:8:4**

n) $^{31}_{15}$P has - in nucleus **14** proton and **15** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**5** electron in the 3rd energy levels thus

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Nucleus
Energy levels
Electrons (represented by dot (.)
Electronic structure of Phosphorus is thus: **2:8:5**

o) $^{32}_{16}S$ has - in nucleus **16** proton and **16** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**6** electron in the 3rd energy levels thus

Nucleus
Energy levels
Electrons (represented by dot (.)
Electronic structure of Sulphur is thus: **2:8:6**

p) $^{35}_{17}Cl$ has - in nucleus **18** proton and **17** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**7** electron in the 3rd energy levels thus

Nucleus
Energy levels
Electrons (represented by dot (.)
Electronic structure of Chlorine is thus: **2:8:7**

p) $^{40}_{18}Ar$ has - in nucleus **22** proton and **18** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**8** electron in the 3rd energy levels thus

Nucleus
Energy levels
Electrons (represented by dot (.)
Electronic structure of Argon is thus: **2:8:8**

q) $^{39}_{19}K$ has - in nucleus **20** proton and **19** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**8** electron in the 3rd energy levels
-**1** electron in the 4th energy levels thus

Nucleus
Energy levels
Electrons (represented by dot (.)
Electronic structure of Potassium is thus: **2:8:8:1**

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r) $^{40}_{20}Ca$ has - in nucleus **20** proton and **20** neutrons
- **2** electron in the 1st energy levels
-**8** electron in the 2nd energy levels
-**8** electron in the 3rd energy levels
-**2** electron in the 4th energy levels thus

Nucleus
Energy levels
Electrons (represented by dot (.)
Electronic structure of Calcium is thus: **2:8:8:2**

B.PERIODIC TABLE

There are over 100 elements so far discovered. Scientists have tried to group them together in a periodic table.

A periodic table is a horizontal and vertical arrangement of elements according to their atomic numbers.

This table was successfully arranged in 1913 by the British scientist Henry Moseley from the previous work of the Russian Scientist Dmitri Mendeleev.

The horizontal arrangement forms **period**. Atoms in the same period have the same the same number of energy levels in their electronic structure. i.e.

The number of **energy levels** in the electronic configuration of an element determine the **period** to which the element is in the periodic table.

e.g.

Which period of the periodic table are the following isotopes/elements/atoms?

a) $^{12}_{6}C$
   Electron structure 2:4 => 2 energy levels used thus **Period 2**
b) $^{23}_{11}Na$
   Electron structure 2:8:1 => 3 energy levels used thus **Period 3**
c) $^{39}_{19}K$
   Electron structure 2:8:8:1 => 4 energy levels used thus **Period 4**
d) $^{1}_{1}H$
   Electron structure 1: => 1 energy level used thus **Period 1**

The vertical arrangement of elements forms a **group.** Atoms in the same have the same the same group have the same number of outer energy level electrons as per their electronic structure. i.e.

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The number of electrons in the outer energy level an element determine the **group** to which the element is, in the periodic table.

a) $^{12}_{6}C$

Electron structure 2:**4** => 4 electrons in outer energy level thus **Group IV**

b) $^{23}_{11}C$

Electron structure 2:8:**1** => 1 electron in outer energy level thus **Group I**

c) $^{39}_{19}K$

Electron structure 2:8:8:**1**=>1 electron in outer energy level thus **Group I**

d) $^{1}_{1}H$

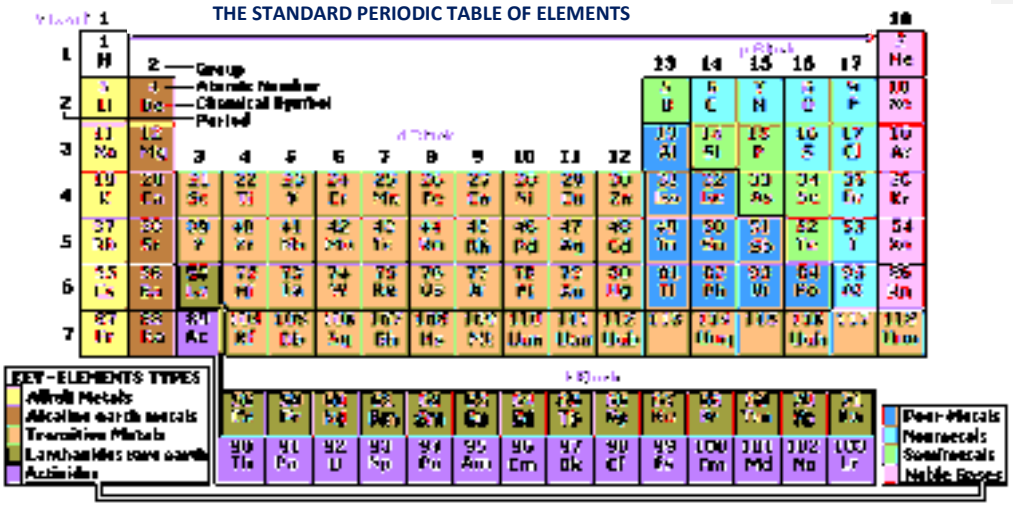
Electron structure **1**: => 1 electron in outer energy level thus **Group I**

By convention;

(i)**Periods** are named using English numerals **1, 2, 3, 4**…

(ii)**Groups** are named using Roman numerals **I, II, III, IV**…

There are eighteen groups in a standard periodic table.
There are seven periods in a standard periodic table.



When an atom has maximum number of electrons in its outer energy level, it is said to be **stable.**

When an atom has no maximum number of electrons in its outer energy level, it is said to be **unstable.**

All stable atoms are in group **8/18** of the periodic table. All other elements are unstable.

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All unstable atoms/isotopes try to be stable through chemical reactions. A chemical reaction involves gaining or losing outer electrons (electron transfer) .When electron transfer take place, an ion is formed.

An ion is formed when an unstable atom gain or donate electrons in its outer energy level in order to be stable. Whether an atom gain or donate electrons depend on the relative energy required to donate or gain extra electrons i.e.

Examples

1. $^{19}_{9}F$ has electronic structure/configuration 2:7.

It can donate the seven outer electrons to have stable electronic structure/configuration 2:.

It can gain one extra electron to have stable electronic structure/configuration 2:8. Gaining requires less energy, and thus Fluorine reacts by gaining one extra electrons.

2. $^{23}_{13}Al$ has electronic structure/configuration 2:8:3

It can donate the three outer electrons to have stable electronic structure/configuration 2:8.

It can gain five extra electrons to have stable electronic structure/configuration 2:8:8. Donating requires less energy, and thus Aluminium reacts by donating its three outer electrons.

Elements with **less** than four electrons in the outer energy level donates /lose the outer electrons to be stable and form a positively charged ion called **cation.**

A cation therefore has more protons (positive charge) than electrons (negative charge)

Generally metals usually form cation

Elements with **more** than four electrons in the outer energy level gain /acquire extra electrons in the outer energy level to be stable and form a negatively charged ion called **anion.**

An anion therefore has fewer protons (positive charge) than electrons (negative charge)

Generally non metals usually form anion. Except Hydrogen

The charge carried by an ion is equal to the number of electrons gained/acquired or donated/lost.

Examples of ion formation

1. $^{1}_{1}H$

H	->	H^+	+	e
(atom)		(monovalent cation)		(electrons donated/lost)

Electronic configuration 1: (No electrons remains)

2. $^{27}_{13}Al$

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	Al	->	Al^{3+}	+ 3e
	(atom)		(trivalent cation)	(3 electrons donated/lost)
Electron	2:8:3		2:8	
structure	(unstable)		(stable)	

3. $^{23}_{11}Na$

	Na	->	Na^{+}	+ e
	(atom)		(cation)	(1 electrons donated/lost)
Electron	2:8:1		2:8	
structure	(unstable)		(stable)	

4. $^{24}_{12}Mg$

	Mg	->	Mg^{2+}	+ 2e
	(atom)		(cation)	(2 electrons donated/lost)
Electron	2:8:1		2:8	
structure	(unstable)		(stable)	

5. $^{16}_{8}O$

	O	+ 2e	->	O^{2-}
	(atom)	(2 electrons gained/acquired)		(anion)
Electron	2:6			2:8
structure	(unstable)			(stable)

6. $^{14}_{7}N$

	N	+ 3e	->	N^{3-}
	(atom)	(3 electrons gained/acquired)		(anion)
Electron	2:5			2:8
structure	(unstable)			(stable)

7. $^{31}_{15}P$

	P	+ 3e	->	P^{3-}
	(atom)	(3 electrons gained/acquired)		(anion)
Electron	2:5			2:8
structure	(unstable)			(stable)

8. $^{19}_{9}F$

	F	+ e	->	F^{-}
	(atom)	(1 electrons gained/acquired)		(anion)
Electron	2:7			2:8
structure	(unstable)			(stable)

9. $^{35}_{17}Cl$

	Cl	+ e	->	Cl^{-}
	(atom)	(1 electrons gained/acquired)		(anion)

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Electron structure 2:8:7 (unstable) 2:8:8 (stable)

3. $^{39}_{19}K$

K -> K^+ + e

(atom) (cation) (1 electrons donated/lost)

Electron structure 2:8:8:1 (unstable) 2:8:8 (stable)

When an element donates/loses its outer electrons, the process is called oxidation. When an element acquires/gains extra electrons in its outer energy level, the process is called reduction. The charge carried by an atom, cation or anion is its oxidation state.

Table showing the oxidation states of some isotopes

Element	Symbol of element / isotopes	Charge of ion	Oxidation state
Hydrogen	$^{1}_{1}H$ $^{2}_{1}H$(deuterium) $^{3}_{1}H$(Tritium)	H^+ H^+ H^+	+1 +1 +1
Chlorine	$^{35}_{17}Cl$ $^{37}_{17}Cl$	Cl^- Cl^-	-1 -1
Potassium	$^{39}_{19}K$ $^{40}_{19}K$ $^{41}_{19}K$	K^+ K^+ K^+	+1 +1 +1
Oxygen	$^{16}_{8}O$ $^{18}_{8}O$	O^{2-} O^{2-}	-2 -2
Magnesium	$^{24}_{12}Mg$	Mg^{2+}	+2
sodium	$^{23}_{11}Na$	Na^+	+1
Copper	Cu	Cu^+ Cu^{2+}	+1 +2
Iron		Fe^{2+} Fe^{3+}	+2 +3
Lead		Pb^{2+} Pb^{4+}	+2 +4

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Manganese		Mn^{2+} Mn^{7+}	+2 +7
Chromium		Cr^{3+} Cr^{6+}	+3 +6
Sulphur		S^{4+} S^{6+}	+4 +6
Carbon		C^{2+} C^{4+}	+2 +4

Note:
Some elements can exist in more than one oxidation state. They are said to have variable oxidation state.
Roman capital numeral is used to indicate the oxidation state of an element with a variable oxidation state in a compound.

Examples:

(i) Copper (I) means Cu^{+} as in Copper(I)oxide
(ii) Copper (II) means Cu^{2+} as in Copper(II)oxide
(iii) Iron (II) means Fe^{2+} as in Iron(II)sulphide
(iv) Iron (III) means Fe^{3+} as in Iron(III)chloride
(iv) Sulphur(VI)mean S^{6+} as in Iron(III)sulphate(VI)
(v) Sulphur(VI)mean S^{6+} as in sulphur(VI)oxide
(vi) Sulphur(IV)mean S^{4+} as in sulphur(IV)oxide
(vii) Sulphur(IV)mean S^{4+} as in sodium sulphate(IV)
(ix) Carbon(IV)mean C^{4+} as in carbon(IV)oxide
(x) Carbon(IV)mean C^{4+} as in Lead(II)carbonate(IV)
(xi) Carbon(II)mean C^{2+} as in carbon(II)oxide
(xii) Manganese(IV)mean Mn^{4+} as in Manganese(IV)oxide

A compound is a combination of two or more elements in fixed proportions. The ratio of the atoms making a compound is called the chemical formulae. Elements combine together to form a compound depending on their combining power.
The combining power of atoms in an element is called Valency. Valency of an element is equal to the **number** of:

(i) Hydrogen atoms that an atom of element can combine with or displace.
(ii) Electrons gained /acquired in outer energy level by non metals to be stable/attain duplet/octet.
(iii) Electrons donated/lost by outer energy level of metals to be stable/attain octet/duplet.

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(iv) Charges carried by ions/cations/ions

Group of atoms that react as a unit during chemical reactions are called **radicals**.

Elements with variable oxidation state also have more than one valency.

Table showing the valency of common radicals

Radical name	**Chemical formulae**	**Combining power / Valency**
Ammonium	NH_4^+	1
Hydroxide	OH^-	1
Nitrate(V)	NO_3^-	1
Hydrogen carbonate	HCO_3^-	1
Hydrogen sulphate(VI)	HSO_4^-	1
Hydrogen sulphate(IV)	HSO_3^-	1
Manganate(VII)	MnO_4^-	1
Chromate(VI)	CrO_4^{2-}	2
Dichromate(VI)	$Cr_2O_7^{2-}$	2
Sulphate(VI)	SO_4^{2-}	2
Sulphate(IV)	SO_3^{2-}	2
Carbonate(IV)	CO_3^{2-}	2
Phosphate(V)	PO_4^{2-}	3

Table showing the valency of some common metal and non metals

Element/metal	Valency	**Element/non metal**	**Valency**
Hydrogen	1	**Florine**	**1**
Lithium	1	**Chlorine**	**1**
Beryllium	2	**Bromine**	**1**
Boron	3	**Iodine**	**1**
Sodium	1	**Carbon**	**4**
Magnesium	2	**Nitrogen**	**3**
Aluminium	3	**Oxygen**	**2**
Potassium	1	**Phosphorus**	**3**
Calcium	2		
Zinc	2		
Barium	2		
Mercury	2		
Iron	2 and 3		

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Copper	1 and 2		
Manganese	2 and 4		
Lead	2 and 4		

From the valency of elements, the chemical formular of a compound can be derived using the following procedure:

(i)Identify the elements and radicals making the compound
(ii)Write the symbol/formular of the elements making the compound starting with the metallic element
(iii)Assign the valency of each element /radical as superscript.
(iv)Interchange/exchange the valencies of each element as subscript.
(v)Divide by the smallest/lowest valency to derive the smallest whole number ratios
Ignore a valency of 1.
This is the chemical formula.

Practice examples

Write the chemical formula of

(a)Aluminium oxide

Elements making compound	Aluminium	Oxygen
Symbol of elements/radicals in compound	Al	O
Assign valencies as superscript	Al^{3}	O^{2}
Exchange/Interchange the valencies as subscript	Al_{2}	O_{3}
Divide by smallest valency to get whole number	-	-

Chemical formula of Aluminium oxide is thus: **$Al_2 O_3$**

This means: 2atoms of Aluminium combine with 3 atoms of Oxygen

(b)Sodium oxide

Elements making compound	Sodium	Oxygen
Symbol of elements/radicals in compound	Na	O
Assign valencies as superscript	Na^{1}	O^{2}
Exchange/Interchange the valencies as subscript	Na_{2}	O_{1}
Divide by smallest valency to get whole number	-	-

Chemical formula of Sodium oxide is thus: **$Na_2 O$**

This means: 2atoms of Sodium combine with 1 atom of Oxygen

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(c)Calcium oxide

Elements making compound	Calcium	Oxygen
Symbol of elements/radicals in compound	Ca	O
Assign valencies as superscript	Ca^{2}	O^{2}
Exchange/Interchange the valencies as subscript	Ca_{2}	O_{2}
Divide by two to get smallest whole number ratio	Ca_{1}	O_{1}

Chemical formula of Calcium oxide is thus**: CaO**

This means: 1 atom of calcium combine with 1 atom of Oxygen.

(d)Lead (IV) oxide

Elements making compound	Lead	Oxygen
Symbol of elements/radicals in compound	Pb	O
Assign valencies as superscript	Pb^{4}	O^{2}
Exchange/Interchange the valencies as subscript	Pb_{2}	O_{4}
Divide by two to get smallest whole number ratio	Pb_{1}	O_{2}

Chemical formula of Lead (IV) oxide is thus**: PbO_2**

This means: 1 atom of lead combine with 2 atoms of Oxygen.

(e)Lead (II) oxide

Elements making compound	Lead	Oxygen
Symbol of elements/radicals in compound	Pb	O
Assign valencies as superscript	Pb^{2}	O^{2}
Exchange/Interchange the valencies as subscript	Pb_{2}	O_{2}
Divide by two to get smallest whole number ratio	Pb_{1}	O_{1}

Chemical formula of Lead (II) oxide is thus**: PbO**

This means: 1 atom of lead combine with 1 atom of Oxygen.

(e)Iron (III) oxide

Elements making compound	Iron	Oxygen
Symbol of elements/radicals in compound	Fe	O
Assign valencies as superscript	Fe^{3}	O^{2}
Exchange/Interchange the valencies as subscript	Fe_{2}	O_{3}
Divide by two to get smallest whole number ratio	-	-

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Chemical formula of Iron(III) oxide is thus: **Fe_2O_3**

This means: 2 atom of lead combine with 3 atom of Oxygen.

(f)Iron (II) sulphate (VI)

Elements making compound	Iron	sulphate(VI)
Symbol of elements/radicals in compound	Fe	SO_4
Assign valencies as superscript	Fe^{2}	$SO_4{}^{2}$
Exchange/Interchange the valencies as subscript	Fe_2	$SO_{4\ 2}$
Divide by two to get smallest whole number ratio	Fe_1	$SO_{4\ 1}$

Chemical formula of Iron (II) sulphate (VI) is thus: **$FeSO_4$**

This means: 1 atom of Iron combine with 1 sulphate (VI) radical.

(g)Copper (II) sulphate (VI)

Elements making compound	Copper	sulphate(VI)
Symbol of elements/radicals in compound	Cu	SO_4
Assign valencies as superscript	Cu^{2}	$SO_4{}^{2}$
Exchange/Interchange the valencies as subscript	Cu_2	$SO_{4\ 2}$
Divide by two to get smallest whole number ratio	Cu_1	$SO_{4\ 1}$

Chemical formula of Cu(II)sulphate(VI) is thus: **$CuSO_4$**

This means: 1 atom of Copper combine with 1 sulphate (VI) radical.

(h)Aluminium sulphate (VI)

Elements making compound	Aluminium	sulphate(VI)
Symbol of elements/radicals in compound	Al	SO_4
Assign valencies as superscript	Al^{3}	$SO_4{}^{2}$
Exchange/Interchange the valencies as subscript	Al_2	$SO_{4\ 3}$
Divide by two to get smallest whole number ratio	-	-

Chemical formula of Aluminium sulphate (VI) is thus: **$Al_2(SO_4)_3$**

This means: 2 atom of Aluminium combine with 3 sulphate (VI) radical.

(i)Aluminium nitrate (V)

Elements making compound	Aluminium	nitrate(V)
Symbol of elements/radicals in compound	Al	NO_3
Assign valencies as superscript	Al^{3}	$NO_3{}^{1}$
Exchange/Interchange the valencies as subscript	Al_1	$NO_{3\ 3}$

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Divide by two to get smallest whole number ratio	-	-

Chemical formula of Aluminium sulphate (VI) is thus: **$Al(NO_3)_3$**

This means: 1 atom of Aluminium combine with 3 nitrate (V) radical.

(j)Potassium manganate (VII)

Elements making compound	Potassium	manganate(VII)
Symbol of elements/radicals in compound	K	MnO_4
Assign valencies as superscript	K^1	MnO_4^1
Exchange/Interchange the valencies as subscript	K_1	$MnO_{4\,1}$
Divide by two to get smallest whole number ratio	-	-

Chemical formula of Potassium manganate (VII) is thus: **$KMnO_4$**

This means: 1 atom of Potassium combine with 4 manganate (VII) radical.

(k)Sodium dichromate (VI)

Elements making compound	Sodium	dichromate(VI)
Symbol of elements/radicals in compound	Na	Cr_2O_7
Assign valencies as superscript	Na^1	$Cr_2O_7^2$
Exchange/Interchange the valencies as subscript	Na_2	$Cr_2O_{7\,1}$
Divide by two to get smallest whole number ratio	-	-

Chemical formula of Sodium dichromate (VI) is thus: **$Na_2\,Cr_2O_7$**

This means: 2 atom of Sodium combine with 1 dichromate (VI) radical.

(l)Calcium hydrogen carbonate

Elements making compound	Calcium	Hydrogen carbonate
Symbol of elements/radicals in compound	Ca	CO_3
Assign valencies as superscript	Ca^2	HCO_3^1
Exchange/Interchange the valencies as subscript	Ca_1	$HCO_{3\ 2}$
Divide by two to get smallest whole number ratio	-	-

Chemical formula of Calcium hydrogen carbonate is thus: **$Ca(HCO_3)_2$**

This means: 1 atom of Calcium combine with 2 hydrogen carbonate radical.

(l)Magnesium hydrogen sulphate (VI)

Elements making compound	Magnesium	Hydrogen

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		sulphate(VI)
Symbol of elements/radicals in compound	Mg	HSO_4
Assign valencies as superscript	Mg^{2}	HSO_4^{1}
Exchange/Interchange the valencies as subscript	Mg_1	$HSO_{4\ 2}$
Divide by two to get smallest whole number ratio	-	-

Chemical formula of Magnesium hydrogen sulphate (VI) is thus**: $Mg(HSO_4)_2$**

This means: 1 atom of Magnesium combine with 2 hydrogen sulphate (VI) radical.

Compounds are formed from chemical reactions. A chemical reaction is formed when atoms of the reactants break free to bond again and form products. A chemical reaction is a statement showing the movement of reactants to form products. The following procedure is used in writing chemical equations:

1. Write the word equation
2. Write the correct chemical formula for each of the reactants and products
3. Check if the number of atoms of **each** element on the reactant side is **equal** to the number of atoms of each element on the product side.
4. Multiply the chemical formula containing the unbalanced atoms with the lowest common multiple if the number of atoms on one side is not equal. This is called **balancing.**

Do not change the chemical formula of the products/reactants.

5. Assign in brackets, the physical state/state symbols of the reactants and products after each chemical formula as:
   (i) **(s)** for solids
   (ii) **(l)** for liquids
   (iii) **(g)** for gas
   (iv) **(aq)** for aqueous/dissolved in water to make a solution.

Practice examples

Write a balanced chemical equation for the following

(a) Hydrogen gas is prepared from reacting Zinc granules with dilute hydrochloric acid.

Procedure

1. Write the word equation
   Zinc + Hydrochloric acid -> Zinc chloride + hydrogen gas
2. Write the correct chemical formula for each of the reactants and products
   $Zn + HCl \rightarrow ZnCl_2 + H_2$

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3. Check if the number of atoms of **each** element on the reactant side is **equal** to the number of atoms of each element on the product side.

Number of atoms of Zn on the reactant side is equal to product side

One atom of H in HCl on the reactant side is not equal to two atoms in H_2 on product side.

One atom of Cl in HCl on the reactant side is not equal to two atoms in $ZnCl_2$ on product side.

4. Multiply the chemical formula containing the **unbalanced** atoms with the lowest common multiple if the number of atoms on one side is not equal.

Multiply HCl by "2" to get "2" Hydrogen and "2" Chlorine on product and reactant side.

$Zn + 2HCl \rightarrow ZnCl_2 + H_2$

5. Assign in brackets, the physical state/state symbols .

$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

(b) Oxygen gas is prepared from decomposition of Hydrogen peroxide solution to water

Procedure

1. Write the word equation

Hydrogen peroxide -> Water + oxygen gas

2. Write the correct chemical formula for each of the reactants and products

$H_2O_2 \rightarrow H_2O + O_2$

3. Check if the number of atoms of **each** element on the reactant side is **equal** to the number of atoms of each element on the product side.

Number of atoms of H on the reactant side is equal to product side

Two atom of O in H_2O_2 on the reactant side is not equal to three atoms (one in H_2O and two in O_2) on product side.

4. Multiply the chemical formula containing the **unbalanced** atoms with the lowest common multiple if the number of atoms on one side is not equal.

Multiply H_2O_2 by "2" to get "4" Hydrogen and "4" Oxygen on reactants

Multiply H_2O by "2" to get "4" Hydrogen and "2" Oxygen on product side

When the "2" Oxygen in O_2 and the"2" in H_2O are added on product side they are equal to the"4" Oxygen on reactants side.

$2H_2O_2 \rightarrow 2H_2O + O_2$

5. Assign in brackets, the physical state/state symbols.

$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$

(c) Chlorine gas is prepared from Potassium manganate (VII) reacting with hydrochloric acid to form potassium chloride solution, manganese (II) chloride solution, water and chlorine gas.

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Procedure

1. Write the word equation

Potassium manganate (VII) + Hydrochloric acid ->

Potassium chloride + manganese (II) chloride + chlorine **+**water

2. Write the correct chemical formula for each of the reactants and products

$KMnO_4$ + HCl -> KCl + $MnCl_2$ +H_2O + Cl_2

3. Check if the number of atoms of **each** element on the reactant side is **equal** to the number of atoms of each element on the product side.

Number of atoms of K and Mn on the reactant side is equal to product side

Two atom of H in H_2O on the product side is not equal to one atom on reactant side.

Four atom of O in $KMnO_4$ is not equal to one in H_2O

One atom of Cl in HCl on reactant side is not equal to three (one in H_2O and two in Cl_2)

4. Multiply the chemical formula containing the **unbalanced** atoms with the lowest common multiple if the number of atoms on one side is not equal.

Multiply HCl by "16" to get "16" Hydrogen and "16" Chlorine on reactants

Multiply KMnO4 by "2" to get "2" Potassium and "2" manganese, "2 x4 =8" Oxygen on reactant side.

Balance the product side to get:

2 $KMnO_4$ +16 HCl -> 2 KCl + 2 $MnCl_2$ +8 H_2O + 5 Cl_2

5. Assign in brackets, the physical state/state symbols.

2$KMnO_4$(s) +16 HCl(aq)-> 2 KCl (aq) + 2$MnCl_2$(aq)+8 H_2O(l)+5 Cl_2(g)

(d)Carbon (IV) oxide gas is prepared from Calcium carbonate reacting with hydrochloric acid to form calcium chloride solution, water and carbon (IV) oxide gas.

Procedure

1. Write the word equation

Calcium carbonate + Hydrochloric acid ->

Calcium **chloride solution+ water +carbon (IV)oxide**

2. Write the correct chemical formula for each of the reactants and products

$CaCO_3$ + HCl -> $CaCl_2$ +H_2O + CO_2

3. Check if the number of atoms of **each** element on the reactant side is **equal** to the number of atoms of each element on the product side.

4. Multiply the chemical formula containing the **unbalanced** atoms with the lowest common multiple if the number of atoms on one side is not equal.

5. Assign in brackets, the physical state/state symbols.

$CaCO_3$(s) + 2 HCl(aq) -> $CaCl_2$(aq) + H_2O(l) + CO_2(g)

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(d)Sodium hydroxide solution neutralizes hydrochloric acid to form salt and water.

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

(e)Sodium reacts with water to form sodium hydroxide and hydrogen gas.

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

(f)Calcium reacts with water to form calcium hydroxide and hydrogen gas

$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

(g)Copper (II) Oxide solid reacts with dilute hydrochloric acid to form copper (II) chloride and water.

$CuO(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2O(l)$

(h)Hydrogen sulphide reacts with Oxygen to form sulphur (IV) Oxide and water.

$2H_2S(g) + 3O_2(g) \rightarrow 2SO_2(g) + 2H_2O(l)$

(i)Magnesium reacts with steam to form Magnesium Oxide and Hydrogen gas.

$Mg(s) + 2H_2O(g) \rightarrow MgO(s) + H_2(g)$

(j)Ethane (C_2H_6) gas burns in air to form Carbon (IV) Oxide and water.

$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$

(k)Ethene (C_2H_4) gas burns in air to form Carbon (IV) Oxide and water.

$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$

(l)Ethyne (C_2H_2) gas burns in air to form Carbon (IV) Oxide and water.

$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$

C. PERIODICITY OF CHEMICAL FAMILES/DOWN THE GROUP

The number of valence electrons and the number of occupied energy levels in an atom of an element determine the position of an element in the periodic table .i.e The number of occupied energy levels determines the Period and the valence electrons determine the Group.

Elements in the same group have similar physical and chemical properties. The trends in physical and chemical properties of elements in the same group vary down the group. Elements in the same group thus constitute a chemical family.

(a) Group I elements: Alkali metals

Group I elements are called **Alkali metals** except Hydrogen which is a non metal. The alkali metals include:

Element	Symbol	Atomic number	Electron structure	Oxidation state	Valency
Lithium	Li	3	2:**1**	Li^+	1
Sodium	Na	11	2:8:**1**	Na^+	1

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Potassium	K	19	2:8:8:**1**	K^+	1
Rubidium	Rb	37	2:8:18:8:**1**	Rb^+	1
Caesium	Cs	55	2:8:18:18:8:**1**	Cs^+	1
Francium	Fr	87	2:8:18:32:18:8:**1**	Fr^+	1

All alkali metals atom has one electron in the outer energy level. They therefore are **monovalent**. They donate /lose the outer electron to have oxidation state M^+
The number of energy levels increases down the group from Lithium to Francium.
The more the number of energy levels the bigger/larger the atomic size. e.g.
The atomic size of Potassium is bigger/larger than that of sodium because Potassium has more/4 energy levels than sodium (3 energy levels).

Atomic and ionic radius

The distance between the centre of the nucleus of an **atom** and the outermost energy level occupied by electron/s is called **atomic radius.** Atomic radius is measured in **nanometers**(n).The higher /bigger the atomic radius the bigger /larger the atomic size.

The distance between the centre of the nucleus of an **ion** and the outermost energy level occupied by electron/s is called **ionic radius.** Ionic radius is also measured in **nanometers** (n).The higher /bigger the ionic radius the bigger /larger the size of the ion.

Atomic radius and ionic radius depend on the number of energy levels occupied by electrons. The more the number of energy levels the bigger/larger the atomic /ionic radius. e.g.

The atomic radius of Francium is bigger/larger than that of sodium because Francium has more/7 energy levels than sodium (3 energy levels).

Atomic radius and ionic radius of alkali metals increase down the group as the number of energy levels increases.

The atomic radius of alkali metals is bigger than the ionic radius. This is because alkali metals react by losing/donating the outer electron and hence lose the outer energy level.

Table showing the atomic and ionic radius of some alkali metals

Element	Symbol	Atomic number	Atomic radius(nM)	Ionic radius(nM)
Lithium	Li	3	0.133	0.060
Sodium	Na	11	0.157	0.095
Potassium	K	19	0.203	0.133

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The atomic radius of sodium is 0.157nM .The ionic radius of Na^+ is 0.095nM. This is because sodium reacts by donating/losing the outer electrons and hence the outer energy level. The remaining electrons/energy levels experience more effective / greater nuclear attraction/pull towards the nucleus reducing the atomic radius.

Electropositivity

The ease of donating/losing electrons is called electropositivity. All alkali metals are electropositive. Electropositivity increase as atomic radius increase. This is because the effective nuclear attraction on outer electrons decreases with increase in atomic radius. The outer electrons experience less nuclear attraction and can be lost/ donated easily/with ease. Francium is the most electropositive element in the periodic table because it has the highest/biggest atomic radius.

Ionization energy

The minimum amount of energy required to remove an electron from an atom of element in its gaseous state is called 1st **ionization energy**. The SI unit of ionization energy is **kilojoules per mole/kJmole^{-1} .**Ionization energy depend on atomic radius. The higher the atomic radius, the less effective the nuclear attraction on outer electrons/energy level and thus the lower the ionization energy. For alkali metals the 1st ionization energy decrease down the group as the atomic radius increase and the effective nuclear attraction on outer energy level electrons decrease.

e.g. The 1st ionization energy of sodium is 496 kJmole^{-1} while that of potassium is 419 kJmole^{-1} **.**This is because atomic radius increase and thus effective nuclear attraction on outer energy level electrons decrease down the group from sodium to Potassium. It requires therefore less energy to donate/lose outer electrons in Potassium than in sodium.

Physical properties

Soft/Easy to cut: Alkali metals are soft and easy to cut with a knife. The softness and ease of cutting increase down the group from Lithium to Francium. This is because an increase in atomic radius, decreases the strength of metallic bond and the packing of the metallic structure

Appearance: Alkali metals have a shiny grey metallic luster when freshly cut. The surface rapidly/quickly tarnishes on exposure to air. This is because the metal surface rapidly/quickly reacts with elements of air/oxygen.

Melting and boiling points: Alkali metals have a relatively low melting/boiling point than common metals like Iron. This is because alkali metals use only one delocalized electron to form a weak metallic bond/structure.

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Electrical/thermal conductivity: Alkali metals are good thermal and electrical conductors. Metals conduct using the outer mobile delocalized electrons. The delocalized electrons move randomly within the metallic structure.

Summary of some physical properties of the 1st three alkali metals

Alkali metal	Appearance	Ease of cutting	Melting point (°C)	Boiling point (°C)	Conductivity	1st ionization energy
Lithium	Silvery white	Not easy	180	1330	Good	520
Sodium	Shiny grey	Easy	98	890	Good	496
Potassium	Shiny grey	Very easy	64	774	Good	419

Chemical properties

(i)Reaction with air/oxygen

On exposure to air, alkali metals react with the elements in the air.

Example

On exposure to air, Sodium first reacts with Oxygen to form sodium oxide.

$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$

The sodium oxide formed further reacts with water/moisture in the air to form sodium hydroxide solution.

$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$

Sodium hydroxide solution reacts with carbon (IV) oxide in the air to form sodium carbonate.

$2NaOH\ (aq) + CO_2\ (g) \rightarrow Na_2CO_3\ (g) + H_2O\ (l)$

(ii)Burning in air/oxygen

Lithium burns in air with a **crimson**/deep red flame to form Lithium oxide

$4Li\ (s) + O_2\ (g) \rightarrow 2Li_2O(s)$

Sodium burns in air with a **yellow** flame to form sodium oxide

$4Na\ (s) + O_2\ (g) \rightarrow 2Na_2O(s)$

Sodium burns in oxygen with a **yellow** flame to form sodium peroxide

$2Na\ (s) + O_2\ (g) \rightarrow Na_2O_2\ (s)$

Potassium burns in air with a **lilac/purple** flame to form potassium oxide

$4K\ (s) + O_2\ (g) \rightarrow 2K_2O\ (s)$

(iii) Reaction with water:

Experiment

Measure 500 cm3 of water into a beaker.

Put three drops of phenolphthalein indicator.

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Put about 0.5g of Lithium metal into the beaker.
Determine the pH of final product
Repeat the experiment using about 0.1 g of Sodium and Potassium.
Caution: Keep a distance
Observations

Alkali metal	Observations	Comparative speed/rate of the reaction
Lithium	-Metal floats in water -rapid effervescence/fizzing/bubbling -colourless gas produced (that extinguishes burning splint with explosion /"pop" sound) -resulting solution turn phenolphthalein indicator pink -pH of solution = 12/13/14	Moderately vigorous
Sodium	-Metal floats in water -very rapid effervescence /fizzing /bubbling -colourless gas produced (that extinguishes burning splint with explosion /"pop" sound) -resulting solution turn phenolphthalein indicator pink -pH of solution = 12/13/14	Very vigorous
Potassium	-Metal floats in water -explosive effervescence /fizzing /bubbling -colourless gas produced (that extinguishes burning splint with explosion /"pop" sound) -resulting solution turn phenolphthalein indicator pink -pH of solution = 12/13/14	Explosive/burst into flames

Explanation

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Alkali metals are less dense than water. They therefore float in water. They react with water to form a strongly alkaline solution of their hydroxides and producing hydrogen gas. The rate of this reaction increase down the group. i.e. Potassium is more reactive than sodium .Sodium is more reactive than Lithium.

The reactivity increases as electropositivity increases of the alkali increases. This is because as the atomic radius increases, the ease of donating/losing outer electron increases during chemical reactions.

Chemical equations

$2Li(s)$	+	$2H_2O(l)$	->	$2LiOH(aq)$	+	$H_2(g)$
$2Na(s)$	+	$2H_2O(l)$	->	$2NaOH(aq)$	+	$H_2(g)$
$2K(s)$	+	$2H_2O(l)$	->	$2KOH(aq)$	+	$H_2(g)$
$2Rb(s)$	+	$2H_2O(l)$	->	$2RbOH(aq)$	+	$H_2(g)$
$2Cs(s)$	+	$2H_2O(l)$	->	$2CsOH(aq)$	+	$H_2(g)$
$2Fr(s)$	+	$2H_2O(l)$	->	$2FrOH(aq)$	+	$H_2(g)$

Reactivity increase down the group

(iv) Reaction with chlorine:

Experiment

Cut about 0.5g of sodium into a deflagrating spoon with a lid cover. Introduce it on a Bunsen flame until it catches fire. Quickly and carefully lower it into a gas jar containing dry chlorine to cover the gas jar.

Repeat with about 0.5g of Lithium.

Caution: This experiment should be done in fume chamber because chlorine is poisonous /toxic.

Observation

Sodium metal continues to burn with a yellow flame forming white solid/fumes.

Lithium metal continues to burn with a crimson flame forming white solid / fumes.

Alkali metals react with chlorine gas to form the corresponding metal chlorides.

The reactivity increase as electropositivity increase down the group from Lithium to Francium. The ease of donating/losing the outer electrons increase as the atomic radius increase and the outer electron is less attracted to the nucleus.

Chemical equations

$2Li(s)$	+	$Cl_2(g)$	->	$2LiCl(s)$
$2Na(s)$	+	$Cl_2(g)$	->	$2NaCl(s)$
$2K(s)$	+	$Cl_2(g)$	->	$2KCl(s)$
$2Rb(s)$	+	$Cl_2(g)$	->	$2RbCl(s)$
$2Cs(s)$	+	$Cl_2(g)$	->	$2CsCl(s)$
$2Fr(s)$	+	$Cl_2(g)$	->	$2FrCl(s)$

Reactivity increase down the group

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The table below shows some compounds of the 1st three alkali metals

	Lithium	sodium	Potassium
Hydroxide	LiOH	NaOH	KOH
Oxide	Li_2O	Na_2O	K_2O
Sulphide	Li_2S	Na_2S	K_2S
Chloride	LiCl	NaCl	KCl
Carbonate	Li_2CO_3	Na_2CO_3	K_2CO_3
Nitrate(V)	$LiNO_3$	$NaNO_3$	KNO_3
Nitrate(III)	-	$NaNO_2$	KNO_2
Sulphate(VI)	Li_2SO_4	Na_2SO_4	K_2SO_4
Sulphate(IV)	-	Na_2SO_3	K_2SO_3
Hydrogen carbonate	-	$NaHCO_3$	$KHCO_3$
Hydrogen sulphate(VI)	-	$NaHSO_4$	$KHSO_4$
Hydrogen sulphate(IV)	-	$NaHSO_3$	$KHSO_3$
Phosphate	-	Na_3PO_4	K_3PO_4
Manganate(VI)	-	$NaMnO_4$	$KMnO_4$
Dichromate(VI)	-	$Na_2Cr_2O_7$	$K_2Cr_2O_7$
Chromate(VI)	-	Na_2CrO_4	K_2CrO_4

Some **uses** of alkali metals include:

(i)Sodium is used in making sodium cyanide for extracting gold from gold ore.
(ii)Sodium chloride is used in seasoning food.
(iii)Molten mixture of sodium and potassium is used as coolant in nuclear reactors.
(iv)Sodium is used in making sodium hydroxide used in making soapy and soapless detergents.
(v)Sodium is used as a reducing agent for the extraction of titanium from Titanium (IV) chloride.
(vi)Lithium is used in making special high strength glasses
(vii)Lithium compounds are used to make dry cells in mobile phones and computer laptops.

Group II elements: Alkaline earth metals

Group II elements are called **Alkaline earth metals** . The alkaline earth metals include:

Element	Symbol	Atomic number	Electron structure	Oxidation state	Valency
Beryllium	Be	4	2:**2**	Be^{2+}	2

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Magnesium	Mg	12	2:8:**2**	Mg^{2+}	2
Calcium	Ca	20	2:8:8:**2**	Ca^{2+}	2
Strontium	Sr	38	2:8:18:8:**2**	Sr^{2+}	2
Barium	Ba	56	2:8:18:18:8:**2**	Ba^{2+}	2
Radium	Ra	88	2:8:18:32:18:8:**2**	Ra^{2+}	2

All alkaline earth metal atoms have two electrons in the outer energy level. They therefore are **divalent**. They donate /lose the two outer electrons to have oxidation state M^{2+}

The number of energy levels increases down the group from Beryllium to Radium. The more the number of energy levels the bigger/larger the atomic size. e.g.

The atomic size/radius of Calcium is bigger/larger than that of Magnesium because Calcium has more/4 energy levels than Magnesium (3 energy levels).

Atomic radius and ionic radius of alkaline earth metals increase down the group as the number of energy levels increases.

The atomic radius of alkaline earth metals is bigger than the ionic radius. This is because they react by losing/donating the two outer electrons and hence lose the outer energy level.

Table showing the atomic and ionic radius of the 1st three alkaline earth metals

Element	Symbol	Atomic number	Atomic radius(nM)	Ionic radius(nM)
Beryllium	Be	4	0.089	0.031
Magnesium	Mg	12	0.136	0.065
Calcium	Ca	20	0.174	0.099

The atomic radius of Magnesium is 0.136nM .The ionic radius of Mg^{2+} is 0.065nM. This is because Magnesium reacts by donating/losing the two outer electrons and hence the outer energy level. The remaining electrons/energy levels experience more effective / greater nuclear attraction/pull towards the nucleus reducing the atomic radius.

Electropositivity

All alkaline earth metals are also electropositive like alkali metals. The electropositivity increase with increase in atomic radius/size. Calcium is more electropositive than Magnesium. This is because the effective nuclear attraction on outer electrons decreases with increase in atomic radius. The two outer electrons in calcium experience less nuclear attraction and can be lost/ donated easily/with ease because of the higher/bigger atomic radius.

Ionization energy

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For alkaline earth metals the 1st ionization energy decrease down the group as the atomic radius increase and the effective nuclear attraction on outer energy level electrons decrease. e.g. The 1st ionization energy of Magnesium is 900 $kJmole^{-1}$ while that of Calcium is 590 $kJmole^{-1}$ **.**This is because atomic radius increase and thus effective nuclear attraction on outer energy level electrons decrease down the group from magnesium to calcium.

It requires therefore less energy to donate/lose outer electron in calcium than in magnesium.

The minimum amount of energy required to remove a second electron from an ion of an element in its gaseous state is called the **2nd ionization energy**.

The 2nd ionization energy is always higher /bigger than the 1st ionization energy. This is because once an electron is donated /lost form an atom, the overall effective nuclear attraction on the remaining electrons/energy level increase. Removing a second electron from the ion require therefore more energy than the first electron.

The atomic radius of alkali metals is higher/bigger than that of alkaline earth metals. This is because across/along the period from left to right there is an increase in nuclear charge from additional number of protons and still additional number of electrons entering the same energy level. Increase in nuclear charge increases the effective nuclear attraction on the outer energy level which pulls it closer to the nucleus. e.g.

Atomic radius of Sodium (0.157nM) is higher than that of Magnesium (0.137nM). This is because Magnesium has more effective nuclear attraction on the outer energy level than Sodium hence pulls outer energy level more nearer to its nucleus.

Physical properties

Soft/Easy to cut: Alkaline earth metals are **not** soft and easy to cut with a knife like alkali metals. This is because of the decrease in atomic radius of corresponding alkaline earth metal, increases the strength of metallic bond and the packing of the metallic structure. Alkaline earth metals are:

(i) ductile(able to form wire/thin long rods)
(ii) malleable(able to be hammered into sheet/long thin plates)
(iii) have high tensile strength(able to be coiled without breaking/ not brittle/withstand stress)

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Appearance: Alkali earth metals have a shiny grey metallic luster when their surface is freshly polished /scrubbed. The surface slowly tarnishes on exposure to air. This is because the metal surface slowly undergoes oxidation to form an oxide. This oxide layer should be removed before using the alkaline earth metals.

Melting and boiling points: Alkaline earth metals have a relatively high melting/ boiling point than alkali metals. This is because alkali metals use only one delocalized electron to form a weaker metallic bond/structure. Alkaline earth metals use two delocalized electrons to form a stronger metallic bond /structure. The melting and boiling points decrease down the group as the atomic radius/size increase reducing the strength of metallic bond and packing of the metallic structure. e.g.

Beryllium has a melting point of 1280°C. Magnesium has a melting point of 650°C.Beryllium has a smaller atomic radius/size than magnesium .The strength of metallic bond and packing of the metallic structure is thus stronger in beryllium.

Electrical/thermal conductivity: Alkaline earth metals are good thermal and electrical conductors. The two delocalized valence electrons move randomly within the metallic structure.

Electrical conductivity increase down the group as the atomic radius/size increase making the delocalized outer electrons less attracted to nucleus. Alkaline earth metals are better thermal and electrical conductors than alkali metals because they have more/two outer delocalized electrons. e.g.

Magnesium is a better conductor than sodium because it has more/two delocalized electrons than sodium. The more delocalized electrons the better the electrical conductor.

Calcium is a better conductor than magnesium.

Calcium has bigger/larger atomic radius than magnesium because the delocalized electrons are less attracted to the nucleus of calcium and thus more free /mobile and thus better the electrical conductor

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Summary of some physical properties of the 1st three alkaline earth metals

Alkaline earth metal	Appearance	Ease of cutting	Melting point (^{o}C)	Boiling point (^{o}C)	Conduct-ivity	1st ionization energy	2nd ionization energy
Berylliu m	Shiny grey	Not easy	1280	3450	Good	900	1800
Magnesi	Shiny	Not	650	1110	Good	736	1450

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um	grey	Easy					
calcium	Shiny grey	Not easy	850	1140	Good	590	970

Chemical properties

(i) Reaction with air/oxygen

On exposure to air, the surface of alkaline earth metals is slowly oxidized to its oxide on prolonged exposure to air.

Example

On exposure to air, the surface of magnesium ribbon is oxidized to form a thin film of Magnesium oxide

. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

(ii) Burning in air/oxygen

Experiment

Hold a about 2cm length of Magnesium ribbon on a Bunsen flame. Stop heating when it catches fire/start burning.

Caution: Do not look directly at the flame

Put the products of burning into 100cm3 beaker. Add about 5cm3 of distilled water. Swirl. Test the mixture using litmus papers.

Repeat with Calcium

Observations

-Magnesium burns with a bright blindening flame
-White solid /ash produced
-Solid dissolves in water to form a colourless solution
-Blue litmus paper remain blue
-Red litmus paper turns blue
-colourless gas with pungent smell of urine

Explanation

Magnesium burns in air with a bright blindening flame to form a mixture of Magnesium oxide and Magnesium nitride.

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

$3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$

Magnesium oxide dissolves in water to form magnesium hydroxide.

$MgO(s) + H_2O(l) \rightarrow Mg(OH)_2(aq)$

Magnesium nitride dissolves in water to form magnesium hydroxide and produce ammonia gas.

$Mg_3N_2(s) + 6H_2O(l) \rightarrow 3Mg(OH)_2(aq) + 2NH_3(g)$

Magnesium hydroxide and ammonia are weakly alkaline with pH 8/9/10/11 and turns red litmus paper blue.

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Calcium burns in air with faint orange/red flame to form a mixture of both Calcium oxide and calcium nitride.

$2Ca\ (s) + O_2\ (g) \rightarrow 2CaO(s)$

$3Ca\ (s) + N_2\ (g) \rightarrow Ca_3N_2\ (s)$

Calcium oxide dissolves in water to form calcium hydroxide.

$CaO(s) + H_2O\ (l) \rightarrow Ca\ (OH)_2(aq)$

Calcium nitride dissolves in water to form calcium hydroxide and produce ammonia gas.

$Ca_3N_2\ (s) + 6H_2O\ (l) \rightarrow 3Ca\ (OH)_2(aq) + 2NH_3\ (g)$

Calcium hydroxide is also weakly alkaline solution with pH 8/9/10/11 and turns red litmus paper blue.

(ii) Reaction with water

Experiment

Measure 50 cm3 of distilled water into a beaker.
Scrub/polish with sand paper 1cm length of Magnesium ribbon
Place it in the water. Test the product-mixture with blue and red litmus papers.
Repeat with Calcium metal.

Observations

-Surface of magnesium covered by bubbles of colourless gas.
-Colourless solution formed.
-Effervescence/bubbles/fizzing takes place in Calcium.
-Red litmus paper turns blue.
-Blue litmus paper remains blue.

Explanations

Magnesium slowly reacts with cold water to form Magnesium hydroxide and bubbles of Hydrogen gas that stick on the surface of the ribbon.

$Mg(s) + 2H_2O\ (l) \rightarrow Mg(OH)_2(aq) + H_2\ (g)$

Calcium moderately reacts with cold water to form Calcium hydroxide and produce a steady stream of Hydrogen gas.

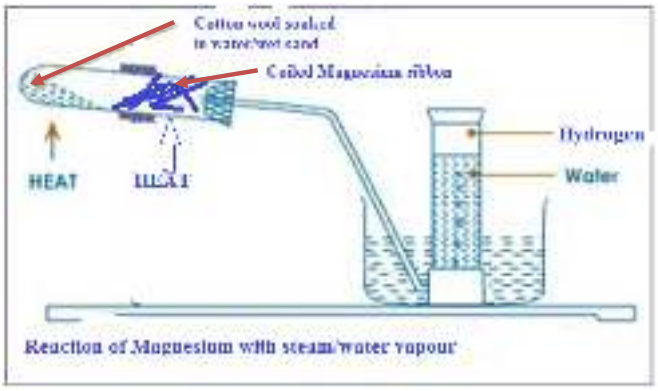
$Ca(s) + 2H2O\ (l) \rightarrow Ca\ (OH)_2(aq) + H_2\ (g)$

(iii) Reaction with water vapour/steam

Experiment
Put some cotton wool soaked in water/wet sand in a long boiling tube.
Coil a well polished magnesium ribbon into the boiling tube.
Ensure the coil touches the side of the boiling tube. Heat the cotton wool/sand slightly then strongly heat the Magnesium ribbon.
Set up of apparatus

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Observations

-Magnesium glows red hot then burns with a blindening flame.
-Magnesium continues to glow/burning even without more heating.
-White solid/residue.
-colourless gas collected over water.

Explanation

On heating wet sand, steam is generated which drives out the air that would otherwise react with /oxidize the ribbon.

Magnesium burns in steam/water vapour generating enough heat that ensures the reaction goes to completion even without further heating. White Magnesium oxide is formed and hydrogen gas is evolved.

To prevent suck back, the delivery tube should be removed from the water before heating is stopped at the end of the experiment.

$Mg(s) + H_2O(l) \rightarrow MgO(s) + H_2(g)$

(iv) Reaction with chlorine gas.

Experiment

Lower slowly burning magnesium ribbon/shavings into a gas jar containing Chlorine gas. Repeat with a hot piece of calcium metal.

Observation

-Magnesium continues to burn in chlorine with a bright blindening flame.
-Calcium continues to burn for a short time.
-White solid formed.
-Pale green colour of chlorine fades.

Explanation

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Magnesium continues to burn in chlorine gas forming white magnesium oxide solid.

$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$

Calcium burns slightly in chlorine gas to form white calcium oxide solid. Calcium oxide formed coat unreacted Calcium stopping further reaction

$Ca(s) + Cl_2(g) \rightarrow CaCl_2(s)$

(v) Reaction with dilute acids.

Experiment

Place about 4.0cm3 of 0.1M dilute sulphuric (VI) acid into a test tube. Add about 1.0cm length of magnesium ribbon into the test tube. Cover the mouth of the test tube using a thumb. Release the gas and test the gas using a burning splint.
Repeat with about 4.0cm3 of 0.1M dilute hydrochloric/nitric (V) acid.
Repeat with 0.1g of Calcium in a beaker with all the above acid

Caution: Keep distance when using calcium

Observation

-Effervescence/fizzing/bubbles with dilute sulphuric (VI) and nitric (V) acids
-Little Effervescence/fizzing/bubbles with calcium and dilute sulphuric (VI) acid.
-Colourless gas produced that extinguishes a burning splint with an explosion/ "pop" sound.
-No gas is produced with Nitric (V) acid.
-Colourless solution is formed.

Explanation

Dilute acids react with alkaline earth metals to form a salt and produce hydrogen gas.
Nitric (V) acid is a strong oxidizing agent. It quickly oxidizes the hydrogen produced to water.
Calcium is very reactive with dilute acids and thus a very **small** piece of very **dilute** acid should be used.

Chemical equations

$Mg(s) + H2SO4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

$Mg(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$

$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

$Ca(s) + H_2SO_4(aq) \rightarrow CaSO_4(\mathbf{s}) + H_2(g)$
(insoluble $CaSO_4(\mathbf{s})$ coat/cover Ca(s))

$Ca(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2(g)$

$Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$

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$Ba(s) + H_2SO_4(aq) \rightarrow BaSO_4(\mathbf{s}) + H_2(g)$
(insoluble $BaSO_4$(**s**) coat/cover Ba(s))
$Ba(s) + 2HNO_3(aq) \rightarrow Ba(NO_3)_2(aq) + H_2(g)$
$Ba(s) + 2HCl(aq) \rightarrow BaCl_2(aq) + H_2(g)$

The table below shows some compounds of some alkaline earth metals

	Beryllium	Magnesium	Calcium	Barium
Hydroxide	$Be(OH)_2$	$Mg(OH)_2$	$Ca(OH)_2$	$Ba(OH)_2$
Oxide	BeO	MgO	CaO	BaO
Sulphide	-	MgS	CaS	BaS
Chloride	$BeCl_2$	$MgCl_2$	$CaCl_2$	$BaCl_2$
Carbonate	$BeCO_3$	$MgCO_3$	$CaCO_3$	$BaCO_3$
Nitrate(V)	$Be(NO_3)_2$	$Mg(NO_3)_2$	$Ca(NO_3)_2$	$Ba(NO_3)_2$
Sulphate(VI)	$BeSO_4$	$MgSO_4$	$CaSO_4$	$BaSO_4$
Sulphate(IV)	-	-	$CaSO_3$	$BaSO_3$
Hydrogen carbonate	-	$Mg(HCO_3)_2$	$Ca(HCO_3)_2$	-
Hydrogen sulphate(VI)	-	$Mg(HSO_4)_2$	$Ca(HSO_4)_2$	-

Some uses of alkaline earth metals include:

(i)Magnesium hydroxide is a non-toxic/poisonous mild base used as an anti acid medicine to relieve stomach acidity.
(ii) Making **duralumin.** Duralumin is an alloy of Magnesium and aluminium used for making aeroplane bodies because it is light.
(iii) Making plaster of Paris-Calcium sulphate (VI) is used in hospitals to set a fractures bone.
(iii)Making cement-Calcium carbonate is mixed with clay and sand then heated to form cement for construction/building.
(iv)Raise soil pH-Quicklime/calcium oxide is added to acidic soils to neutralize and raise the soil pH in agricultural farms.
(v) As nitrogenous fertilizer-Calcium nitrate (V) is used as an agricultural fertilizer because plants require calcium for proper growth.
(vi)In the blast furnace-Limestone is added to the blast furnace to produce more reducing agent and remove slag in the blast furnace for extraction of Iron.

(c)Group VII elements: Halogens

Group VII elements are called **Halogens.** They are all non metals. They include:

Element	**Symbol**	**Atomic number**	**Electronicc configuration**	**Charge of ion**	**Valency**	**State at Room Temperature**

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Fluorine	**F**	**9**	**2:7**	**F^-**	**1**	**Pale yellow gas**
Chlorine	**Cl**	**17**	**2:8:7**	**Cl^-**	**1**	**Pale green gas**
Bromine	**Br**	**35**	**2:8:18:7**	**Br^-**	**1**	**Red liquid**
Iodine	**I**	**53**	**2:8:18:18:7**	**I^-**	**1**	**Grey Solid**
Astatine	**At**	**85**	**2:8:18:32:18:7**	**At^-**	**1**	**Radioactive**

All halogen atoms have seven electrons in the outer energy level. They **acquire/gain one** electron in the outer energy level to be stable. They therefore are therefore **monovalent** .They exist in oxidation state X^-

The number of energy levels increases down the group from Fluorine to Astatine. The more the number of energy levels the bigger/larger the atomic size. e.g.

The atomic size/radius of Chlorine is bigger/larger than that of Fluorine because Chlorine has more/3 energy levels than Fluorine (2 energy levels).

Atomic radius and ionic radius of Halogens increase down the group as the number of energy levels increases.

The atomic radius of Halogens is smaller than the ionic radius. This is because they react by gaining/acquiring extra one electron in the outer energy level. The effective nuclear attraction on the more/extra electrons decreases. The incoming extra electron is also repelled causing the outer energy level to expand to reduce the repulsion and accommodate more electrons.

Table showing the atomic and ionic radius of four Halogens

Element	Symbol	Atomic number	Atomic radius(nM)	Ionic radius(nM)
Fluorine	F	9	0.064	0.136
Chlorine	Cl	17	0.099	0.181
Bromine	Br	35	0.114	0.195
Iodine	I	53	0.133	0.216

The atomic radius of Chlorine is 0.099nM .The ionic radius of Cl^- is 0.181nM. This is because Chlorine atom/molecule reacts by gaining/acquiring extra one electron. The more/extra electrons/energy level experience less effective nuclear attraction /pull towards the nucleus. The outer energy level expand/increase to reduce the repulsion of the existing and incoming gained /acquired electrons.

Electronegativity

The ease of gaining/acquiring extra electrons is called electronegativity. All halogens are electronegative. Electronegativity decreases as atomic radius increase. This is because the effective nuclear attraction on outer electrons decreases with increase in atomic radius.

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The outer electrons experience less nuclear attraction and thus ease of gaining/acquiring extra electrons decrease.
It is measured using Pauling's scale.
Where Fluorine with Pauling scale 4.0 is the most electronegative element and thus the highest tendency to acquire/gain extra electron.

Table showing the electronegativity of the halogens

Halogen	F	Cl	Br	I	At
Electronegativity (Pauling scale)	4.0	3.0	2.8	2.5	2.2

The electronegativity of the halogens decrease down the group from fluorine to Astatine. This is because atomic radius increases down the group and thus decrease electron – attracting power down the group from fluorine to astatine.
Fluorine is the most electronegative element in the periodic table because it has the small atomic radius.

Electron affinity

The minimum amount of energy required to gain/acquire an extra electron by an atom of element in its gaseous state is called 1st **electron affinity**. The SI unit of electron affinity is **kilojoules per mole/kJmole^{-1}.** Electron affinity depends on atomic radius. The higher the atomic radius, the less effective the nuclear attraction on outer energy level electrons and thus the lower the electron affinity.
For halogens the 1st electron affinity decrease down the group as the atomic radius increase and the effective nuclear attraction on outer energy level electrons decrease. Due to its small size/atomic radius Fluorine shows exceptionally low electron affinity. This is because a lot of energy is required to overcome the high repulsion of the existing and incoming electrons.
Table showing the election affinity of halogens for the process

$X + e \rightarrow X^-$

Halogen	F	Cl	Br	I
Electron affinity kJmole^{-1}	-333	-364	-342	-295

The higher the electron affinity the more stable theion.i.e
Cl^- is a more stable ion than Br^- because it has a more negative / exothermic electron affinity than Br^-
Electron affinity is different from:

(i) Ionization energy.

Ionization energy is the energy required to lose/donate an electron in an atom of an element in its gaseous state while electron affinity is the energy required to gain/acquire extra electron by an atom of an element in its gaseous state.

(ii) Electronegativity.

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-Electron affinity is the energy required to gain an electron in an atom of an element in gaseous state. It involves the process:

$$X(g) + e \rightarrow X^-(g)$$

Electronegativity is the ease/tendency of gaining/ acquiring electrons by an element during chemical reactions.

It does not involve use of energy but theoretical arbitrary Pauling' scale of measurements.

Physical properties

State at room temperature

Fluorine and Chlorine are gases, Bromine is a liquid and Iodine is a solid. Astatine is radioactive.

All halogens exist as **diatomic** molecules bonded by strong covalent bond. Each molecule is joined to the other by weak intermolecular forces/ Van-der-waals forces.

Melting/Boiling point

The strength of intermolecular/Van-der-waals forces of attraction increase with increase in molecular size/atomic radius.

Iodine has therefore the largest atomic radius and thus strongest intermolecular forces to make it a solid.

Iodine sublimes when heated to form (**caution:** highly toxic/poisonous) purple vapour.

This is because Iodine molecules are held together by weak van-der-waals/intermolecular forces which require little heat energy to break.

Electrical conductivity

All Halogens are poor conductors of electricity because they have no free delocalized electrons.

Solubility in polar and non-polar solvents

All halogens are soluble in water (polar solvent).

When a boiling tube containing either chlorine gas or bromine vapour is separately inverted in a beaker containing distilled water and tetrachloromethane (non-polar solvent), the level of solution in boiling tube rises in both water and tetrachloromethane.

This is because halogen are soluble in both polar and non-polar solvents.

Solubility of halogens in water/polar solvents decrease down the group. Solubility of halogens in non-polar solvent increases down the group.

The level of water in chlorine is higher than in bromine and the level of tetrachloromethane in chlorine is lower than in bromine.

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Caution: Tetrachloromethane , Bromine vapour and Chlorine gas are all **highly** toxic/poisonous.

Table showing the physical properties of Halogens

Halogen	**Formula of molecule**	**Electrical conductivity**	**Solubility in water**	**Melting point(°C)**	**Boiling point(°C)**
Fluorine	F_2	Poor	Insoluble/soluble in tetrachloromethane	-238	-188
Chlorine	Cl_2	Poor	Insoluble/soluble in tetrachloromethane	-101	-35
Bromine	Br_2	Poor	Insoluble/soluble in tetrachloromethane	7	59
Iodine	I_2	Poor	Insoluble/soluble in tetrachloromethane	114	sublimes

Chemical properties

(i)Displacement

Experiment

Place separately in test tubes about $5cm^3$ of sodium chloride, Sodium bromide and Sodium iodide solutions.

Add 5 drops of chlorine water to each test tube:

Repeat with 5 drops of bromine water instead of chlorine water

Observation

Using Chlorine water

-Yellow colour of chlorine water fades in all test tubes except with sodium chloride.

-Coloured Solution formed.

Using Bromine water

-Yellow colour of bromine water fades in test tubes containing sodium iodide.

-Coloured Solution formed.

Explanation

The halogens displace each other from their solution. The more electronegative displace the less electronegative from their solution.

Chlorine is more electronegative than bromine and iodine.

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On adding chlorine water, bromine and Iodine are displaced from their solutions by chlorine.
Bromine is more electronegative than iodide but less 6than chlorine.
On adding Bromine water, iodine is displaced from its solution but not chlorine.

Table showing the displacement of the halogens

(V) means there is displacement (x) means there is no displacement

Halogen ion in solution / Halogen	F^-	Cl^-	Br^-	I^-
F_2	X			
Cl_2	X	X		
Br_2	X	X	X	
I_2	X	X	X	X

Chemical /ionic equations

With Fluorine

$F_2(g) + 2NaCl^-(aq) \rightarrow 2NaF(aq) + Cl_2(aq)$
$F_2(g) + 2Cl^-(aq) \rightarrow 2F^-(aq) + Cl_2(aq)$
$F_2(g) + 2NaBr^-(aq) \rightarrow 2NaF(aq) + Br_2(aq)$
$F_2(g) + 2Br^-(aq) \rightarrow 2F^-(aq) + Br_2(aq)$
$F_2(g) + 2NaI^-(aq) \rightarrow 2NaF(aq) + I_2(aq)$
$F_2(g) + 2I^-(aq) \rightarrow 2F^-(aq) + I_2(aq)$

With chlorine

$Cl_2(g) + 2NaCl^-(aq) \rightarrow 2NaCl(aq) + Br_2(aq)$
$Cl_2(g) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)$
$Cl_2(g) + 2NaI^-(aq) \rightarrow 2NaCl(aq) + I_2(aq)$
$Cl_2(g) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$

With Bromine

$Br_2(g) + 2NaI^-(aq) \rightarrow 2NaBr(aq) + I_2(aq)$
$Br_2(g) + 2I^-(aq) \rightarrow 2Br^-(aq) + I_2(aq)$

Uses of halogens

(i) Florine – manufacture of P.T.F.E (Poly tetra fluoroethene) synthetic fiber.
- Reduce tooth decay when added in small amounts/quantities in tooth paste.

NB –large small quantities of fluorine /fluoride ions in water cause browning of teeth/flourosis.

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- Hydrogen fluoride is used to engrave words /pictures in glass.

(ii) Bromine - Silver bromide is used to make light sensitive photographic paper/films.

(iii) Iodide – Iodine dissolved in alcohol is used as medicine to kill bacteria in skin cuts. It is called tincture of iodine.

The table below to show some compounds of halogens

Element / Halogen	H	Na	Mg	Al	Si	C	P
F	**HF**	**NaF**	**MgH_2**	**AlF_3**	**SiF_4**	**CF_4**	**PF_3**
Cl	**HCl**	**NaCl**	**MgCl**	**$AlCl_3$**	**$SiCl_3$**	**CCl_4**	**PCl_3**
Br	**HBr**	**NaBr**	**$MgBr_2$**	**$AlBr_3$**	**$SiBr_4$**	**CBr_4**	**PBr_3**
I	**HI**	**NaI**	**MgI_2**	**AlI_3**	**SiI_4**	**I_2**	**PBr_3**

(i) Below is the table showing the bond energy of four halogens.

Bond	Bond energy k J $mole^{-1}$
Cl-Cl	242
Br-Br	193
I-I	151

I. What do you understand by the term “bond energy”

Bond energy is the energy required to break/ form one mole of chemical bond

II. Explain the trend in bond Energy of the halogens above:

-**Decrease down the group from chlorine to Iodine**

-Atomic radius increase down the group decreasing the energy required to break the covalent bonds between the larger atom with reduced effective nuclear @ charge an outer energy level that take part in bonding.

(c)Group VIII elements: Noble gases

Group VIII elements are called **Noble gases.** They are all non metals. Noble gases occupy about 1.0% of the atmosphere as colourless gaseous mixture. Argon is the most abundant with 0.9%.

They exist as **monatomic** molecules with very weak van-der-waals /intermolecular forces holding the molecules.

They include:

Element	Symbol	Atomic number	Electron structure	State at room temperature

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Helium	He	2	2:	Colourless gas
Neon	Ne	10	2:8	Colourless gas
Argon	Ar	18	2:8:8	Colourless gas
Krypton	Kr	36	2:8:18:8	Colourless gas
Xenon	Xe	54	2:8:18:18:8	Colourless gas
Radon	Rn	86	2:8:18:32:18:8	Radioctive

All noble gas atoms have a stable duplet(two electrons in the 1st energy level) or octet(eight electrons in other outer energy level)in the outer energy level. They therefore **do not** acquire/gain extra electron in the outer energy level or donate/lose. They therefore are therefore **zerovalent** .

The number of energy levels increases down the group from Helium to Randon. The more the number of energy levels the bigger/larger the atomic size/radius.
e.g.

The atomic size/radius of Argon is bigger/larger than that of Neon because Argon has more/3 energy levels than Neon (2 energy levels).

Atomic radius noble gases increase down the group as the number of energy levels increases.

The effective nuclear attraction on the outer electrons thus decrease down the group.

The noble gases are generally unreactive because the outer energy level has the stable octet/duplet. The stable octet/duplet in noble gas atoms lead to comparatively very high 1st ionization energy. This is because losing /donating an electron from the stable atom require a lot of energy to lose/donate and make it unstable.

As atomic radius increase down the group and the 1st ionization energy decrease, very electronegative elements like Oxygen and Fluorine are able to react and bond with lower members of the noble gases.e.g

Xenon reacts with Fluorine to form a covalent compound XeF_6.This is because the outer electrons/energy level if Xenon is far from the nucleus and thus experience less effective nuclear attraction.

Noble gases have low melting and boiling points. This is because they exist as monatomic molecules joined by very weak intermolecular/van-der-waals forces that require very little energy to weaken and form liquid and break to form a gas. The intermolecular/van-der-waals forces increase down the group as the atomic radius/size increase from Helium to Radon. The melting and boiling points thus increase also down the group.

Noble gases are insoluble in water and are poor conductors of electricity.

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Element	Formula of molecule	Electrical conductivity	Solubility in water	Atomic radius(nM)	1st ionization energy	Melting point(^{0}C)	Boiling point(^{0}C)
Helium	He	Poor	Insoluble	0.128	2372	-270	-269
Neon	Ne	Poor	Insoluble	0.160	2080	-249	-246
Argon	Ar	Poor	Insoluble	0.192	1520	-189	-186
Krypton	Kr	Poor	Insoluble	0.197	1350	-157	-152
Xenon	Xe	Poor	Insoluble	0.217	1170	-112	-108
Radon	Rn	Poor	Insoluble	0.221	1134	-104	-93

Uses of noble gases

Argon is used in light bulbs to provide an inert environment to prevent oxidation of the bulb filament

Argon is used in arch welding as an insulator.

Neon is used in street and advertisement light

Helium is mixed with Oxygen during deep sea diving and mountaineering.

Helium is used in weather balloon for meteorological research instead of Hydrogen because it is unreactive/inert.Hydrogen when impure can ignite with an explosion.

Helium is used in making thermometers for measuring very low temperatures.

CHEMISTRY OF CARBON

A: CARBON

Carbon is an element in Group IV(Group 4)of the Periodic table .It has atomic number 6 and electronic configuration 2:4 and thus has four valence electrons(tetravalent).It does not easily ionize but forms strong covalent bonds with other elements including itself.

(a)Occurrence

Carbon mainly naturally occurs as:

(i)allotropes of carbon i.e graphite, diamond and fullerenes.

(ii)amorphous carbon in coal, peat ,charcoal and coke.

(iii)carbon(IV)oxide gas accounting 0.03% by volume of normal air in the atmosphere.

(b)Allotropes of Carbon

Carbon naturally occur in two main crystalline allotropic forms, carbon-graphite and carbon-diamond

Carbon-diamond	Carbon-graphite
Shiny crystalline solid	Black/dull crystalline solid
Has a very high melting/boiling point because it has a very closely packed	Has a high melting/boiling point because it has a very closely packed

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giant tetrahedral structure joined by strong covalent bonds	giant hexagonal planar structure joined by strong covalent bonds
Has very high density(Hardest known natural substance)	Soft
Abrassive	Slippery
Poor electrical conductor because it has no free delocalized electrons	Good electrical conductor because it has free 4th valency delocalized electrons
Is used in making Jewels, drilling and cutting metals	Used in making Lead-pencils,electrodes in batteries and as a lubricant
Has giant tetrahedral structure	Has giant hexagonal planar structure

c) Properties of Carbon

(i)Physical properties of carbon

Carbon occur widely and naturally as a black solid

It is insoluble in water but soluble in carbon disulphide and organic solvents.

It is a poor electrical and thermal conductor.

(ii)Chemical properties of carbon

I. Burning

Experiment

Introduce a small piece of charcoal on a Bunsen flame then lower it into a gas jar containing Oxygen gas. Put three drops of water. Swirl. Test the solution with blue and red litmus papers.

Observation

-Carbon chars then burns with a blue flame

-Colourless and odourless gas produced

-Solution formed turn blue litmus paper faint red. Red litmus paper remains red.

Explanation

Carbon burns in air and faster in Oxygen with a blue non-sooty/non-smoky flame forming Carbon (IV) oxide gas. Carbon burns in limited supply of air with a blue non-sooty/non-smoky flame forming Carbon (IV) oxide gas. Carbon (IV) oxide gas dissolve in water to form weak acidic solution of Carbonic (IV)acid.

Chemical Equation

$C(s) + O_2(g) \rightarrow CO_2(g)$ (in excess air)

$2C(s) + O_2(g) \rightarrow 2CO(g)$ (in limited air)

$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$ (very weak acid)

II. Reducing agent

Experiment

Mix thoroughly equal amounts of powdered charcoal and copper (II)oxide into a crucible. Heat strongly.

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Observation
Colour change from black to brown
Explanation
Carbon is a reducing agent. For ages it has been used to reducing metal oxide ores to metal, itself oxidized to carbon (IV) oxide gas. Carbon reduces black copper (II) oxide to brown copper metal

Chemical Equation

$2CuO(s) + C(s) \rightarrow 2Cu(s) + CO_2(g)$
(black) (brown)

$2PbO(s) + C(s) \rightarrow 2Pb(s) + CO_2(g)$
(brown when hot/ yellow when cool) (grey)

$2ZnO(s) + C(s) \rightarrow 2Zn(s) + CO_2(g)$
(yellow when hot/ white when cool) (grey)

$Fe_2O_3(s) + 3C(s) \rightarrow 2Fe(s) + 3CO_2(g)$
(brown when hot/cool) (grey)

$Fe_3O_4(s) + 4C(s) \rightarrow 3Fe(s) + 4CO_2(g)$
(brown when hot/cool) (grey)

B: COMPOUNDS OF CARBON

The following are the main compounds of Carbon

(i)Carbon(IV)Oxide(CO_2)
(ii)Carbon(II)Oxide(CO)
(iii)Carbonate(IV) (CO_3^{2-})and hydrogen carbonate(IV(HCO_3^-)
(iv)Sodium carbonate(Na_2CO_3)
(i) Carbon(IV)Oxide (CO_2)

(a)Occurrence

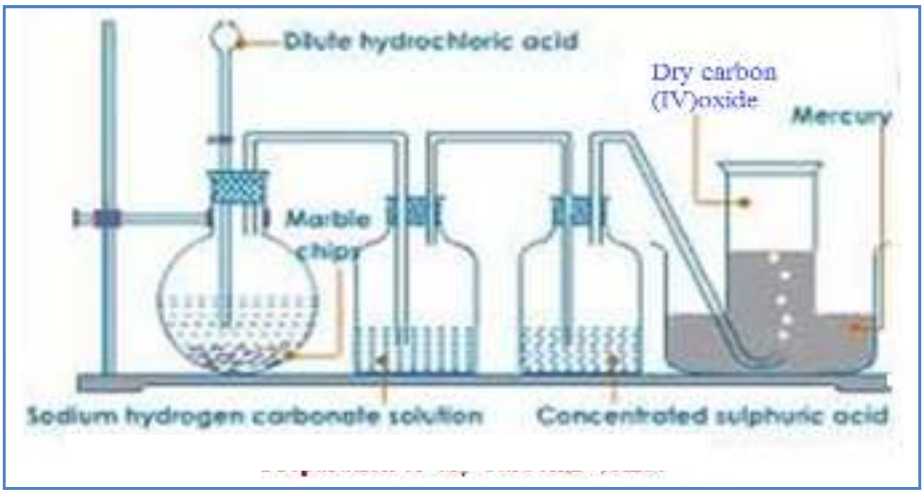
Carbon(IV)oxide is found:
-in the air /atmosphere as 0.03% by volume.
-a solid carbon(IV)oxide mineral in Esageri near Eldame Ravine and Kerita near Limuru in Kenya.

(b)School Laboratory preparation

In the school laboratory carbon(IV)oxide can be prepared in the school laboratory from the reaction of marble chips($CaCO_3$)or sodium hydrogen carbonate($NaHCO_3$) with dilute hydrochloric acid.

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(c)Properties of carbon(IV)oxide gas(Questions)

1.Write the equation for the reaction for the school laboratory preparation of carbon (IV)oxide gas.

Any carbonate reacted with dilute hydrochloric acid should be able to generate carbon (IV)oxide gas.

Chemical equations

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCO_3(aq) + H_2O(l) + CO_2(g)$

$ZnCO_3(s) + 2HCl(aq) \rightarrow ZnCO_3(aq) + H_2O(l) + CO_2(g)$

$MgCO_3(s) + 2HCl(aq) \rightarrow MgCO_3(aq) + H_2O(l) + CO_2(g)$

$CuCO_3(s) + 2HCl(aq) \rightarrow CuCO_3(aq) + H_2O(l) + CO_2(g)$

$NaHCO_3(s) + HCl(aq) \rightarrow Na_2CO_3(aq) + H_2O(l) + CO_2(g)$

$KHCO_3(s) + HCl(aq) \rightarrow K_2CO_3(aq) + H_2O(l) + CO_2(g)$

2.What method of gas collection is used in preparation of Carbon(IV)oxide gas. Explain.

Downward delivery /upward displacement of air/over mercury

Carbon(IV)oxide gas is about 1½ times denser than air.

3.What is the purpose of :

(a)water?

To absorb the more volatile hydrogen chloride fumes produced during the vigorous reaction.

(b)sodium hydrogen carbonate?

To absorb the more volatile hydrogen chloride fumes produced during the vigorous reaction and by reacting with the acid to produce more carbon (IV)oxide gas .

Chemical equation

$NaHCO_3(s) + HCl(aq) \rightarrow Na_2CO_3(aq) + H_2O(l) + CO_2(g)$

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(c)concentrated sulphuric(VI)acid?

To dry the gas/as a drying agent

4.Describe the smell of carbon(IV)oxide gas

Colourless and odourless

5. Effect on lime water.

Experiment

Bubbled carbon(IV)oxide gas into a test tube containing lime water for about three minutes

Observation

White precipitate is formed.

White precipitate dissolved when excess carbon(IV)oxide gas is bubbled .

Explanation

Carbon(IV)oxide gas reacts with lime water($Ca(OH)_2$) to form an insoluble white precipitate of calcium carbonate. Calcium carbonate reacts with more Carbon(IV) oxide gas to form soluble Calcium hydrogen carbonate.

Chemical equation

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

$CaCO_3(aq) + H_2O(l) + CO_2(g) \rightarrow Ca(HCO_3)_2(aq)$

6. Effects on burning Magnesium ribbon

Experiment

Lower a piece of burning magnesium ribbon into a gas jar containing carbon (IV)oxide gas.

Observation

The ribbon continues to burn with difficulty

White ash/solid is formed.

Black speck/solid/particles formed on the side of gas jar.

Explanation

Carbon (IV) oxide gas does not support combustion/burning.Magnesium burn to produce/release enough heat energy to decompose Carbon (IV) oxide gas to carbon and oxygen.Magnesium continues to burn in Oxygen forming white Magnesium Oxide solid/ash.Black speck/particle of carbon/charcoal residue forms on the sides of reaction flask. During the reaction Carbon (IV) oxide is reduced (Oxidizing agent)to carbon while Magnesium is Oxidized to Magnesium Oxide.

Chemical equation

$2Mg(s) + CO_2(g) \rightarrow C(s) + 2MgO(l)$

7. Dry and wet litmus papers were separately put in a gas jar containing dry carbon (IV) oxide gas. State and explain the observations made.

Observation

Blue dry litmus paper remain blue

Red dry litmus paper remain Red

Blue wet/damp/moist litmus paper turns red

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Red wet/damp/moist litmus paper remain red
Explanation
Dry Carbon (IV) oxide gas is a molecular compound that does not dissociate/ionize to release $\mathbf{H^+}$ and thus has no effect on litmus papers.
Wet/damp/moist litmus paper contains water that dissolves/react with dry carbon (IV) oxide gas to form the weak solution of carbonic (IV) acid (H_2CO_3).
Carbonic (IV) acid dissociate/ionizes to a few /little free $\mathbf{H^+}$ and $\mathbf{CO_3^{2-}}$.
The few H^+ (aq) ions are responsible for turning blue litmus paper to faint red showing the gas is very weakly acidic.
Chemical equation

$$H_2CO_3(aq) \rightarrow 2H^+(aq) + CO_3^{2-}(aq)$$

8. Explain why Carbon (IV) oxide cannot be prepared from the reaction of:

(i) Marble chips with dilute sulphuric (VI) acid.

Explanation
Reaction forms insoluble calcium sulphate (VI) that cover/coat unreacted marble chips stopping further reaction
Chemical equation

$$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(\mathbf{s}) + H_2O(l) + CO_2(g)$$
$$PbCO_3(s) + H_2SO_4(aq) \rightarrow PbSO_4(\mathbf{s}) + H_2O(l) + CO_2(g)$$
$$BaCO_3(s) + H_2SO_4(aq) \rightarrow BaSO_4(\mathbf{s}) + H_2O(l) + CO_2(g)$$

(ii) Lead (II) carbonate with dilute Hydrochloric acid.

Reaction forms insoluble Lead (II)Chloride that cover/coat unreacted Lead(II) carbonate stopping further reaction unless the reaction mixture is heated. Lead (II) Chloride is soluble in hot water.
Chemical equation

$$PbCO_3(s) + 2HCl(aq) \rightarrow PbCl_2(\mathbf{s}) + H_2O(l) + CO_2(g)$$

9. Describe the test for the presence of Carbon (IV) oxide.

Using burning splint
Lower a burning splint into a gas jar suspected to contain Carbon (IV) oxide gas. The burning splint is extinguished.
Using Lime water.
Bubble the gas suspected to be Carbon (IV) oxide gas. A white precipitate that dissolves in excess bubbling is formed.
Chemical equation

$$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$$
$$CaCO_3(aq) + H_2O(l) + CO_2(g) \rightarrow Ca(HCO_3)_2(aq)$$

10.State three main uses of Carbon (IV)oxide gas

(i)In the Solvay process for the manufacture of soda ash/sodium carbonate
(ii)In preservation of aerated drinks
(iii)As fire extinguisher because it does not support combustion and is denser than air.

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(iv)In manufacture of Baking powder.

(ii) Carbon (II) Oxide (CO)

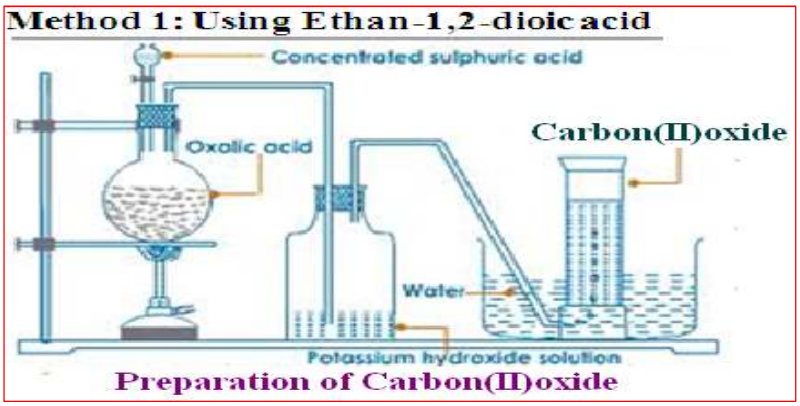
(a)Occurrence

Carbon (II) oxide is found is found from incomplete combustion of fuels like petrol charcoal, liquefied Petroleum Gas/LPG.

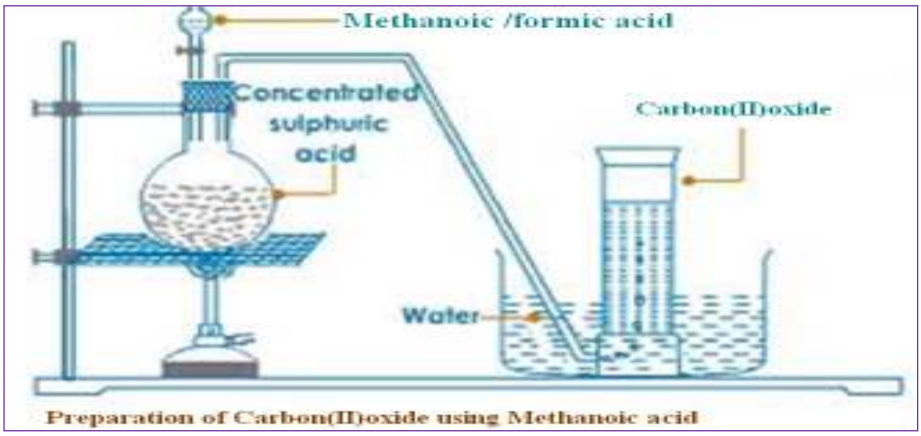
(b)School Laboratory preparation

In the school laboratory carbon(II)oxide can be prepared from dehydration of methanoic acid/Formic acid(HCOOH) or Ethan-1,2-dioic acid/Oxalic acid(HOOCCOOH) using concentrated sulphuric(VI) acid. Heating is necessary.

METHOD 1:Preparation of Carbon (IV)Oxide from dehydration of Oxalic/ethan-1,2-dioic acid



METHOD 2:Preparation of Carbon (IV)Oxide from dehydration of Formic/Methanoic acid



(c)Properties of Carbon (II) Oxide(Questions)

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1. Write the equation for the reaction for the preparation of carbon(II)oxide using;

(i)Method 1;
Chemical equation
$HOOCCOOH(s) \xrightarrow{-Conc.H_2SO_4} CO(g) + CO_2(g) + H_2O(l)$
$H_2C_2O_4(s) \xrightarrow{-Conc.H_2SO_4} CO(g) + CO_2(g) + H_2O(l)$
(ii)Method 2;
Chemical equation
$HCOOH(s) \xrightarrow{-Conc.H_2SO_4} CO(g) + H_2O(l)$
$H_2CO_2(s) \xrightarrow{-Conc.H_2SO_4} CO(g) + H_2O(l)$

2. What method of gas collection is used during the preparation of carbon (II) oxide

Over water because the gas is insoluble in water.
Downward delivery because the gas is 1 ½ times denser than air .

3. What is the purpose of :

(i) Potassium hydroxide/sodium hydroxide in Method 1

To absorb/ remove carbon (II) oxide produced during the reaction.

$2KOH(aq) + CO_2(g) \rightarrow K_2CO_3(s) + H_2O(l)$

$2NaOH(aq) + CO_2(g) \rightarrow Na_2CO_3(s) + H_2O(l)$

(ii) Concentrated sulphuric(VI)acid in Method 1 and 2.

Dehydrating agent –removes the element of water (Hydrogen and Oxygen in ratio 2:1) present in both methanoic and ethan-1,2-dioic acid.

4. Describe the smell of carbon (II) oxide.

Colourless and odourless.

5. State and explain the observation made when carbon(IV)oxide is bubbled in lime water for a long time.

No white precipitate is formed.

6. Dry and wet/moist/damp litmus papers were separately put in a gas jar containing dry carbon (IV) oxide gas. State and explain the observations made.

Observation

-blue dry litmus paper remains blue
-red dry litmus paper remains red
- wet/moist/damp blue litmus paper remains blue
- wet/moist/damp red litmus paper remains red

Explanation

Carbon(II)oxide gas is a molecular compound that does not dissociate /ionize to release H+ ions and thus has no effect on litmus papers. Carbon(II)oxide gas is therefore a **neutral** gas.

7. Carbon (II) oxide gas was ignited at the end of a generator as below.

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(i)State the observations made in flame K.

Gas burns with a blue flame

(ii)Write the equation for the reaction taking place at flame K.

$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$

8. Carbon (II) oxide is a reducing agent. Explain

Experiment

Pass carbon (II) oxide through glass tube containing copper (II) oxide. Ignite any excess poisonous carbon (II) oxide.

Observation

Colour change from black to brown. Excess carbon (II) oxide burn with a blue flame.

Explanation

Carbon is a reducing agent. It is used to reduce metal oxide ores to metal, itself oxidized to carbon (IV) oxide gas. Carbon (II) Oxide reduces black copper (II) oxide to brown copper metal

Chemical Equation

$CuO(s) + CO(g) \rightarrow Cu(s) + CO_2(g)$
(black) (brown)

$PbO(s) + CO(g) \rightarrow Pb(s) + CO_2(g)$
(brown when hot/ yellow when cool) (grey)

$ZnO(s) + CO(g) \rightarrow Zn(s) + CO_2(g)$
(yellow when hot/ white when cool) (grey)

$Fe_2O_3(s) + 3CO(s) \rightarrow 2Fe(s) + 3CO_2(g)$
(brown when hot/cool (grey)

$Fe_3O_4(s) + 4CO(g) \rightarrow 3Fe(s) + 4CO_2(g)$
(brown when hot/cool (grey)

These reactions are used during the extraction of many metals from their ore.

9. Carbon (II) oxide is a pollutant. Explain.

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Carbon (II) oxide is highly poisonous/toxic.It preferentially combine with haemoglobin to form stable carboxyhaemoglobin in the blood instead of oxyhaemoglobin.This reduces the free haemoglobin in the blood causing nausea , coma then death.

10. The diagram below show a burning charcoal stove/burner/jiko. Use it to answer the questions that follow.



Explain the changes that take place in the burner

Explanation

Charcoal stove has air holes through which air enters. Air oxidizes carbon to carbon (IV) oxide gas at region I. This reaction is exothermic (-ΔH) producing more heat.

Chemical equation

$C(s) + O_2(g) \rightarrow CO_2(g)$

Carbon (IV) oxide gas formed rises up to meet more charcoal which reduces it to Carbon (II) oxide gas.

Chemical equation

$2CO_2(g) + O_2(g) \rightarrow 2CO(g)$

At the top of burner in region II, Carbon (II)oxide gas is further oxidized to Carbon(IV)oxide gas if there is plenty of air but escape if the air is limited poisoning the living things around.

Chemical equation

$2CO(g) + O_2(g) \xrightarrow{\text{(excess air)}} 2CO_2(g)$

11. Describe the test for the presence of carbon(II)oxide gas.

Experiment

Burn/Ignite the pure sample of the gas. Pass/Bubble the products into lime water/Calcium hydroxide.

Observation

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Colourless gas burns with a blue flame. A white precipitate is formed that dissolve on further bubbling of the products.

Chemical equation

$2CO\ (g) + O_2(g) \rightarrow 2CO_2\ (g)$ (gas burns with blue flame)

Chemical equation

$Ca(OH)_2\ (aq) + CO_2\ (g) \rightarrow CaCO_3\ (s) + H_2O(l)$

Chemical equation

$CO_2\ (g) + CaCO_3\ (s) + H_2O(l) \rightarrow Ca(HCO_3)_2\ (aq)$

12. State the main uses of carbon (II)oxide gas.

(i) As a fuel /water gas

(ii)As a reducing agent in the blast furnace for extracting iron from iron ore(Magnetite/Haematite)

(iii)As a reducing agent in extraction of Zinc from Zinc ore/Zinc blende

(iv) As a reducing agent in extraction of Lead from Lead ore/Galena

(v) As a reducing agent in extraction of Copper from Copper iron sulphide/Copper pyrites.

(iii)Carbonate(IV) (CO_3^{2-})and hydrogen carbonate(IV(HCO_3^-)

1.Carbonate (IV) (CO_3^{2-}) are normal salts derived from carbonic(IV)acid (H_2CO_3) and hydrogen carbonate (IV) (HCO_3^-) are acid salts derived from carbonic(IV)acid. Carbonic(IV)acid(H_2CO_3) is formed when carbon(IV)oxide gas is bubbled in water. It is a dibasic acid with two ionizable hydrogens.

$H_2CO_3(aq) \rightarrow 2H^+(aq) + CO_3^{2-}(aq)$

$H_2CO_3(aq) \rightarrow H^+(aq) + HCO_3^-(aq)$

2.Carbonate (IV) (CO_3^{2-}) are insoluble in water **except** Na_2CO_3 , K_2CO_3 and $(NH_4)_2CO_3$

3.Hydrogen carbonate (IV) (HCO_3^-) are soluble in water. Only five hydrogen carbonates exist. **$NaHCO_3$, $KHCO_3$,NH_4HCO_3 $Ca(HCO_3)_2$ and $Mg(HCO_3)_2$**

$Ca(HCO_3)_2$ and $Mg(HCO_3)_2$ exist only in aqueous solutions.

3.The following experiments show the effect of heat on Carbonate (IV) (CO_3^{2-}) and Hydrogen carbonate (IV) (HCO_3^-) salts:

Experiment

In a clean dry test tube place separately about 1.0 of the following:

Zinc(II)carbonate(IV), sodium hydrogen carbonate(IV), sodium carbonate(IV), Potassium carbonate(IV) ammonium carbonate(IV), potassium hydrogen carbonate(IV), Lead(II)carbonate(IV), Iron(II)carbonate(IV), and copper(II)carbonate(IV). Heat each portion gently the strongly. Test any gases produced with lime water.

Observation

(i)Colorless droplets form on the cooler parts of test tube in case of sodium carbonate(IV) and Potassium carbonate(IV).

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(ii)White residue/solid left in case of sodium hydrogen carbonate(IV), sodium carbonate(IV), Potassium carbonate(IV) and potassium hydrogen carbonate(IV).
(iii)Colour changes from blue/green to black in case of copper(II)carbonate(IV).
(iv) Colour changes from green to brown/yellow in case of Iron (II)carbonate(IV).
(v) Colour changes from white when cool to yellow when hot in case of Zinc (II) carbonate(IV).
(vi) Colour changes from yellow when cool to brown when hot in case of Lead (II) carbonate(IV).
(vii)Colourless gas produced that forms a white precipitate with lime water in all cases.

Explanation

1. Sodium carbonate(IV) and Potassium carbonate(IV) exist as hydrated salts with 10 molecules of water of crystallization that condenses and collects on cooler parts of test tube as a colourless liquid.

Chemical equation

$Na_2CO_3 .10H_2O(s) \rightarrow Na_2CO_3 (s) + 10H_2O(l)$

$K_2CO_3 .10H_2O(s) \rightarrow K_2CO_3 (s) + 10H_2O(l)$

2. Carbonate (IV) (CO_3^{2-}) and Hydrogen carbonate (IV) (HCO_3^-) salts decompose on heating except Sodium carbonate(IV) and Potassium carbonate(IV).

(a) Sodium hydrogen carbonate(IV) and Potassium hydrogen carbonate(IV) decompose on heating to form sodium carbonate(IV) and Potassium carbonate(IV).Water and carbon(IV)oxide gas are also produced.

Chemical equation

$2NaHCO_3 (s) \rightarrow Na_2CO_3 (s) + H_2O(l) + CO_2 (g)$
(white) (white)

$2KHCO_3 (s) \rightarrow K_2CO_3 (s) + H_2O(l) + CO_2 (g)$
(white) (white)

(b) Calcium hydrogen carbonate(IV) and Magnesium hydrogen carbonate(IV) decompose on heating to form insoluble Calcium carbonate(IV) and Magnesium carbonate(IV).Water and carbon(IV)oxide gas are also produced.

Chemical equation

$Ca(HCO_3)_2 (aq) \rightarrow CaCO_3 (s) + H_2O(l) + CO_2 (g)$
(Colourless solution) (white)

$Mg(HCO_3)_2 (aq) \rightarrow MgCO_3 (s) + H_2O(l) + CO_2 (g)$
(Colourless solution) (white)

(c) Ammonium hydrogen carbonate(IV) decompose on heating to form ammonium carbonate(IV) .Water and carbon(IV)oxide gas are also produced.

Chemical equation

$2NH_4HCO_3 (s) \rightarrow (NH_4)_2CO_3 (s) + H_2O(l) + CO_2 (g)$
(white) (white)

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(d)All other carbonates decompose on heating to form the metal oxide and produce carbon(IV)oxide gas e.g.

Chemical equation

$MgCO_3$ (s) -> MgO (s) + CO_2 (g)
(white solid) (white solid)

Chemical equation

$BaCO_3$ (s) -> BaO (s) + CO_2 (g)
(white solid) (white solid)

Chemical equation

$CaCO_3$ (s) -> CaO (s) + CO_2 (g)
(white solid) (white solid)

Chemical equation

$CuCO_3$ (s) -> CuO (s) + CO_2 (g)
(blue/green solid) (black solid)

Chemical equation

$ZnCO_3$ (s) -> ZnO (s) + CO_2 (g)
(white solid) (white solid when cool/
Yellow solid when hot)

Chemical equation

$PbCO_3$ (s) -> PbO (s) + CO_2 (g)
(white solid) (yellow solid when cool/
brown solid when hot)

4.The following experiments show the presence of Carbonate (IV) (CO_3^{2-}) and Hydrogen carbonate (IV) (HCO_3^-) ions in sample of a salt:

(a)Using Lead(II) nitrate(V)

I. Using a portion of salt solution in a test tube .add four drops of Lead(II)nitrate(V)solution.Preserve.

Observation	inference
White precipitate/ppt	CO_3^{2-} ,SO_3^{2-} ,SO_4^{2-} ,Cl^-

II. To the preserved solution, add six drops of dilutte nitric(V)acid. Preserve.

Observation	inference
White precipitate/ppt persists	SO_4^{2-} ,Cl^-
White precipitate/ppt dissolves	CO_3^{2-} ,SO_3^{2-}

II. To the preserved sample(that forms a precipitate),heat to boil.

Observation	inference
White precipitate/ppt persists	SO_4^{2-}

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White precipitate/ppt dissolves	Cl^-

II. To the preserved sample(that do not form a precipitate),add three drops of acidified potassium manganate(VII)/lime water

Observation	inference
Effervescence/bubbles/fizzing colourless gas produced Acidified $KMnO_4$ decolorized/no white precipitate on lime water	SO_3^{2-}
Effervescence/bubbles/fizzing colourless gas produced Acidified $KMnO_4$ not decolorized/ white precipitate on lime water	CO_3^{2-}

Experiments/Observations:

(b)Using Barium(II)nitrate(V)/ Barium(II)chloride

I. To about 5cm3 of a salt solution in a test tube add four drops of Barium(II) nitrate (V) / Barium(II)chloride. Preserve.

Observation	Inference
White precipitate/ppt	SO_4^{2-} , SO_3^{2-} , CO_3^{2-} ions

II. To the preserved sample in (I) above, add six drops of 2M nitric(V) acid . Preserve.

Observation 1

Observation	Inference
White precipitate/ppt persists	SO_4^{2-} , ions

Observation 2

Observation	Inference
White precipitate/ppt dissolves	SO_3^{2-} , CO_3^{2-} , ions

III.To the preserved sample observation 2 in (II) above, add 4 drops of acidified potassium manganate(VII) /dichromate(VI).

Observation 1

Observation	Inference
(i)acidified potassium manganate(VII)decolorized (ii)Orange colour of acidified potassium dichromate(VI) turns to green	SO_3^{2-} ions

Observation 2

Observation	Inference
(i)acidified potassium manganate(VII) not decolorized	CO_3^{2-} ions

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(ii)Orange colour of acidified potassium dichromate(VI) does not turns to green	

Explanations

Using Lead(II)nitrate(V)

(i)Lead(II)nitrate(V) solution reacts with chlorides(Cl^-), Sulphate (VI) salts (SO_4^{2-}), Sulphate (IV)salts (SO_3^{2-}) and carbonates(CO_3^{2-}) to form the insoluble white precipitate of Lead(II)chloride, Lead(II)sulphate(VI), Lead(II) sulphate (IV) and Lead(II)carbonate(IV).

Chemical/ionic equation:

$Pb^{2+}(aq) + Cl^-(aq) \rightarrow PbCl_2(s)$

$Pb^{2+}(aq) + SO_4^{2+}(aq) \rightarrow PbSO_4(s)$

$Pb^{2+}(aq) + SO_3^{2+}(aq) \rightarrow PbSO_3(s)$

$Pb^{2+}(aq) + CO_3^{2+}(aq) \rightarrow PbCO_3(s)$

(ii)When the insoluble precipitates are acidified with nitric(V) acid,

- Lead(II)chloride and Lead(II)sulphate(VI) do not react with the acid and thus their white precipitates remain/ persists.
- Lead(II) sulphate (IV) and Lead(II)carbonate(IV) reacts with the acid to form **soluble** Lead(II) nitrate (V) and produce/effervesces/fizzes/bubbles out **sulphur(IV)oxide** and **carbon(IV)oxide** gases respectively.

. Chemical/ionic equation:

$PbSO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + Pb^{2+}(aq) + SO_2(g)$

$PbCO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + Pb^{2+}(aq) + CO_2(g)$

(iii)When Lead(II)chloride and Lead(II)sulphate(VI) are heated/warmed;

- Lead(II)chloride dissolves in hot water/on boiling(recrystallizes on cooling)
- Lead(II)sulphate(VI) do not dissolve in hot water thus its white precipitate persists/remains on heating/boiling.

(iv)When sulphur(IV)oxide and carbon(IV)oxide gases are produced;

- **sulphur(IV)oxide** will decolorize acidified potassium manganate(VII) and / or Orange colour of acidified potassium dichromate(VI) will turns to green.

Carbon(IV)oxide will not.

Chemical equation:

$5SO_3^{2-}(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(l)$

(purple) (colourless)

$3SO_3^{2-}(aq) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \rightarrow 3SO_4^{2-}(aq) + 2Cr^{3+}(aq) + 4H_2O(l)$

(Orange) (green)

- **Carbon(IV)oxide** forms an insoluble white precipitate of calcium carbonate if three drops of lime water are added into the reaction test tube when effervescence is taking place. **Sulphur(IV)oxide will not.**

Chemical equation:

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

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These tests should be done immediately after acidifying to ensure the gases produced react with the oxidizing agents/lime water.

Using Barium(II)nitrate(V)/ Barium(II)Chloride

(i)Barium(II)nitrate(V) and/ or Barium(II)chlo**ride** solution reacts with Sulphate (VI) salts (SO_4^{2-}), Sulphate (IV)salts (SO_3^{2-}) and carbonates(CO_3^{2-}) to form the insoluble white precipitate of Barium(II)sulphate(VI), Barium(II) sulphate (IV) and Barium(II)carbonate(IV).

Chemical/ionic equation:

$Ba^{2+}(aq) + SO_4^{2+}(aq) \rightarrow BaSO_4(s)$

$Ba^{2+}(aq) + SO_3^{2+}(aq) \rightarrow BaSO_3(s)$

$Ba^{2+}(aq) + CO_3^{2+}(aq) \rightarrow BaCO_3(s)$

(ii)When the insoluble precipitates are acidified with nitric(V) acid,

- Barium (II)sulphate(VI) do not react with the acid and thus its white precipitates remain/ persists.
- Barium(II) sulphate (IV) and Barium(II)carbonate(IV) reacts with the acid to form **soluble** Barium(II) nitrate (V) and produce /effervesces /fizzes/ bubbles out **sulphur(IV)oxide** and **carbon(IV)oxide** gases respectively.

. Chemical/ionic equation:

$BaSO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + Ba^{2+}(aq) + SO_2(g)$

$BaCO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + Ba^{2+}(aq) + CO_2(g)$

(iii) When sulphur(IV)oxide and carbon(IV)oxide gases are produced;

- **sulphur(IV)oxide** will decolorize acidified potassium manganate(VII) and / or Orange colour of acidified potassium dichromate(VI) will turns to green.

Carbon(IV)oxide will not.

Chemical equation:

$5SO_3^{2-}(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(l)$
(purple) (colourless)

$3SO_3^{2-}(aq) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \rightarrow 3SO_4^{2-}(aq) + 2Cr^{3+}(aq) + 4H_2O(l)$
(Orange) (green)

- **Carbon(IV)oxide** forms an insoluble white precipitate of calcium carbonate if three drops of lime water are added into the reaction test tube when effervescence is taking place. **Sulphur(IV)oxide will not.**

Chemical equation:

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

These tests should be done immediately after acidifying to ensure the gases produced react with the oxidizing agents/lime water.

(iii) Sodium carbonate(IV) (Na_2CO_3)

(a)Extraction of sodium carbonate from soda ash

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Sodium carbonate naturally occurs in Lake Magadi in Kenya as Trona.trona is the double salt ; sodium sesquicarbonate. $NaHCO_3 .Na_2CO_3 .H_2O$.It is formed from the volcanic activity that takes place in Lake Naivasha, Nakuru ,Bogoria and Elementeita .All these lakes drain into Lake Magadi through underground rivers. Lake Magadi has no outlet.

Solubility of Trona decrease with increase in temperature.High temperature during the day causes trona to naturally crystallize .It is mechanically scooped/dredged/dug and put in a furnace.

Inside the furnace, trona decompose into soda ash/sodium carbonate.

Chemical equation

$2NaHCO_3 .Na_2CO_3 .H_2O\ (s) \rightarrow 3Na_2CO_3\ (s) + 5H_2O(l) + CO_2\ (g)$

(trona) (soda ash)

Soda ash is then bagged and sold as Magadi soda.It is mainly used:

(i)in making glass to lower the melting point of raw materials (sand/SiO_2 from 1650°C and CaO from 2500°C to around 1500°C)

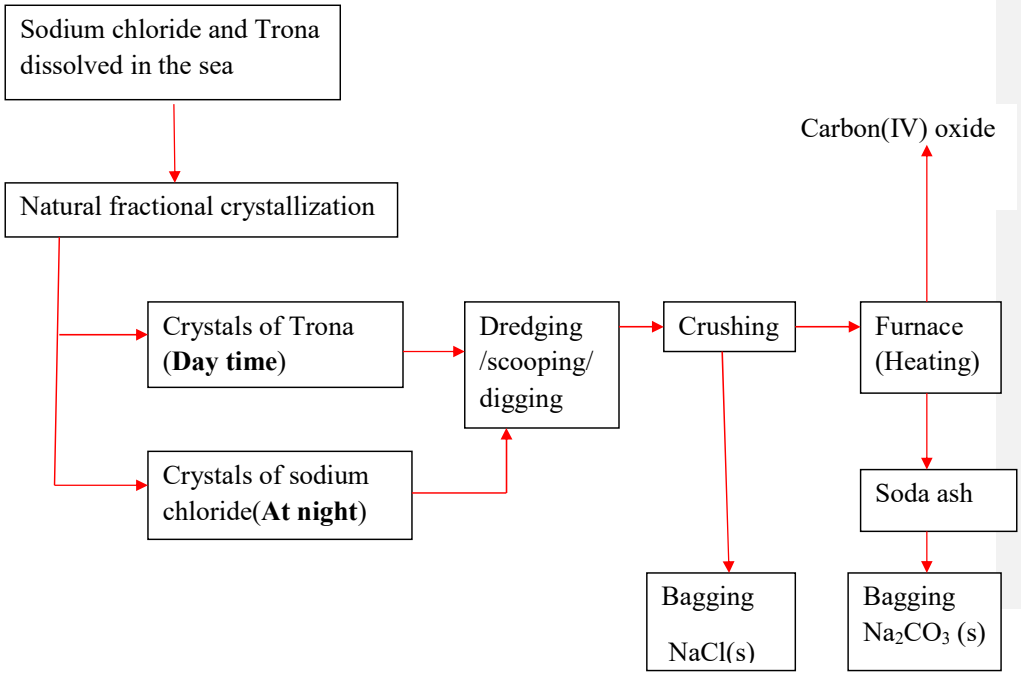
(ii)in softening hard water

(iii)in the manufacture of soapless detergents.

(iv)Swimming pool "pH increaser"

Sodium chloride is also found dissolved in the lake. Solubility of sodium chloride decrease with decreases in temperature/ sodium chloride has lower solubility at lower temperatures. When temperatures decrease at night it crystallize out .The crystals are then mechanically dug/dredged /scooped then packed for sale as animal/cattle feeds and seasoning food.

Summary flow diagram showing the extraction of Soda ash from Trona



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b)The Solvay process for industrial manufacture of sodium carbonate(IV)

(i)Raw materials.

-**Brine** /Concentrated Sodium chloride from salty seas/lakes.

-**Ammonia** gas from Haber.

-**Limestone** /Calcium carbonate from chalk /limestone rich rocks.

-**Water** from rivers/lakes.

(ii)Chemical processes

Ammonia gas is passed **up** to meet a **downward** flow of sodium chloride solution / brine to form **ammoniated** brine/**ammoniacal** brine **mixture** in the **ammoniated brine chamber**

The ammoniated brine mixture is then pumped up, atop the carbonator/ solvay tower.

In the carbonator/ solvay tower, ammoniated brine/ammoniacal brine mixture slowly trickle down to meet an upward flow of carbon(IV)oxide gas.

The carbonator is shelved /packed with quartz/broken glass to

(i) reduce the rate of flow of ammoniated brine/ammoniacal brine mixture.

(ii)increase surface area of the liquid mixture to ensure a lot of ammoniated brine/ammoniacal brine mixture react with carbon(IV)oxide gas.

Insoluble sodium hydrogen carbonate and soluble ammonium chloride are formed from the reaction.

Chemical equation

$CO_2(g) + H_2O(l) + NaCl\ (aq) + NH_3(g) \rightarrow NaHCO_3(s) + NH_4Cl(aq)$

The products are then filtered. **Insoluble** sodium hydrogen carbonate forms the **residue** while soluble ammonium chloride forms the **filtrate**.

Sodium hydrogen carbonate itself can be used:

(i) as baking powder and preservation of some soft drinks.

(ii) as a buffer agent and antacid in animal feeds to improve fibre digestion.

(iii) making dry chemical fire extinguishers.

In the Solvay process Sodium hydrogen carbonate is then heated to form Sodium carbonate/soda ash, water and carbon (IV) oxide gas.

Chemical equation

$2NaHCO_3\ (s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(l)$

Sodium carbonate is stored ready for use in:

(i) during making glass/lowering the melting point of mixture of sand/SiO_2 from 1650^oC and CaO from 2500^oC to around 1500^oC

(ii) in softening hard water

(iii) in the manufacture of soapless detergents.

(iv) swimming pool “pH increaser”.

Water and carbon(IV)oxide gas are recycled back to the ammoniated brine/ammoniacal brine chamber.

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More carbon(IV)oxide is produced in the kiln/furnace. Limestone is heated to decompose into Calcium oxide and carbon(IV)oxide.
Chemical equation

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

Carbon(IV)oxide is recycled to the carbonator/solvay tower. Carbon (IV)oxide is added water in the **slaker** to form Calcium hydroxide. This process is called **slaking**.
Chemical equation

$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$

Calcium hydroxide is mixed with ammonium chloride from the carbonator/solvay tower in the **ammonia regeneration chamber** to form Calcium chloride, water and more ammonia gas.
Chemical equation

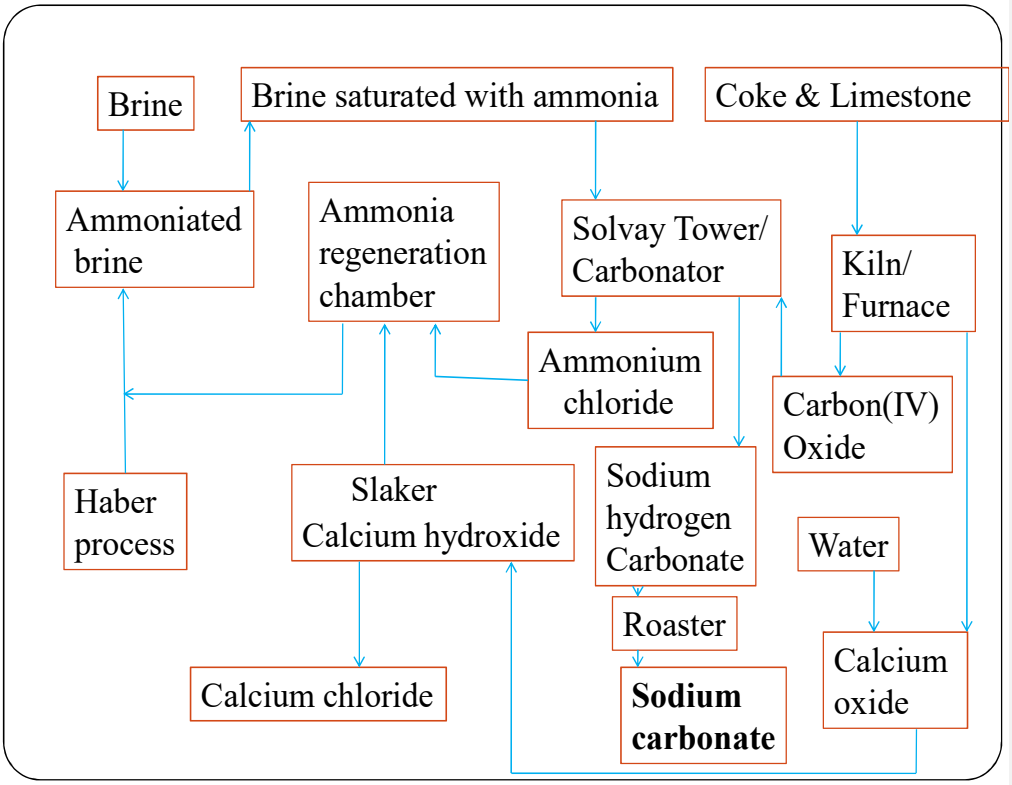
$Ca(OH)_2(aq) + 2NH_4Cl(aq) \rightarrow CaCl_2(s) + 2NH_3(g) + H_2O(l)$

$NH_3(g)$ and $H_2O(l)$ are recycled.
Calcium chloride may be used:
(i)as drying agent in the school laboratory during gas preparation (except ammonia gas)
(ii)to lower the melting point of solid sodium chloride / rock salt salts during the Downs process for industrial extraction of sodium metal.

Detailed Summary flow diagram of Solvay Process

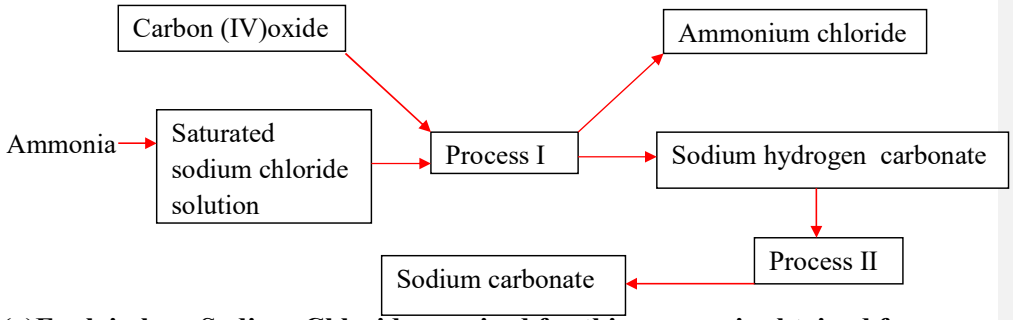
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Practice

1. The diagram below shows part of the Solvay process used in manufacturing sodium carbonate. Use it to answer the questions that follow.



(a)Explain how Sodium Chloride required for this process is obtained from the sea.

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Sea water is pumped /scooped into shallow pods. Evaporation of most of the water takes place leaving a very concentrated solution.

(b)(i) Name process:

I. Filtration

II. Decomposition

(ii) Write the equation for the reaction in process:

Process I

Chemical equation

$CO_2(g) + H_2O(l) + NaCl\ (aq) + NH_3(g) \rightarrow NaHCO_3(s) + NH_4Cl(aq)$

Process II

Chemical equation

$2NaHCO_3\ (s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(l)$

(c)(i) Name two substances recycled in the solvay process

Ammonia gas , Carbon(IV)Oxide and Water.

(ii)Which is the by-product of this process?

Calcium(II)Chloride /$CaCl_2$

(iii)State two uses that the by-product can be used for:

1. As a drying agent in the school laboratory preparation of gases.
2. In the Downs cell/process for extraction of Sodium to lower the melting point of rock salt.

(iv)Write the chemical equation for the formation of the by-products in the Solvay process.

Chemical equation

$Ca(OH)_2\ (aq) + 2NH_4Cl\ (aq) \rightarrow CaCl_2(s) + 2NH_3(g) + H_2O(l)$

(d)In an experiment to determine the % purity of Sodium carbonate produced in the Solvay process, 2.15g of the sample reacted with exactly 40.0cm3 of 0.5M Sulphuric (VI)acid.

(i)Calculate the number of moles of sodium carbonate that reacted.

Chemical equation

$Na_2CO_3\ (aq) + H_2SO_4\ (aq) \rightarrow Na_2SO_4\ (aq) + CO_2(g) + H_2O(l)$

Mole ratio $Na_2CO_3 : H_2SO_4 \Rightarrow 1:1$

Moles H_2SO_4 = $\frac{\text{Molarity} \times \text{Volume}}{1000} \Rightarrow \frac{0.5 \times 40.0}{1000}$ = **0.02 Moles**

Moles of Na_2CO_3 = **0.02 Moles**

(ii)Determine the % of sodium carbonate in the sample.

Molar mass of Na_2CO_3 = **106g**

Mass of Na_2CO_3 = moles x Molar mass => 0.02 x 106 = **2.12 g**

% of Na_2CO_3 = $\frac{(2.12\ g \times 100)}{2.15}$ = **98.6047%**

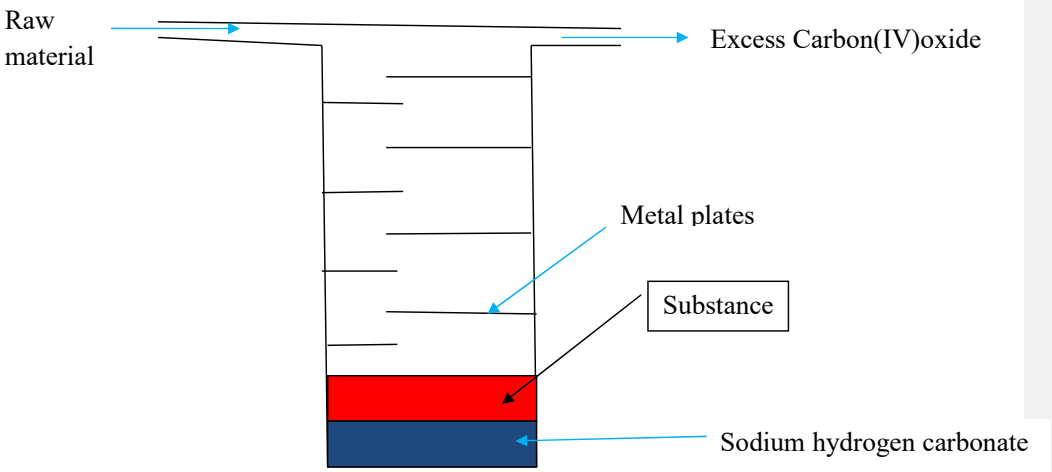
(e) State two uses of soda ash.

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(i) during making glass/lowering the melting point of mixture of sand/SiO_2 from 1650°C and CaO from 2500°C to around 1500°C
(ii) in softening hard water
(iii) in the manufacture of soapless detergents.
(iv) swimming pool "pH increaser".

(f)The diagram below shows a simple ammonia soda tower used in manufacturing sodium carbonate .Use it to answer the questions that follow:



(i)Name the raw materials needed in the above process

-Ammonia
-Water
-Carbon(IV)oxide
-Limestone
-Brine/ Concentrated sodium chloride

(ii)Identify substance A

Ammonium chloride /NH_4Cl

(iii) Write the equation for the reaction taking place in:

I.Tower.

Chemical equation

$CO_2(g) + NaCl\ (aq) + H_2O(l) + NH_3(g) \rightarrow NaHCO_3(s) + NH_4Cl(aq)$

II. Production of excess carbon (IV)oxide.

Chemical equation

$CaCO_3\ (s) \rightarrow CaO(s) + CO_2(g)$

III. The regeneration of ammonia

Chemical equation

$Ca(OH)_2\ (aq) + 2NH_4Cl\ (aq) \rightarrow CaCl_2(s) + 2NH_3(g) + H_2O(l)$

(iv)Give a reason for having the circular metal plates in the tower.

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-To slow the downward flow of brine.
-To increase the rate of dissolving of ammonia.
-To increase the surface area for dissolution

(v)Name the gases recycled in the process illustrated above.

Ammonia gas , Carbon(IV)Oxide and Water.

2. Describe how you would differentiate between carbon (IV)oxide and carbon(II)oxide using chemical method.

Method I

-Bubble both gases in lime water/$Ca(OH)_2$
-white precipitate is formed if the gas is carbon (IV) oxide
- **No** white precipitate is formed if the gas is carbon (II) oxide

Method II

-ignite both gases
- Carbon (IV) oxide does **not** burn/ignite
- Carbon (II) oxide burn with a blue non-sooty flame.

Method III

-Lower a burning splint into a gas containing each gas separately.
-burning splint is extinguished if the gas is carbon (IV) oxide
-burning splint is **not** extinguished if the gas is carbon (II) oxide.

3.Using Magnesium sulphate(VI)solution ,describe how you can differentiate between a solution of sodium carbonate from a solution of sodium hydrogen carbonate

-Add Magnesium sulphate(VI) solution to separate portions of a solution of sodium carbonate and sodium hydrogen carbonate in separate test tubes
-White precipitate is formed in test tube containing sodium carbonate
-No white precipitate is formed in test tube containing sodium hydrogen carbonate.

Chemical equation

Na_2CO_3 (aq) +$MgSO_4$ (aq) -> Na_2SO_4 (aq) + $MgCO_3$(s)
(white ppt)

Ionic equation

CO_3^{2-} (aq) + Mg^{2+} (aq) -> $MgCO_3$(s)
(white ppt)

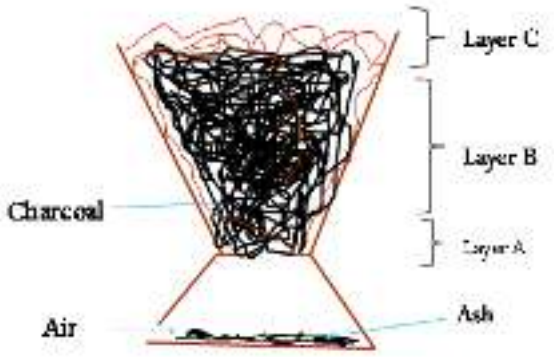
Chemical equation

$2NaHCO_3$ (aq) +$MgSO_4$ (aq) -> Na_2SO_4 (aq) + $Mg(HCO_3)_2$ (aq)
(colourless solution)

4. The diagram below shows a common charcoal burner .Assume the burning take place in a room with sufficient supply of air.

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(a)Explain what happens around:

(i)Layer A

Sufficient/excess air /oxygen enter through the air holes into the burner .It reacts with/oxidizes Carbon to carbon(IV)oxide

Chemical equation

$C(s) + O_2(g) \rightarrow CO_2(g)$

(ii)Layer B

Hot carbon(IV)oxide rises up and is reduced by more carbon/charcoal to carbon (II)oxide.

Chemical equation

$C(s) + CO_2(g) \rightarrow 2CO(g)$

(ii)Layer C

Hot carbon(II)oxide rises up and burns with a blue flame to be oxidized by the excess air to form carbon(IV)oxide.

$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$

(b)State and explain what would happen if the burner is put in an enclosed room.

The hot poisonous /toxic carbon(II)oxide rising up will not be oxidized to Carbon(IV)oxide.

(c)Using a chemical test, describe how you would differentiate two unlabelled black solids suspected to be charcoal and copper(II)oxide.

Method I

-Burn/Ignite the two substances separately.
-Charcoal burns with a blue flame
- Copper(II)oxide does not burn

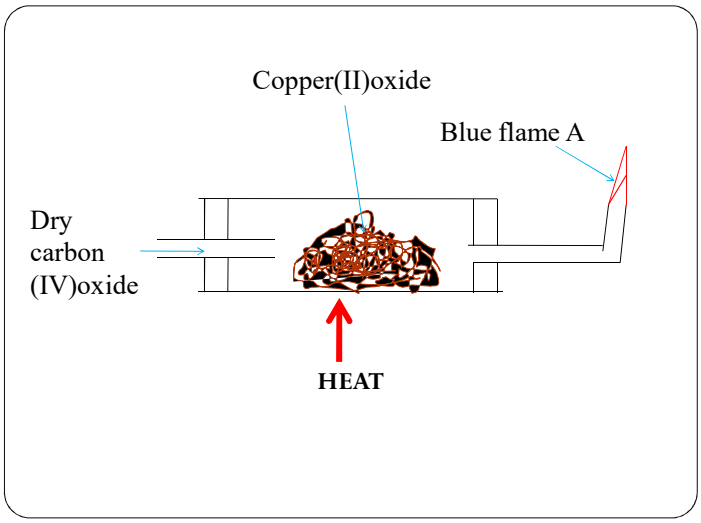
Method II

-Add dilute sulphuric(VI)acid/Nitric(V)acid/Hydrochloric acid separately.
-Charcoal does not dissolve.
- Copper(II)oxide dissolves to form a colourless solution.

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5. Excess Carbon(II)oxide was passed over heated copper(II)oxide as in the set up shown below for five minutes.



(a)State and explain the observations made in the combustion tube.

Observation

Colour change from black to brown

Explanation

Carbon (II)oxide reduces black copper(II)oxide to brown copper metal itself oxidized to Carbon(IV)oxide.

Chemical equation

$CO(g) + CuO(s) \rightarrow Cu(s) + CO_2(g)$

(black) (brown)

(b) (i)Name the gas producing flame A

Carbon(II)oxide

(ii)Why should the gas be burnt?

It is toxic/poisonous

(iii)Write the chemical equation for the production of flame A

$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$

(c)State and explain what happens when carbon(IV)oxide is prepared using Barium carbonate and dilute sulphuric(VI)acid.

Reaction starts then stops after sometime producing small/little quantity of carbon(IV)oxide gas.

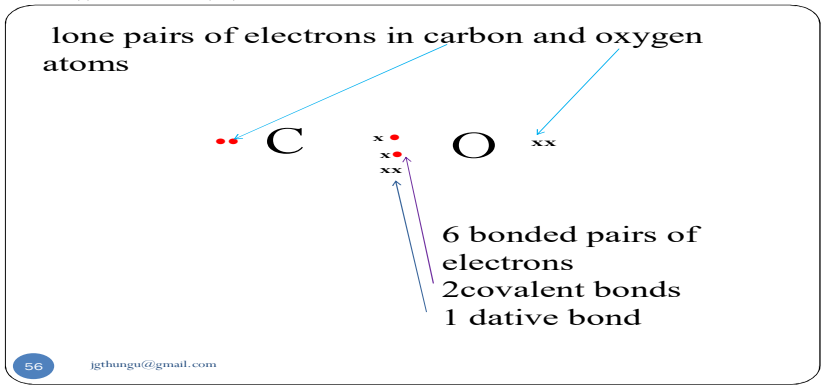
Barium carbonate react with dilute sulphuric(VI)acid to form insoluble Barium sulphate(VI) that cover/coat unreacted Barium carbonate stopping further reaction to produce more Carbon(IV)oxide.

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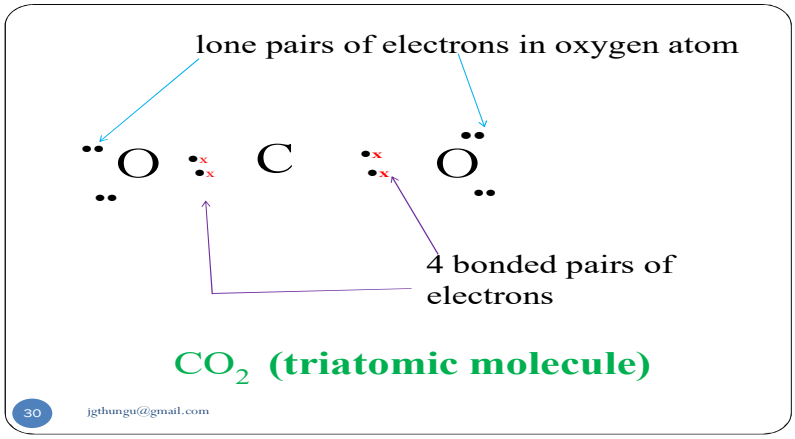
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(d) Using dot () and cross(x) to represent electrons show the bonding in a molecule of :

(i) Carbon(II)oxide



(ii) Carbon(IV)Oxide.



(e) Carbon (IV)oxide is an environmental pollutant of global concern. Explain.

-It is a green house gas thus causes global warming.

-It dissolves in water to form acidic carbonic acid which causes "acid rain"

(f)Explain using chemical equation why lime water is used to test for the presence of Carbon (IV) oxide instead of sodium hydroxide.

Using lime water/calcium hydroxide:

- a visible white precipitate of calcium carbonate is formed that dissolves on bubbling excess Carbon (IV) oxide gas

Chemical equation

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

(white precipitate)

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$CaCO_3$ (aq) + H_2O(l) + CO_2 (g) -> $Ca(HCO_3)_2$ (aq)

Using sodium hydroxide:

- **No** precipitate of sodium carbonate is formed Both sodium carbonate and sodium hydrogen carbonate are soluble salts/dissolves.

Chemical equation

2NaOH (aq) + CO_2 (g) -> Na_2CO_3 (s) + H_2O(l)

(No white precipitate)

Na_2CO_3 (s) + H_2O(l) + CO_2 (g) -> 2$NaHCO_3$ (s)

(g)Ethan-1,2-dioic acid and methanoic acid may be used to prepare small amount of carbon(II)oxide in a school laboratory.

(i) Explain the modification in the set up when using one over the other.

Before carbon(II)oxide is collected:

-when using methanoic acid, **no** concentrated sodium/potassium hydroxide is needed to absorb Carbon(IV)oxide.

-when using ethan-1,2-dioic acid, concentrated sodium/potassium hydroxide is needed to absorb Carbon(IV)oxide.

(ii)Write the equation for the reaction for the formation of carbon(II)oxide from:

I.Methanoic acid.

Chemical equation HCOOH(aq) -> CO(g) + H_2O(l)

II. Ethan-1,2-dioic acid

Chemical equation HOOCCOOH(aq) -> CO_2(g)+CO(g)+H_2O(l)

(h)Both carbon(II)oxide and carbon(IV)oxide affect the environment. Explain why carbon(II)oxide is more toxic/poisonous.

-Both gases are colourless,denser than water and odourless.

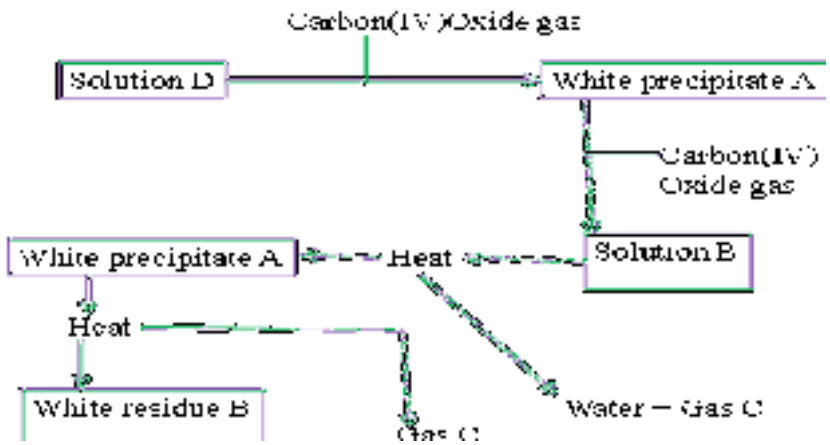
-Carbon(II)oxide is preferentially absorbed by human/mammalian haemoglobin when inhaled forming stable carboxyhaemoglobin instead of oxyhaemoglobin.This reduces the free haemoglobin in the blood leading to suffocation and quick death. **--** Carbon(IV)oxide is a green house gas that increases global warming.

-Carbon(II)oxide is readily oxidized to carbon(IV)oxide

6.Study the flow chart below and use it to answer the questions that follow.

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(a)Name:

(i)the white precipitate A

Calcium carbonate

(ii) solution B

Calcium hydrogen carbonate

(iii) gas C

Carbon(IV)oxide

(iv) white residue B

Calcium oxide

(v) solution D

Calcium hydroxide/lime water

(b)Write a balanced chemical equation for the reaction for the formation of:

(i) the white precipitate A from solution D

Chemical equation

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

(ii) the white precipitate A from solution B

Chemical equation

$Ca(HCO_3)_2(aq) \rightarrow CO_2(g) + CaCO_3(s) + H_2O(l)$

(iii) solution B from the white precipitate A

Chemical equation

$CO_2(g) + CaCO_3(s) + H_2O(l) \rightarrow Ca(HCO_3)_2(aq)$

(iv) white residue B from the white precipitate A

Chemical equation

$CaCO_3(s) \rightarrow CO_2(g) + CaO(s)$

(iv) reaction of white residue B with water

Chemical equation

$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$

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INTRODUCTION TO SALTS

1.(a) A salt is an ionic compound formed when the cation from a base combine with the anion derived from an acid.

A salt is therefore formed when the hydrogen ions in an acid are replaced wholly/fully or partially/partly, directly or indirectly by a metal or ammonium radical.

(b) The number of ionizable/replaceable hydrogen in an acid is called **basicity** of an acid.

Some acids are therefore:

(i)**mono**basic acids generally denoted **H**X e.g.

HCl, **H**NO_3,HCOO**H**,CH3COO**H**.

(ii)**di**basic acids ; generally denoted $\mathbf{H_2}$X e.g.

$\mathbf{H_2}SO_4$, $\mathbf{H_2}SO_3$, $\mathbf{H_2}CO_3$,**H**OOCOO**H**.

(iii)**tri**basic acids ; generally denoted $\mathbf{H_3}$X e.g.

$\mathbf{H_3}PO_4$.

(c) Some salts are **normal** salts while others are **acid** salts.

(i)A normal salt is formed when all the ionizable /replaceable hydrogen in an acid is replaced by a metal or metallic /ammonium radical.

(ii)An acid salt is formed when part/portion the ionizable /replaceable hydrogen in an acid is replaced by a metal or metallic /ammonium radical.

Table showing normal and acid salts derived from common acids

Acid name	Chemical formula	Basicity	Normal salt	Acid salt
Hydrochloric acid	HCl	Monobasic	Chloride(Cl^-)	None
Nitric(V)acid	HNO_3	Monobasic	Nitrate(V)(NO_3^-)	None
Nitric(III)acid	HNO_2	Monobasic	Nitrate(III)(NO_2^-)	None
Sulphuric(VI) acid	H_2SO_4	Dibasic	Sulphate(VI) (SO_4^{2-})	Hydrogen sulphate(VI) (HSO_4^-)
Sulphuric(IV) acid	H_2SO_3	Dibasic	Sulphate(IV) (SO_3^{2-})	Hydrogen sulphate(IV) (HSO_3^-)
Carbonic(IV)acid	H_2CO_3	Dibasic	Carbonate(IV)(CO_3^{2-})	Hydrogen carbonate(IV)

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				(HCO_3^-)
Phosphoric(V) acid	H_3PO_4	Tribasic	Phosphate(V)(PO_4^{3-})	Dihydrogen phosphate(V) ($H_2PO_4^{2-}$) Hydrogen diphosphate(V) ($HP_2O_4^{2-}$)

The table below show shows some examples of salts.

Base/alkali	Cation	Acid	Anion	Salt	Chemical name of salts
NaOH	Na^+	HCl	Cl^-	NaCl	Sodium(I)chloride
$Mg(OH)_2$	Mg^{2+}	H_2SO_4	SO_4^{2-}	$MgSO_4$ $Mg(HSO_4)_2$	Magnesium sulphate(VI) Magnesium hydrogen sulphate(VI)
$Pb(OH)_2$	Pb^{2+}	HNO_3	NO_3^-	$Pb(NO_3)_2$	Lead(II)nitrate(V)
$Ba(OH)_2$	Ba^{2+}	HNO_3	NO_3^-	$Ba(NO_3)_2$	Barium(II)nitrate(V)
$Ca(OH)_2$	Ba^{2+}	H_2SO_4	SO_4^{2-}	$MgSO_4$	Calcium sulphate(VI)
NH_4OH	NH_4^+	H_3PO_4	PO_4^{3-}	$(NH_4)_3PO_4$ $(NH_4)_2HPO_4$ $NH_4H_2PO_4$	Ammonium phosphate(V) Diammonium phosphate(V) Ammonium diphosphate(V)
KOH	K^+	H_3PO_4	PO_4^{3-}	K_3PO_4	Potassium phosphate(V)
$Al(OH)_3$	Al^{3+}	H_2SO_4	SO_4^{2-}	$Al_2(SO_4)_2$	Aluminium(III)sulphate(VI)
$Fe(OH)_2$	Fe^{2+}	H_2SO_4	SO_4^{2-}	$FeSO_4$	Iron(II)sulphate(VI)
$Fe(OH)_3$	Fe^{3+}	H_2SO_4	SO_4^{2-}	$Fe_2(SO_4)_2$	Iron(III)sulphate(VI)

(d) Some salts undergo **hygroscopy**, **deliquescence** and **efflorescence**.

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(i) Hygroscopic salts /compounds are those that absorb water from the atmosphere but do not form a solution.
Some salts which are hygroscopic include anhydrous copper(II)sulphate(VI), anhydrous cobalt(II)chloride, potassium nitrate(V) common table salt.
(ii)Deliquescent salts /compounds are those that absorb water from the atmosphere and form a solution.
Some salts which are deliquescent include: Sodium nitrate(V),Calcium chloride, Sodium hydroxide, Iron(II)chloride, Magnesium chloride.
(iii)Efflorescent salts/compounds are those that lose their water of crystallization to the atmosphere.
Some salts which effloresces include: sodium carbonate decahydrate, Iron(II)sulphate(VI)heptahydrate, sodium sulphate (VI)decahydrate.
(e)Some salts contain water of crystallization.They are hydrated.Others do not contain water of crystallization. They are anhydrous.

Table showing some hydrated salts.

Name of hydrated salt	Chemical formula
Copper(II)sulphate(VI)pentahydrate	$CuSO_4.5H_2O$
Aluminium(III)sulphate(VI)hexahydrate	$Al_2(SO_4)_3.6H_2O$
Zinc(II)sulphate(VI)heptahydrate	$ZnSO_4.7H_2O$
Iron(II)sulphate(VI)heptahydrate	$FeSO_4.7H_2O$
Calcium(II)sulphate(VI)heptahydrate	$CaSO_4.7H_2O$
Magnesium(II)sulphate(VI)heptahydrate	$MgSO_4.7H_2O$
Sodium sulphate(VI)decahydrate	$Na_2SO_4.10H_2O$
Sodium carbonate(IV)decahydrate	$Na_2CO_3.10H_2O$
Potassium carbonate(IV)decahydrate	$K_2CO_3.10H_2O$
Potassium sulphate(VI)decahydrate	$K_2SO_4.10H_2O$

(f)Some salts exist as a simple salt while some as complex salts. Below are some complex salts.

Table of some complex salts

Name of complex salt	Chemical formula	Colour of the complex salt
Tetraamminecopper(II)sulphate(VI)	$Cu(NH_3)_4SO_4H_2O$	Royal/deep blue solution
Tetraamminezinc(II)nitrate(V)	$Zn(NH_3)_4(NO_3)_2$	Colourless solution
Tetraamminecopper(II) nitrate(V)	$Cu(NH_3)_4(NO_3)_2$	Royal/deep blue solution
Tetraamminezinc(II)sulphate(VI)	$Zn(NH_3)_4SO_4$	Colourless solution

(g)Some salts exist as two salts in one. They are called **double salts**.

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Table of some double salts

Name of double salts	Chemical formula
Trona(sodium sesquicarbonate)	Na_2CO_3 $NaHCO_3.2H_2O$
Ammonium iron(II)sulphate(VI)	$FeSO_4(NH_4)_2SO_4.2H_2O$
Ammonium aluminium(III)sulphate(VI)	$Al_2(SO_4)_3(NH_4)_2SO_4.H_2O$

(h)Some salts dissolve in water to form a **solution**. They are said to be soluble. Others do not dissolve in water. They form a **suspension/precipitate** in water.

Table of solubility of salts

Soluble salts	Insoluble salts
All nitrate(V)salts	
All sulphate(VI)/SO_4^{2-} salts **except** →	Barium(II) sulphate(VI)/$BaSO_4$ Calcium(II) sulphate(VI)/$CaSO_4$ Lead(II) sulphate(VI)/$PbSO_4$
All sulphate(IV)/SO_3^{2-} salts **except** →	Barium(II) sulphate(IV)/$BaSO_3$ Calcium(II) sulphate(IV)/$CaSO_3$ Lead(II) sulphate(IV)/$PbSO_3$
All chlorides/Cl^- **except** →	Silver chloride/AgCl Lead(II)chloride/$PbCl_2$(dissolves in hot water)
All phosphate(V)/PO_4^{3-}	
All sodium,potassium and ammonium salts	
All hydrogen carbonates/HCO_3^-	
All hydrogen sulphate(VI)/ HSO_4^-	
Sodium carbonate/Na_2CO_3, potassium carbonate/ K_2CO_3, ammonium carbonate $(NH_4)_2CO_3$ ←	**except** All carbonates
All alkalis(KOH,NaOH, NH_4OH) ←	**except** All bases

13 Salts can be prepared in a school laboratory by a method that uses its solubility in water.

(a) Soluble salts may be prepared by using any of the following methods:

(i)Direct displacement/reaction of a metal with an acid.

By reacting a metal higher in the reactivity series than hydrogen with a dilute acid,a salt is formed and hydrogen gas is evolved.

Excess of the metal must be used to ensure all the acid has reacted.

When effervescence/bubbling /fizzing has stopped ,excess metal is filtered.

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The filtrate is heated to concentrate then allowed to crystallize.
Washing with distilled water then drying between filter papers produces a sample crystal of the salt. i.e.

$M(s) + H_2X \rightarrow MX(aq) + H_2(g)$

Examples

$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$
$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$
$Pb(s) + 2HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + H_2(g)$
$Ca(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2(g)$
$Mg(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$
$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

(ii)Reaction of an insoluble base with an acid

By adding an insoluble base (oxide/hydroxide) to a dilute acid until **no more dissolves**, in the acid, a salt and water are formed. Excess of the base is filtered off. The filtrate is heated to concentrate, allowed to crystallize then washed with distilled water before drying between filter papers e.g.

$PbO(s) + 2HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + H_2O(l)$
$Pb(OH)_2(s) + 2HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + 2H_2O(l)$
$CaO(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l)$
$MgO(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2O(l)$
$MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$
$ZnO(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$
$Zn(OH)_2(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2H_2O(l)$
$CuO(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2O(l)$
$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$
$Ag_2O(s) + 2HNO_3(aq) \rightarrow 2AgNO_3(aq) + H_2O(l)$
$Na_2O(s) + 2HNO_3(aq) \rightarrow 2NaNO_3(aq) + H_2O(l)$

(iii) Reaction of insoluble /soluble carbonate /hydrogen carbonate with an acid.

By adding an excess of a soluble /insoluble carbonate or hydrogen carbonate to a dilute acid, effervescence /fizzing/bubbling out of carbon (IV) oxide gas shows the reaction is taking place. When effervescence /fizzing/bubbling out of the gas is over, excess of the insoluble carbonate is filtered off. The filtrate is heated to concentrate, allowed to crystallize then washed with distilled water before drying between filter paper papers e.g.

$PbCO_3(s) + 2HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + H_2O(l) + CO_2(g)$
$ZnCO_3(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2O(l) + CO_2(g)$
$CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l) + CO_2(g)$
$MgCO_3(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2O(l) + CO_2(g)$
$Cu\,CO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l) + CO_2(g)$
$Ag_2CO_3(s) + 2HNO_3(aq) \rightarrow 2AgNO_3(aq) + H_2O(l) + CO_2(g)$

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Na_2CO_3 (s) + $2HNO_3$(aq) -> $2NaNO_3$(aq) + H_2O (l) + CO_2(g)
K_2CO_3 (s) + 2HCl(aq) -> 2KCl(aq) + H_2O (l) + CO_2(g)
$NaHCO_3$ (s) + HNO_3(aq) -> $NaNO_3$(aq) + H_2O (l) + CO_2(g)
$KHCO_3$ (s) + HCl(aq) -> KCl(aq) + H_2O (l) + CO_2(g)

(iv) neutralization/reaction of soluble base/alkali with dilute acid

By adding an acid to a burette into a known volume of an alkali with 2-3 drops of an indicator, the colour of the indicator changes when the acid has completely reacted with an alkali at the **end point.** The procedure is then repeated without the indicator .The solution mixture is then heated to concentrate, allowed to crystallize ,washed with distilled water before drying with filter papers. e.g.

NaOH (aq) + HNO_3(aq) -> $NaNO_3$(aq) + H_2O (l)
KOH (aq) + HNO_3(aq) -> KNO_3(aq) + H_2O (l)
KOH (aq) + HCl(aq) -> KCl(aq) + H_2O (l)
2KOH (aq) + H_2SO_4(aq) -> K_2SO_4(aq) + $2H_2O$ (l)
2 NH_4OH (aq) + H_2SO_4(aq) -> $(NH_4)_2SO_4$(aq) + $2H_2O$ (l)
NH_4OH (aq) + HNO_3(aq) -> NH_4NO_3(aq) + H_2O (l)

(iv) Direct synthesis/combination.

When a metal **burn** in a gas jar containing a non metal , the two directly combine to form a salt. e.g.

2Na(s) + Cl_2(g) -> 2NaCl(s)
2K(s) + Cl_2(g) -> 2KCl(s)
Mg(s) + Cl_2(g) -> $Mg Cl_2$ (s)
Ca(s) + Cl_2(g) -> $Ca Cl_2$ (s)

Some salts once formed undergo **sublimation** and **hydrolysis**. Care should be taken to avoid water/moisture into the reaction flask during their preparation. Such salts include aluminum (III) chloride ($AlCl_3$) and iron (III) chloride ($FeCl_3$)

1. Heated aluminium foil reacts with chlorine to form aluminium(III)chloride that sublimes away from the source of heating then deposited as solid again

2Al(s) + $3Cl_2$(g) -> $2AlCl_3$ (**s/g**)

Once formed aluminium(III)chloride hydrolyses/reacts with water vapour / moisture present to form aluminium hydroxide solution and highly acidic fumes of hydrogen chloride gas.

$AlCl_3$(s)+ $3H_2O$(g) -> $Al(OH)_3$ (**aq**) + 3HCl(g)

2. Heated iron filings reacts with chlorine to form iron(**III**)chloride that sublimes away from the source of heating then deposited as solid again

2Fe(s) + $3Cl_2$(g) -> $2FeCl_3$ (**s/g**)

Once formed , aluminium(III)chloride hydrolyses/reacts with water vapour / moisture present to form aluminium hydroxide solution and highly acidic fumes of hydrogen chloride gas.

$FeCl_3$(s)+ $3H_2O$(g) -> $Fe(OH)_3$ (**aq**) + 3HCl(g)

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(b)**Insoluble** salts can be prepared by reacting **two** suitable **soluble** salts to form **one soluble** and **one insoluble**. This is called **double decomposition** or **precipitation**. The mixture is filtered and the **residue** is washed with distilled water then dried.

CuSO_4(aq) + Na_2**CO_3** (aq) -> **$CuCO_3$ (s)** + Na_2 SO_4(aq)
BaCl_2(aq) + K_2**SO_4** (aq) -> **$BaSO_4$ (s)** + 2KCl (aq)
Pb$(NO_3)_2$(aq) + K_2**SO_4** (aq) -> **$PbSO_4$ (s)** + 2KNO_3 (aq)
2**Ag**NO_3(aq) + Mg**Cl_2** (aq) -> 2**AgCl(s)** + $Mg(NO_3)_2$ (aq)
Pb$(NO_3)_2$(aq) + $(NH_4)_2$**SO_4** (aq) -> **$PbSO_4$ (s)** + 2NH_4NO_3(aq)
BaCl_2(aq) + K_2**SO_3** (aq) -> **$BaSO_3$ (s)** + 2KCl (aq)

14. Salts may lose their water of crystallization, decompose, melt or sublime on heating on a Bunsen burner flame.
The following shows the behavior of some salts on heating gently /or strongly in a laboratory school burner:

(a)effect of heat on chlorides

All chlorides have very high melting and boiling points and therefore are not affected by laboratory heating except ammonium chloride. Ammonium chloride **sublimes** on **gentle** heating. It **dissociate** into the constituent **ammonia** and **hydrogen chloride** gases on strong heating.

$NH_4Cl(s) \rightleftharpoons NH_4Cl(g) \rightleftharpoons NH_3(g) + HCl(g)$
(sublimation) (dissociation)

(b)effect of heat on nitrate(V)

(i) Potassium nitrate(V)/KNO_3 and sodium nitrate(V)/$NaNO_3$ decompose on heating to form Potassium nitrate(III)/KNO_2 and sodium nitrate(III)/$NaNO_2$ and producing Oxygen gas in each case.

2**KNO_3** (s) -> 2KNO_2(s) + O_2(g)
2**$NaNO_3$** (s) -> 2$NaNO_2$(s) + O_2(g)

(ii)Heavy metal nitrates(V) salts decompose on heating to form the oxide and a mixture of brown acidic nitrogen(IV)oxide and oxygen gases. e.g.

2$Ca(NO_3)_2$ (s) -> 2CaO(s) + 4NO_2(g) + O_2(g)
2$Mg(NO_3)_2$(s) -> 2MgO(s) + 4NO_2(g) + O_2(g)
2$Zn(NO_3)_2$(s) -> 2ZnO(s) + 4NO_2(g) + O_2(g)
2$Pb(NO_3)_2$(s) -> 2PbO(s) + 4NO_2(g) + O_2(g)
2$Cu(NO_3)_2$(s) -> 2CuO(s) + 4NO_2(g) + O_2(g)
2$Fe(NO_3)_2$(s) -> 2FeO(s) + 4NO_2(g) + O_2(g)

(iii)Silver(I)nitrate(V) and mercury(II) nitrate(V) are lowest in the reactivity series. They decompose on heating to form the **metal**(silver and mercury)and the Nitrogen(IV)oxide and oxygen gas. i.e.

2$AgNO_3$(s) -> 2Ag (s) + 2NO_2(g) + O_2(g)

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$2Hg(NO_3)_2(s) \rightarrow 2Hg(s) + 4NO_2(g) + O_2(g)$

(iv)Ammonium nitrate(V) and Ammonium nitrate(III) decompose on heating to Nitrogen(I)oxide(relights/rekindles glowing splint) and nitrogen gas respectively.Water is also formed.i.e.

$NH_4NO_3(s) \rightarrow N_2O(g) + H_2O(l)$

$NH_4NO_2(s) \rightarrow N_2(g) + H_2O(l)$

(c) effect of heat on nitrate(V)

Only Iron(II)sulphate(VI), Iron(III)sulphate(VI) and copper(II)sulphate(VI) decompose on heating. They form the **oxide**, and produce highly acidic fumes of acidic **sulphur(IV)oxide** gas.

$2FeSO_4(s) \rightarrow Fe_2O_3(s) + SO_3(g) + SO_2(g)$

$Fe_2(SO_4)_3(s) \rightarrow Fe_2O_3(s) + SO_3(g)$

$CuSO_4(s) \rightarrow CuO(s) + SO_3(g)$

(d) effect of heat on carbonates(IV) and hydrogen carbonate(IV).

(i)Sodium carbonate(IV)and potassium carbonate(IV)**do not decompose** on heating.

(ii)Heavy metal nitrate(IV)salts decompose on heating to form the **oxide** and produce **carbon(IV)oxide** gas. Carbon (IV)oxide gas forms a white precipitate when bubbled in lime water. The white precipitate dissolves if the gas is in excess.

e.g. $CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

$PbCO_3(s) \rightarrow PbO(s) + CO_2(g)$

$FeCO_3(s) \rightarrow FeO(s) + CO_2(g)$

$ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$

(iii)Sodium hydrogen carbonate(IV) and Potassium hydrogen carbonate(IV)decompose on heating to give the corresponding carbonate (IV) and form water and carbon(IV)oxide gas. i.e.

$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(l)$

$2KHCO_3(s) \rightarrow K_2CO_3(s) + CO_2(g) + H_2O(l)$

(iii) Calcium hydrogen carbonate (IV) and Magnesium hydrogen carbonate(IV) decompose on heating to give the corresponding carbonate (IV) and form water and carbon(IV)oxide gas. i. e.

$Ca(HCO_3)_2(\mathbf{aq}) \rightarrow CaCO_3(s) + CO_2(g) + H_2O(l)$

$Mg(HCO_3)_2(\mathbf{aq}) \rightarrow$ -

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CHEMISTRY Practice balancing Chemical equations

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Balance the following chemical equations in the spaces provided on the question paper.

$Ca(OH)_2(aq) + Cl_2(g) \rightarrow CaCl_2(aq) + CaOCl_2(aq) + H_2O(l)$
(Cold/dilute) (Calcium Chlorate(I))

$Ca(OH)_2(aq) + Cl_2(g) \rightarrow CaCl_2(aq) + Ca(ClO_3)_2(aq) + H_2O(l)$
(Hot/Concentrated) (Calcium Chlorate(V))
$NaOH + Cl_2(g) \rightarrow NaClO_3(aq) + NaCl(aq) + 3H_2O(l)$ (Sodium Chlorate (V)
$KOH + Cl_2(g) \rightarrow KClO_3(aq) + KCl(aq) + H_2O(l)$ (Potassium Chlorate (V)
$Ca(OH)_2(aq) + Cl_2(g) \rightarrow CaCl_2(aq) + CaOCl_2(aq) + H_2O(l)$
(Cold/dilute) (Calcium Chlorate(I))
$NaCl(s) + H_2SO_4(l) \rightarrow NaHSO_4(aq) + HCl(g)$
$KCl(s) + H_2SO_4(l) \rightarrow KHSO_4(aq) + HCl(g)$
$CaO(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2O(l)$
$Ca(OH)_2(aq) + Cl_2(g) \rightarrow CaOCl_2(aq) + H_2O(l)$
$HCl(s) + NH_3(g) \rightarrow NH_4Cl(s)$
$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$
$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
$2Li(s) + 2HCl(aq) \rightarrow 2LiCl(aq) + H_2(g)$

$Ca(OH)_2(aq) + Cl_2(g) \rightarrow CaCl_2(aq) + Ca(ClO_3)_2(aq) + H_2O(l)$
(Hot/Concentrated) (Calcium Chlorate(V))
$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$
(Colourless solution)

1.Chemical equation:
Ionic equation:
$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$

2.Chemical equation:
$Ag_2CO_3(s) + 2HCl(aq) \rightarrow 2AgCl(s) + H_2O(l) + CO_2(g)$
(Coats/Cover Ag_2CO_3)

3.Chemical equation:
$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$
(colourless solution)
Ionic equation:
$NaHCO_3(s) + H^+(aq) \rightarrow Na^+(aq) + H_2O(l) + CO_2(g)$

4.Chemical equation:
$CuCO_3(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2O(l) + CO_2(g)$

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(Blue Solution)

Ionic equation:

$CuCO_3(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l) + CO_2(g)$

5.Chemical equation:

$FeCO_3(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2O(l) + CO_2(g)$

(green solution)

Ionic equation:

$FeCO_3(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2O(l) + CO_2(g)$

6.Chemical equation:

$(NH_4)_2CO_3(s) + 2HCl(aq) \rightarrow 2NH_4Cl(aq) + H_2O(l) + CO_2(g)$

Ionic equation:

$(NH_4)_2CO_3(s) + 2H^+(aq) \rightarrow 2NH_4^+(aq) + H_2O(l) + CO_2(g)$

Chemical equation:

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

Ionic equation:

$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$

Chemical equation:

$KOH(aq) + HCl(aq) \rightarrow KCl(aq) + H_2O(l)$

Ionic equation:

$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$

Chemical equation:

$NH_4OH(aq) + HCl(aq) \rightarrow NH_4Cl(aq) + H_2O(l)$

Ionic equation:

$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$

$CuO(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l)$

Ionic equation:

$CuO(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l)$

Chemical equation:

$CaO(s) + HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l)$

Ionic equation:

$CuO(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l)$

Chemical equation:

$PbO(s) + 2HCl(aq) \rightarrow PbCl_2(s) + H_2O(l)$

No ionic equation

Chemical equation:

$ZnO(s) + HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$

Ionic equation:

$ZnO(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2O(l)$

$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

$HCl(g) + (aq) \rightarrow HCl(aq)$

$NH_4Cl(s) \rightarrow HCl(g) + NH_3(g)$

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CHEMISTRY Form 2: Practice 1: Time 2 hours 80 marks

1. Study the information in the table below and answer the questions that follow (The letters do not represent the actual symbols of the elements)

		Ionization Energy_kJ/Mole	
Element	Electronic configuration	1st ionization energy	2nd ionization energy
A	2.2	900	1800
B	2.8.2	736	1450
C	2.8.8.2	590	1150

(i) What chemical family do the elements A, B and C belong? (1mk)
(ii)Write the formula and electronic structure of an ion of B(2mks)
Formula
Electronic structure
(iii)What type of bonding exist in(2mks)
I. atoms of C
II. chloride of B
(iv)What is ionization energy(1mk)
(v)Explain the following:
1.The 1st ionization energy is lower that the second ionization energy.(2mk)
2.The 1st ionization energy of B is lower that of C.(2mk)
(vi)Write a chemical equation for the reaction of element B with:
1.Air
2.Chlorine gas
3.Steam(water vapour)

2.Study the information in the table below and answer the questions that follow (the letters do not represent the actual symbol of the substances)

Substance	**Melting Point (^{0}C)**	**Boiling Point (^{0}C)**	**Solubility in water**	**Density at room temperature g/cm^3**
H	-117	78.5	Very soluble	0.8
J	-78	-33	Very soluble	0.77×10^{-3}
K	-23	77	Insoluble	1.6

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L	-219	-183	Slightly soluble	1.33×10^{-3}

(i) Which substance would dissolve in water and could be separated from the solution by fractional distillation? Give a reason (2mk)

(ii) Which substances is a liquid at room temperature and when mixed with water two layers would be formed? Explain (2mk)

(iii) Which letter represents a substance that is gas at room temperature and which can be collected:

I. Over water? Explain (2mk)

II. By downward displacement of air? (Density of air is $1.29 \times 10^{-3} g/cm^3$ at room temperature).

Explain (2mk)

3. The grid below represents part of the periodic table. The letters do not represent the actual symbols.

								A
B			X	G	P	Z	E	V
	J		I	L	R	T		
D	G						M	

a) Select the most reactive

(i)non-metal. (1mk)

(ii)metal. (1mk)

b) Write the formula of the compound consisting of (10mk)

1.D and Z only.

2 . X and Z only

3. Oxide of B

4. Carbonate of J

5. sulphate of D

6. Nitrate of B

7. Chloride of X

8. Sodium compound of E

9. Aluminium compound of Z

10. Hydrogen compound of G

c) Select an element that can form an ion of charge (10mk)

(i) +1

(ii) -1

(iii) +2

(iv) +3

(v) -3

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d) Which element has the least ionization energy? Explain (2mks)

f) To which chemical family do the following elements belong? (3mk)

J

E

B

(g)When a piece of element G is placed in cold water, it sinks to the bottom and effervescence of a colourless gas that burns explosively is produced. Use a simple diagram to illustrate how this gas can be collected during this experiment. (3mks)

h) An element K has relative atomic mass of 40.2.It has two isotopes of masses 39 and 42. Calculate the relative abundance of each isotope. (3mks)

4. Balance the following chemical equation (6mk)

1.$Ca(OH)_2(aq) + Cl_2(g) \rightarrow CaCl_2(aq) + CaOCl_2(aq) + H_2O(l)$

2.$NaOH + Cl_2(g) \rightarrow NaClO_3(aq) + NaCl(aq) + 3H_2O(l)$

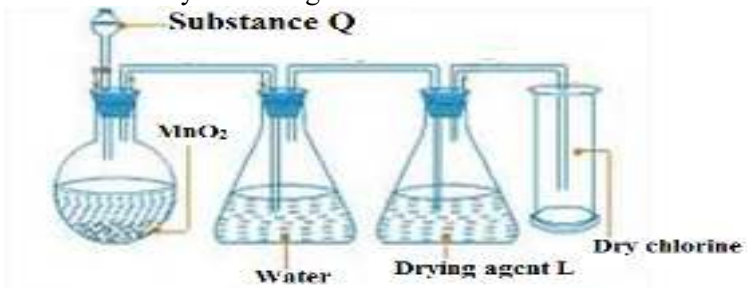
3.$NaCl(s) + H_2SO_4(l) \rightarrow NaHSO_4(aq) + HCl(g)$

4.$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2O(l)$

5.$Fe(s) + HCl(aq) \rightarrow FeCl_3(aq) + H_2(g)$

6.$Zn(s) + HCl(aq) \rightarrow ZnCl_2(aq + H_2(g)$

5.The diagram below shows a set up of apparatus for the school laboratory collection of dry chlorine gas.



A) Name (2mk)

(i) Substance Q

(ii) Suitable drying agent L

b) State a missing condition for the reaction to take place faster. (1mk)

c) Moist red and blue litmus papers were dipped into the chlorine gas from the above set up .State and explain the observations made. (2mk)

d) Write the equation for the reaction taking place in the conical flask (1mk)

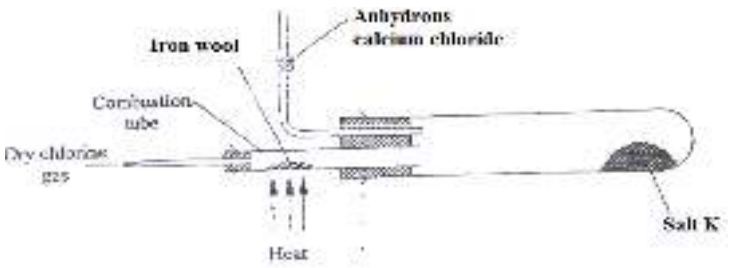
e) Name two other substances that can be used in place of MnO_2 (2mk)

(f)State three uses of chlorine (3mk)

6. Study the set up below.

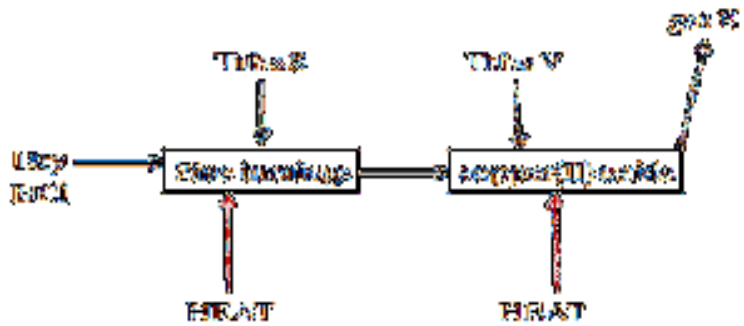
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a) Name salt K (1mk)
(b)Write the equation for the reaction for the formation of salt K (1mk)
(c)What property of salt A is exhibited as shown in the experiment.(1mk)
(d)What is the purpose of anhydrous calcium chloride? Explain (2mk)
(e)Name another metal that can be used to produce similar results (1mk)

7. In an experiment, dry hydrogen chloride gas was passed through heated zinc turnings as in the set up below. The gas produced was the passed through copper (II) oxide



Write the equation for the reaction :
(i) for the preparation of hydrogen chloride gas. (1mk)
(ii)in tube S(1mk)
b)State and explain the observation made in tube V. (2mk)
(c)How would the total mass of tube S and tube V and their contents compare before and after the experiment.
Tube S(2mk)
Tube V(2mk)
(d)Gas K was condensed to liquid K.
(i)Identify liquid K(1mk)
(ii)Describe a simple chemical test to identify Liquid K(3mk)

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(iii)A small piece of sodium metal was placed into a beaker containing liquid K.

I. State three observations made(3mk)

II.Write an equation for the reaction that take place(1mk)

III.What is the pH of the resulting solution. Explain(2mk)

8.Using dot(.) and cross(x) to represent electrons, show the bonding in.

(a)hydroxonium ion(H_3O^+)(2mk)

(b)Carbon(IV)oxide(CO_2)(2mk)

(c)Carbon(II)oxide(CO)(2mk)

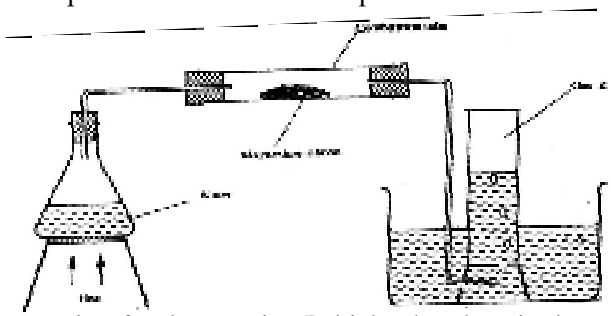
(d)Ammonia(NH_3)(2mk)

(e)Ammonium ion(NH_4^+)(2mk)

(f)Magnesium chloride($MgCl_2$)(2mk)

(g)Ethane(C_2H_6)(2mk)

9.Study the set- up below and answer the questions that follow



(a) Write an equation for the reaction, which take place in the combustion tube.

(b) What property of gas Z allows it to be collected as shown in the diagram

(c) State two uses of gas Z

CHEMISTRY FORM THREE NOTES

CHEMISTRY OF CARBON

CHEMISTRY OF CARBON

A: CARBON

Carbon is an element in Group IV(Group 4)of the Periodic table .It has atomic number 6 and electronic configuration 2:4 and thus has four valence electrons(tetravalent).It does not easily ionize but forms strong covalent bonds with other elements including itself.

(a)Occurrence

Carbon mainly naturally occurs as:

(i)allotropes of carbon i.e graphite, diamond and fullerenes.

(ii)amorphous carbon in coal, peat ,charcoal and coke.

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(iii)carbon(IV)oxide gas accounting 0.03% by volume of normal air in the atmosphere.

(b)Allotropes of Carbon

Carbon naturally occur in two main crystalline allotropic forms, carbon-graphite and carbon-diamond

Carbon-diamond	Carbon-graphite
Shiny crystalline solid	Black/dull crystalline solid
Has a very high melting/boiling point because it has a very closely packed giant tetrahedral structure joined by strong covalent bonds	Has a high melting/boiling point because it has a very closely packed giant hexagonal planar structure joined by strong covalent bonds
Has very high density(Hardest known natural substance)	Soft
Abrassive	Slippery
Poor electrical conductor because it has no free delocalized electrons	Good electrical conductor because it has free 4th valency delocalized electrons
Is used in making Jewels, drilling and cutting metals	Used in making Lead-pencils,electrodes in batteries and as a lubricant
Has giant tetrahedral structure	Has giant hexagonal planar structure

c) Properties of Carbon

(i)Physical properties of carbon

Carbon occur widely and naturally as a black solid

It is insoluble in water but soluble in carbon disulphide and organic solvents.

It is a poor electrical and thermal conductor.

(ii)Chemical properties of carbon

I. Burning

Experiment

Introduce a small piece of charcoal on a Bunsen flame then lower it into a gas jar containing Oxygen gas. Put three drops of water. Swirl. Test the solution with blue and red litmus papers.

Observation

-Carbon chars then burns with a blue flame

-Colourless and odourless gas produced

-Solution formed turn blue litmus paper faint red. Red litmus paper remains red.

Explanation

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Carbon burns in air and faster in Oxygen with a blue non-sooty/non-smoky flame forming Carbon (IV) oxide gas. Carbon burns in limited supply of air with a blue non-sooty/non-smoky flame forming Carbon (IV) oxide gas. Carbon (IV) oxide gas dissolve in water to form weak acidic solution of Carbonic (IV)acid.

Chemical Equation

$C(s) + O_2(g) \rightarrow CO_2(g)$ (in excess air)

$2C(s) + O_2(g) \rightarrow 2CO(g)$ (in limited air)

$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$ (very weak acid)

II. Reducing agent

Experiment

Mix thoroughly equal amounts of powdered charcoal and copper (II)oxide into a crucible. Heat strongly.

Observation

Colour change from black to brown

Explanation

Carbon is a reducing agent. For ages it has been used to reducing metal oxide ores to metal, itself oxidized to carbon (IV) oxide gas. Carbon reduces black copper (II) oxide to brown copper metal

Chemical Equation

$2CuO(s) + C(s) \rightarrow 2Cu(s) + CO_2(g)$
(black) (brown)

$2PbO(s) + C(s) \rightarrow 2Pb(s) + CO_2(g)$
(brown when hot/ yellow when cool) (grey)

$2ZnO(s) + C(s) \rightarrow 2Zn(s) + CO_2(g)$
(yellow when hot/ white when cool) (grey)

$Fe_2O_3(s) + 3C(s) \rightarrow 2Fe(s) + 3CO_2(g)$
(brown when hot/cool) (grey)

$Fe_3O_4(s) + 4C(s) \rightarrow 3Fe(s) + 4CO_2(g)$
(brown when hot/cool) (grey)

B: COMPOUNDS OF CARBON

The following are the main compounds of Carbon

(i)Carbon(IV)Oxide(CO_2)
(ii)Carbon(II)Oxide(CO)
(iii)Carbonate(IV) (CO_3^{2-})and hydrogen carbonate(IV(HCO_3^-)
(iv)Sodium carbonate(Na_2CO_3)
(i) Carbon(IV)Oxide (CO_2)

(a)Occurrence

Carbon(IV)oxide is found:
-in the air /atmosphere as 0.03% by volume.

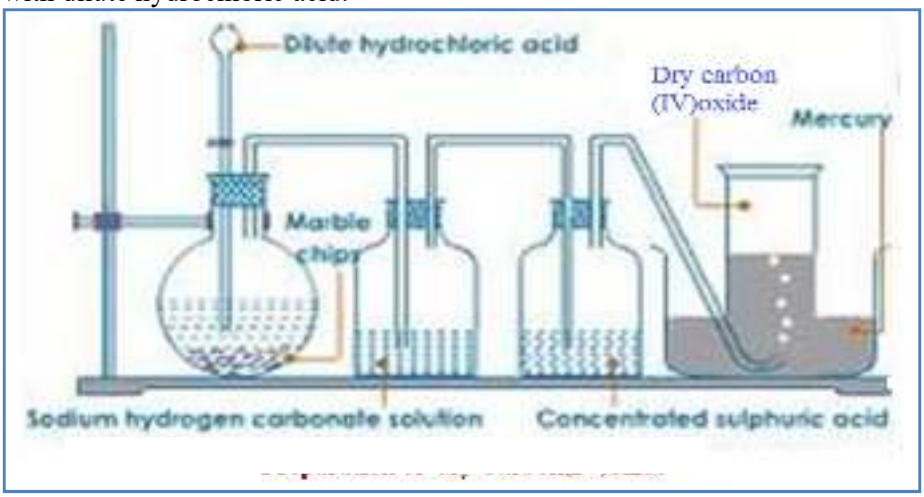
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-a solid carbon(IV)oxide mineral in Esageri near Eldame Ravine and Kerita near Limuru in Kenya.

(b)School Laboratory preparation

In the school laboratory carbon(IV)oxide can be prepared in the school laboratory from the reaction of marble chips($CaCO_3$)or sodium hydrogen carbonate($NaHCO_3$) with dilute hydrochloric acid.



(c)Properties of carbon(IV)oxide gas(Questions)

1.Write the equation for the reaction for the school laboratory preparation of carbon (IV)oxide gas.

Any carbonate reacted with dilute hydrochloric acid should be able to generate carbon (IV)oxide gas.

Chemical equations

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCO_3(aq) + H_2O(l) + CO_2(g)$

$ZnCO_3(s) + 2HCl(aq) \rightarrow ZnCO_3(aq) + H_2O(l) + CO_2(g)$

$MgCO_3(s) + 2HCl(aq) \rightarrow MgCO_3(aq) + H_2O(l) + CO_2(g)$

$CuCO_3(s) + 2HCl(aq) \rightarrow CuCO_3(aq) + H_2O(l) + CO_2(g)$

$NaHCO_3(s) + HCl(aq) \rightarrow Na_2CO_3(aq) + H_2O(l) + CO_2(g)$

$KHCO_3(s) + HCl(aq) \rightarrow K_2CO_3(aq) + H_2O(l) + CO_2(g)$

2.What method of gas collection is used in preparation of Carbon(IV)oxide gas. Explain.

Downward delivery /upward displacement of air/over mercury

Carbon(IV)oxide gas is about 1½ times denser than air.

3.What is the purpose of :

(a)water?

To absorb the more volatile hydrogen chloride fumes produced during the vigorous reaction.

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(b)sodium hydrogen carbonate?
To absorb the more volatile hydrogen chloride fumes produced during the vigorous reaction and by reacting with the acid to produce more carbon (IV)oxide gas .
Chemical equation
$NaHCO_3(s) + HCl(aq) \rightarrow Na_2CO_3(aq) + H_2O(l) + CO_2(g)$
(c)concentrated sulphuric(VI)acid?
To dry the gas/as a drying agent
4.Describe the smell of carbon(IV)oxide gas
Colourless and odourless
5. Effect on lime water.
Experiment
Bubbled carbon(IV)oxide gas into a test tube containing lime water for about three minutes
Observation
White precipitate is formed.
White precipitate dissolved when excess carbon(IV)oxide gas is bubbled .
Explanation
Carbon(IV)oxide gas reacts with lime water($Ca(OH)_2$) to form an insoluble white precipitate of calcium carbonate. Calcium carbonate reacts with more Carbon(IV) oxide gas to form soluble Calcium hydrogen carbonate.
Chemical equation
$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$
$CaCO_3(aq) + H_2O(l) + CO_2(g) \rightarrow Ca(HCO_3)_2(aq)$
6. Effects on burning Magnesium ribbon
Experiment
Lower a piece of burning magnesium ribbon into a gas jar containing carbon (IV)oxide gas.
Observation
The ribbon continues to burn with difficulty
White ash/solid is formed.
Black speck/solid/particles formed on the side of gas jar.
Explanation
Carbon (IV) oxide gas does not support combustion/burning.Magnesium burn to produce/release enough heat energy to decompose Carbon (IV) oxide gas to carbon and oxygen.Magnesium continues to burn in Oxygen forming white Magnesium Oxide solid/ash.Black speck/particle of carbon/charcoal residue forms on the sides of reaction flask. During the reaction Carbon (IV) oxide is reduced (Oxidizing agent)to carbon while Magnesium is Oxidized to Magnesium Oxide.
Chemical equation
$2Mg(s) + CO_2(g) \rightarrow C(s) + 2MgO(l)$

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7. Dry and wet litmus papers were separately put in a gas jar containing dry carbon (IV) oxide gas. State and explain the observations made.

Observation

Blue dry litmus paper remain blue

Red dry litmus paper remain Red

Blue wet/damp/moist litmus paper turns red

Red wet/damp/moist litmus paper remain red

Explanation

Dry Carbon (IV) oxide gas is a molecular compound that does not dissociate/ionize to release $\mathbf{H^+}$ and thus has no effect on litmus papers.

Wet/damp/moist litmus paper contains water that dissolves/react with dry carbon (IV) oxide gas to form the weak solution of carbonic (IV) acid (H_2CO_3).

Carbonic (IV) acid dissociate/ionizes to a few /little free $\mathbf{H^+}$ and $\mathbf{CO_3^{2-}}$.

The few H^+ (aq) ions are responsible for turning blue litmus paper to faint red showing the gas is very weakly acidic.

Chemical equation

$H_2CO_3(aq) \rightarrow 2H^+(aq) + CO_3^{2-}(aq)$

8. Explain why Carbon (IV) oxide cannot be prepared from the reaction of:

(i) Marble chips with dilute sulphuric (VI) acid.

Explanation

Reaction forms insoluble calcium sulphate (VI) that cover/coat unreacted marble chips stopping further reaction

Chemical equation

$CaCO_3(s) + H_2SO_4(aq) \rightarrow \mathbf{CaSO_4(s)} + H_2O(l) + CO_2(g)$

$PbCO_3(s) + H_2SO_4(aq) \rightarrow \mathbf{PbSO_4(s)} + H_2O(l) + CO_2(g)$

$BaCO_3(s) + H_2SO_4(aq) \rightarrow \mathbf{BaSO_4(s)} + H_2O(l) + CO_2(g)$

(ii) Lead (II) carbonate with dilute Hydrochloric acid.

Reaction forms insoluble Lead (II)Chloride that cover/coat unreacted Lead(II) carbonate stopping further reaction unless the reaction mixture is heated. Lead (II) Chloride is soluble in hot water.

Chemical equation

$PbCO_3(s) + 2HCl(aq) \rightarrow \mathbf{PbCl_2(s)} + H_2O(l) + CO_2(g)$

9. Describe the test for the presence of Carbon (IV) oxide.

Using burning splint

Lower a burning splint into a gas jar suspected to contain Carbon (IV) oxide gas. The burning splint is extinguished.

Using Lime water.

Bubble the gas suspected to be Carbon (IV) oxide gas. A white precipitate that dissolves in excess bubbling is formed.

Chemical equation

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

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$CaCO_3$ (aq) + H_2O(l) + CO_2 (g) -> $Ca(HCO_3)_2$ (aq)

10.State three main uses of Carbon (IV)oxide gas

(i)In the Solvay process for the manufacture of soda ash/sodium carbonate

(ii)In preservation of aerated drinks

(iii)As fire extinguisher because it does not support combustion and is denser than air.

(iv)In manufacture of Baking powder.

(ii) Carbon (II) Oxide (CO)

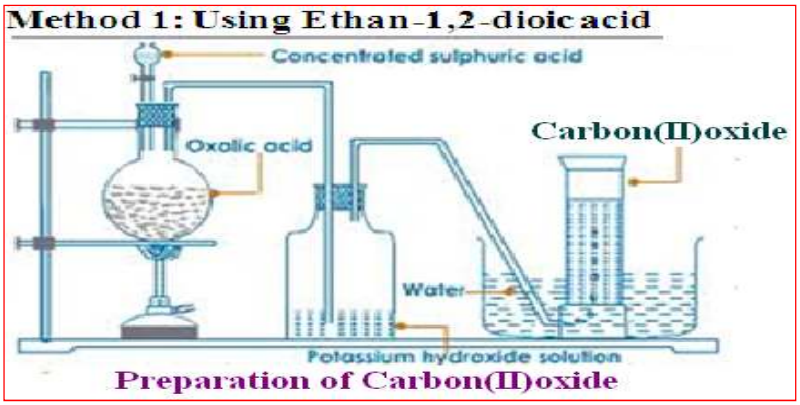
(a)Occurrence

Carbon (II) oxide is found is found from incomplete combustion of fuels like petrol charcoal, liquefied Petroleum Gas/LPG.

(b)School Laboratory preparation

In the school laboratory carbon(II)oxide can be prepared from dehydration of methanoic acid/Formic acid(HCOOH) or Ethan-1,2-dioic acid/Oxalic acid(HOOCCOOH) using concentrated sulphuric(VI) acid. Heating is necessary.

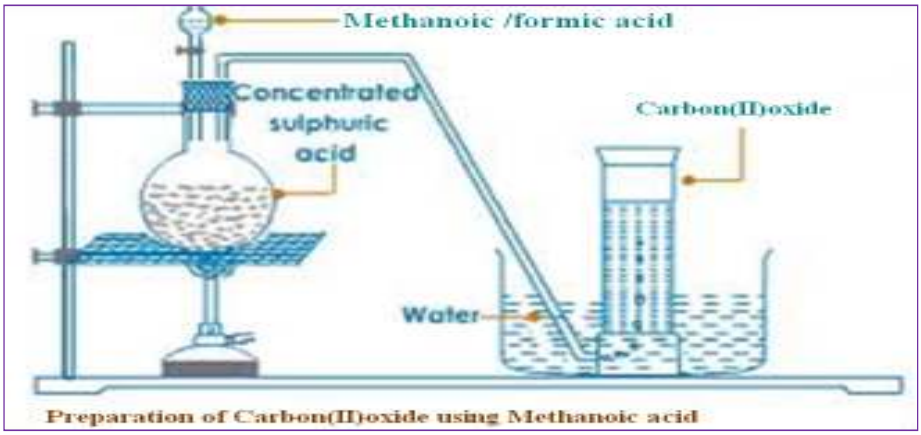
METHOD 1:Preparation of Carbon (IV)Oxide from dehydration of Oxalic/ethan-1,2-dioic acid



METHOD 2:Preparation of Carbon (IV)Oxide from dehydration of Formic/Methanoic acid

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Preparation of Carbon(II)oxide using Methanoic acid

(c)Properties of Carbon (II) Oxide(Questions)

1. Write the equation for the reaction for the preparation of carbon(II)oxide using;

(i)Method 1;

Chemical equation

$HOOCCOOH(s) \xrightarrow{-Conc.H_2SO_4} CO(g) + CO_2(g) + H_2O(l)$

$H_2C_2O_4(s) \xrightarrow{-Conc.H_2SO_4} CO(g) + CO_2(g) + H_2O(l)$

(ii)Method 2;

Chemical equation

$HCOOH(s) \xrightarrow{-Conc.H_2SO_4} CO(g) + H_2O(l)$

$H_2CO_2(s) \xrightarrow{-Conc.H_2SO_4} CO(g) + H_2O(l)$

2. What method of gas collection is used during the preparation of carbon (II) oxide

Over water because the gas is insoluble in water.

Downward delivery because the gas is 1 ½ times denser than air .

3. What is the purpose of :

(i) Potassium hydroxide/sodium hydroxide in Method 1

To absorb/ remove carbon (II) oxide produced during the reaction.

$2KOH(aq) + CO_2(g) \rightarrow K_2CO_3(s) + H_2O(l)$

$2NaOH(aq) + CO_2(g) \rightarrow Na_2CO_3(s) + H_2O(l)$

(ii) Concentrated sulphuric(VI)acid in Method 1 and 2.

Dehydrating agent –removes the element of water (Hydrogen and Oxygen in ratio 2:1) present in both methanoic and ethan-1,2-dioic acid.

4. Describe the smell of carbon (II) oxide.

Colourless and odourless.

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5. State and explain the observation made when carbon(IV)oxide is bubbled in lime water for a long time.

No white precipitate is formed.

6. Dry and wet/moist/damp litmus papers were separately put in a gas jar containing dry carbon (IV) oxide gas. State and explain the observations made.

Observation

-blue dry litmus paper remains blue

-red dry litmus paper remains red

- wet/moist/damp blue litmus paper remains blue

- wet/moist/damp red litmus paper remains red

Explanation

Carbon(II)oxide gas is a molecular compound that does not dissociate /ionize to release H+ ions and thus has no effect on litmus papers. Carbon(II)oxide gas is therefore a **neutral** gas.

7. Carbon (II) oxide gas was ignited at the end of a generator as below.



(i)State the observations made in flame K.

Gas burns with a blue flame

(ii)Write the equation for the reaction taking place at flame K.

$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$

8. Carbon (II) oxide is a reducing agent. Explain

Experiment

Pass carbon (II) oxide through glass tube containing copper (II) oxide. Ignite any excess poisonous carbon (II) oxide.

Observation

Colour change from black to brown. Excess carbon (II) oxide burn with a blue flame.

Explanation

Carbon is a reducing agent. It is used to reduce metal oxide ores to metal, itself oxidized to carbon (IV) oxide gas. Carbon (II) Oxide reduces black copper (II) oxide to brown copper metal

Chemical Equation

$CuO(s) + CO(g) \rightarrow Cu(s) + CO_2(g)$

(black) (brown)

$PbO(s) + CO(g) \rightarrow Pb(s) + CO_2(g)$

(brown when hot/ (grey)

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yellow when cool)

$ZnO(s) + CO(g) \rightarrow Zn(s) + CO_2(g)$

(yellow when hot/ white when cool) (grey)

$Fe_2O_3(s) + 3CO(s) \rightarrow 2Fe(s) + 3CO_2(g)$

(brown when hot/cool (grey)

$Fe_3O_4(s) + 4CO(g) \rightarrow 3Fe(s) + 4CO_2(g)$

(brown when hot/cool (grey)

These reactions are used during the extraction of many metals from their ore.

9. Carbon (II) oxide is a pollutant. Explain.

Carbon (II) oxide is highly poisonous/toxic.It preferentially combine with haemoglobin to form stable carboxyhaemoglobin in the blood instead of oxyhaemoglobin.This reduces the free haemoglobin in the blood causing nausea, coma then death.

10. The diagram below show a burning charcoal stove/burner/jiko. Use it to answer the questions that follow.



Explain the changes that take place in the burner

Explanation

Charcoal stove has air holes through which air enters. Air oxidizes carbon to carbon (IV) oxide gas at region I. This reaction is exothermic ($-\Delta H$) producing more heat.

Chemical equation

$C(s) + O_2(g) \rightarrow CO_2(g)$

Carbon (IV) oxide gas formed rises up to meet more charcoal which reduces it to Carbon (II) oxide gas.

Chemical equation

$2CO_2(g) + O_2(g) \rightarrow 2CO(g)$

At the top of burner in region II, Carbon (II)oxide gas is further oxidized to Carbon(IV)oxide gas if there is plenty of air but escape if the air is limited poisoning the living things around.

Chemical equation

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$2CO\ (g) + O_2(g) \rightarrow 2CO_2\ (g)$
(excess air)

11. Describe the test for the presence of carbon(II)oxide gas.

Experiment

Burn/Ignite the pure sample of the gas. Pass/Bubble the products into lime water/Calcium hydroxide.

Observation

Colourless gas burns with a blue flame. A white precipitate is formed that dissolve on further bubbling of the products.

Chemical equation

$2CO\ (g) + O_2(g) \rightarrow 2CO_2\ (g)$ (gas burns with blue flame)

Chemical equation

$Ca(OH)_2\ (aq) + CO_2\ (g) \rightarrow CaCO_3\ (s) + H_2O(l)$

Chemical equation

$CO_2\ (g) + CaCO_3\ (s) + H_2O(l) \rightarrow Ca(HCO_3)_2\ (aq)$

12. State the main uses of carbon (II)oxide gas.

(i) As a fuel /water gas

(ii)As a reducing agent in the blast furnace for extracting iron from iron ore(Magnetite/Haematite)

(iii)As a reducing agent in extraction of Zinc from Zinc ore/Zinc blende

(iv) As a reducing agent in extraction of Lead from Lead ore/Galena

(v) As a reducing agent in extraction of Copper from Copper iron sulphide/Copper pyrites.

(iii)Carbonate(IV) (CO_3^{2-})and hydrogen carbonate(IV(HCO_3^-)

1.Carbonate (IV) (CO_3^{2-}) are normal salts derived from carbonic(IV)acid (H_2CO_3) and hydrogen carbonate (IV) (HCO_3^-) are acid salts derived from carbonic(IV)acid. Carbonic(IV)acid(H_2CO_3) is formed when carbon(IV)oxide gas is bubbled in water. It is a dibasic acid with two ionizable hydrogens.

$H_2CO_3(aq) \rightarrow 2H^+(aq) + CO_3^{2-}(aq)$

$H_2CO_3(aq) \rightarrow H^+(aq) + HCO_3^-(aq)$

2.Carbonate (IV) (CO_3^{2-}) are insoluble in water **except** Na_2CO_3 , K_2CO_3 and $(NH_4)_2CO_3$

3.Hydrogen carbonate (IV) (HCO_3^-) are soluble in water. Only five hydrogen carbonates exist. **Na HCO_3 , KHCO$_3$,NH_4HCO_3 Ca$(HCO_3)_2$ and Mg$(HCO_3)_2$** Ca$(HCO_3)_2$ and Mg$(HCO_3)_2$ exist only in aqueous solutions.

3.The following experiments show the effect of heat on Carbonate (IV) (CO_3^{2-}) and Hydrogen carbonate (IV) (HCO_3^-) salts:

Experiment

In a clean dry test tube place separately about 1.0 of the following:

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Zinc(II)carbonate(IV), sodium hydrogen carbonate(IV), sodium carbonate(IV), Potassium carbonate(IV) ammonium carbonate(IV), potassium hydrogen carbonate(IV), Lead(II)carbonate(IV), Iron(II)carbonate(IV), and copper(II)carbonate(IV). Heat each portion gently the strongly. Test any gases produced with lime water.

Observation

(i)Colorless droplets form on the cooler parts of test tube in case of sodium carbonate(IV) and Potassium carbonate(IV).

(ii)White residue/solid left in case of sodium hydrogen carbonate(IV), sodium carbonate(IV), Potassium carbonate(IV) and potassium hydrogen carbonate(IV).

(iii)Colour changes from blue/green to black in case of copper(II)carbonate(IV).

(iv) Colour changes from green to brown/yellow in case of Iron (II)carbonate(IV).

(v) Colour changes from white when cool to yellow when hot in case of Zinc (II) carbonate(IV).

(vi) Colour changes from yellow when cool to brown when hot in case of Lead (II) carbonate(IV).

(vii)Colourless gas produced that forms a white precipitate with lime water in all cases.

Explanation

1. Sodium carbonate(IV) and Potassium carbonate(IV) exist as hydrated salts with 10 molecules of water of crystallization that condenses and collects on cooler parts of test tube as a colourless liquid.

Chemical equation

$Na_2CO_3 .10H_2O(s) \rightarrow Na_2CO_3 (s) + 10H_2O(l)$

$K_2CO_3 .10H_2O(s) \rightarrow K_2CO_3 (s) + 10H_2O(l)$

2. Carbonate (IV) (CO_3^{2-}) and Hydrogen carbonate (IV) (HCO_3^-) salts decompose on heating except Sodium carbonate(IV) and Potassium carbonate(IV).

(a) Sodium hydrogen carbonate(IV) and Potassium hydrogen carbonate(IV) decompose on heating to form sodium carbonate(IV) and Potassium carbonate(IV).Water and carbon(IV)oxide gas are also produced.

Chemical equation

$2NaHCO_3 (s) \rightarrow Na_2CO_3 (s) + H_2O(l) + CO_2 (g)$
(white) (white)

$2KHCO_3 (s) \rightarrow K_2CO_3 (s) + H_2O(l) + CO_2 (g)$
(white) (white)

(b) Calcium hydrogen carbonate(IV) and Magnesium hydrogen carbonate(IV) decompose on heating to form insoluble Calcium carbonate(IV) and Magnesium carbonate(IV).Water and carbon(IV)oxide gas are also produced.

Chemical equation

$Ca(HCO_3)_2 (aq) \rightarrow CaCO_3 (s) + H_2O(l) + CO_2 (g)$
(Colourless solution) (white)

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$Mg(HCO_3)_2$ (aq) -> $MgCO_3$ (s) + H_2O(l) + CO_2 (g)
(Colourless solution) (white)

(c) Ammonium hydrogen carbonate(IV) decompose on heating to form ammonium carbonate(IV) .Water and carbon(IV)oxide gas are also produced.

Chemical equation

$2NH_4HCO_3$ (s) -> $(NH_4)_2CO_3$ (s) + H_2O(l) + CO_2 (g)
(white) (white)

(d)All other carbonates decompose on heating to form the metal oxide and produce carbon(IV)oxide gas e.g.

Chemical equation

$MgCO_3$ (s) -> MgO (s) + CO_2 (g)
(white solid) (white solid)

Chemical equation

$BaCO_3$ (s) -> BaO (s) + CO_2 (g)
(white solid) (white solid)

Chemical equation

$CaCO_3$ (s) -> CaO (s) + CO_2 (g)
(white solid) (white solid)

Chemical equation

$CuCO_3$ (s) -> CuO (s) + CO_2 (g)
(blue/green solid) (black solid)

Chemical equation

$ZnCO_3$ (s) -> ZnO (s) + CO_2 (g)
(white solid) (white solid when cool/
Yellow solid when hot)

Chemical equation

$PbCO_3$ (s) -> PbO (s) + CO_2 (g)
(white solid) (yellow solid when cool/
brown solid when hot)

4.The following experiments show the presence of Carbonate (IV) (CO_3^{2-}) and Hydrogen carbonate (IV) (HCO_3^-) ions in sample of a salt:

(a)Using Lead(II) nitrate(V)

I. Using a portion of salt solution in a test tube .add four drops of Lead(II)nitrate(V)solution.Preserve.

Observation	inference
White precipitate/ppt	CO_3^{2-} ,SO_3^{2-} ,SO_4^{2-} ,Cl^-

II. To the preserved solution, add six drops of dilutte nitric(V)acid. Preserve.

Observation	inference
White precipitate/ppt persists	SO_4^{2-} ,Cl^-

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White precipitate/ppt dissolves	CO_3^{2-} ,SO_3^{2-}

II. To the preserved sample(that forms a precipitate),heat to boil.

Observation	inference
White precipitate/ppt persists	SO_4^{2-}
White precipitate/ppt dissolves	Cl^-

II. To the preserved sample(that do not form a precipitate),add three drops of acidified potassium manganate(VII)/lime water

Observation	inference
Effervescence/bubbles/fizzing colourless gas produced Acidified $KMnO_4$ decolorized/no white precipitate on lime water	SO_3^{2-}
Effervescence/bubbles/fizzing colourless gas produced Acidified $KMnO_4$ not decolorized/ white precipitate on lime water	CO_3^{2-}

Experiments/Observations:

(b)Using Barium(II)nitrate(V)/ Barium(II)chloride

I. To about 5cm3 of a salt solution in a test tube add four drops of Barium(II) nitrate (V) / Barium(II)chloride. Preserve.

Observation	Inference
White precipitate/ppt	SO_4^{2-} , SO_3^{2-} , CO_3^{2-} ions

II. To the preserved sample in (I) above, add six drops of 2M nitric(V) acid . Preserve.

Observation 1

Observation	Inference
White precipitate/ppt persists	SO_4^{2-} , ions

Observation 2

Observation	Inference
White precipitate/ppt dissolves	SO_3^{2-} , CO_3^{2-} , ions

III.To the preserved sample observation 2 in (II) above, add 4 drops of acidified potassium manganate(VII) /dichromate(VI).

Observation 1

Observation	Inference
(i)acidified potassium manganate(VII)decolorized	SO_3^{2-} ions

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(ii)Orange colour of acidified potassium dichromate(VI) turns to green	

Observation 2

Observation	Inference
(i)acidified potassium manganate(VII) not decolorized (ii)Orange colour of acidified potassium dichromate(VI) does not turns to green	CO_3^{2-} ions

Explanations

Using Lead(II)nitrate(V)

(i)Lead(II)nitrate(V) solution reacts with chlorides(Cl^-), Sulphate (VI) salts (SO_4^{2-}), Sulphate (IV)salts (SO_3^{2-}) and carbonates(CO_3^{2-}) to form the insoluble white precipitate of Lead(II)chloride, Lead(II)sulphate(VI), Lead(II) sulphate (IV) and Lead(II)carbonate(IV).

Chemical/ionic equation:

$Pb^{2+}(aq) + Cl^-(aq) \rightarrow PbCl_2(s)$

$Pb^{2+}(aq) + SO_4^{2+}(aq) \rightarrow PbSO_4(s)$

$Pb^{2+}(aq) + SO_3^{2+}(aq) \rightarrow PbSO_3(s)$

$Pb^{2+}(aq) + CO_3^{2+}(aq) \rightarrow PbCO_3(s)$

(ii)When the insoluble precipitates are acidified with nitric(V) acid,

- Lead(II)chloride and Lead(II)sulphate(VI) do not react with the acid and thus their white precipitates remain/ persists.
- Lead(II) sulphate (IV) and Lead(II)carbonate(IV) reacts with the acid to form **soluble** Lead(II) nitrate (V) and produce/effervesces/fizzes/bubbles out **sulphur(IV)oxide** and **carbon(IV)oxide** gases respectively.

. Chemical/ionic equation:

$PbSO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + Pb^{2+}(aq) + SO_2(g)$

$PbCO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + Pb^{2+}(aq) + CO_2(g)$

(iii)When Lead(II)chloride and Lead(II)sulphate(VI) are heated/warmed;

- Lead(II)chloride dissolves in hot water/on boiling(recrystallizes on cooling)
- Lead(II)sulphate(VI) do not dissolve in hot water thus its white precipitate persists/remains on heating/boiling.

(iv)When sulphur(IV)oxide and carbon(IV)oxide gases are produced;

- **sulphur(IV)oxide** will decolorize acidified potassium manganate(VII) and / or Orange colour of acidified potassium dichromate(VI) will turns to green.

Carbon(IV)oxide will not.

Chemical equation:

$5SO_3^{2-}(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(l)$

(purple) (colourless)

$3SO_3^{2-}(aq) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \rightarrow 3SO_4^{2-}(aq) + 2Cr^{3+}(aq) + 4H_2O(l)$

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(Orange) (green)

- **Carbon(IV)oxide** forms an insoluble white precipitate of calcium carbonate if three drops of lime water are added into the reaction test tube when effervescence is taking place. **Sulphur(IV)oxide will not.**

Chemical equation:

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

These tests should be done immediately after acidifying to ensure the gases produced react with the oxidizing agents/lime water.

Using Barium(II)nitrate(V)/ Barium(II)Chloride

(i)Barium(II)nitrate(V) and/ or Barium(II)chlo**ride** solution reacts with Sulphate (VI) salts (SO_4^{2-}), Sulphate (IV)salts (SO_3^{2-}) and carbonates(CO_3^{2-}) to form the insoluble white precipitate of Barium(II)sulphate(VI), Barium(II) sulphate (IV) and Barium(II)carbonate(IV).

Chemical/ionic equation:

$Ba^{2+}(aq) + SO_4^{2+}(aq) \rightarrow BaSO_4(s)$

$Ba^{2+}(aq) + SO_3^{2+}(aq) \rightarrow BaSO_3(s)$

$Ba^{2+}(aq) + CO_3^{2+}(aq) \rightarrow BaCO_3(s)$

(ii)When the insoluble precipitates are acidified with nitric(V) acid,

- Barium (II)sulphate(VI) do not react with the acid and thus its white precipitates remain/ persists.
- Barium(II) sulphate (IV) and Barium(II)carbonate(IV) reacts with the acid to form **soluble** Barium(II) nitrate (V) and produce /effervesces /fizzes/ bubbles out **sulphur(IV)oxide** and **carbon(IV)oxide** gases respectively.

. Chemical/ionic equation:

$BaSO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + Ba^{2+}(aq) + SO_2(g)$

$BaCO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + Ba^{2+}(aq) + CO_2(g)$

(iii) When sulphur(IV)oxide and carbon(IV)oxide gases are produced;

- **sulphur(IV)oxide** will decolorize acidified potassium manganate(VII) and / or Orange colour of acidified potassium dichromate(VI) will turns to green.

Carbon(IV)oxide will not.

Chemical equation:

$5SO_3^{2-}(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(l)$

(purple) (colourless)

$3SO_3^{2-}(aq) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \rightarrow 3SO_4^{2-}(aq) + 2Cr^{3+}(aq) + 4H_2O(l)$

(Orange) (green)

- **Carbon(IV)oxide** forms an insoluble white precipitate of calcium carbonate if three drops of lime water are added into the reaction test tube when effervescence is taking place. **Sulphur(IV)oxide will not.**

Chemical equation:

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

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These tests should be done immediately after acidifying to ensure the gases produced react with the oxidizing agents/lime water.

(iii) Sodium carbonate(IV) (Na_2CO_3)

(a)Extraction of sodium carbonate from soda ash

Sodium carbonate naturally occurs in Lake Magadi in Kenya as Trona.trona is the double salt ; sodium sesquicarbonate. $NaHCO_3 .Na_2CO_3 .H_2O$.It is formed from the volcanic activity that takes place in Lake Naivasha, Nakuru ,Bogoria and Elementeita .All these lakes drain into Lake Magadi through underground rivers. Lake Magadi has no outlet.

Solubility of Trona decrease with increase in temperature.High temperature during the day causes trona to naturally crystallize .It is mechanically scooped/dredged/dug and put in a furnace.

Inside the furnace, trona decompose into soda ash/sodium carbonate.

Chemical equation

$2NaHCO_3 .Na_2CO_3 .H_2O\ (s) \rightarrow 3Na_2CO_3\ (s) + 5H_2O(l) + CO_2\ (g)$

(trona) (soda ash)

Soda ash is then bagged and sold as Magadi soda.It is mainly used:

(i)in making glass to lower the melting point of raw materials (sand/SiO_2 from 1650°C and CaO from 2500°C to around 1500°C)

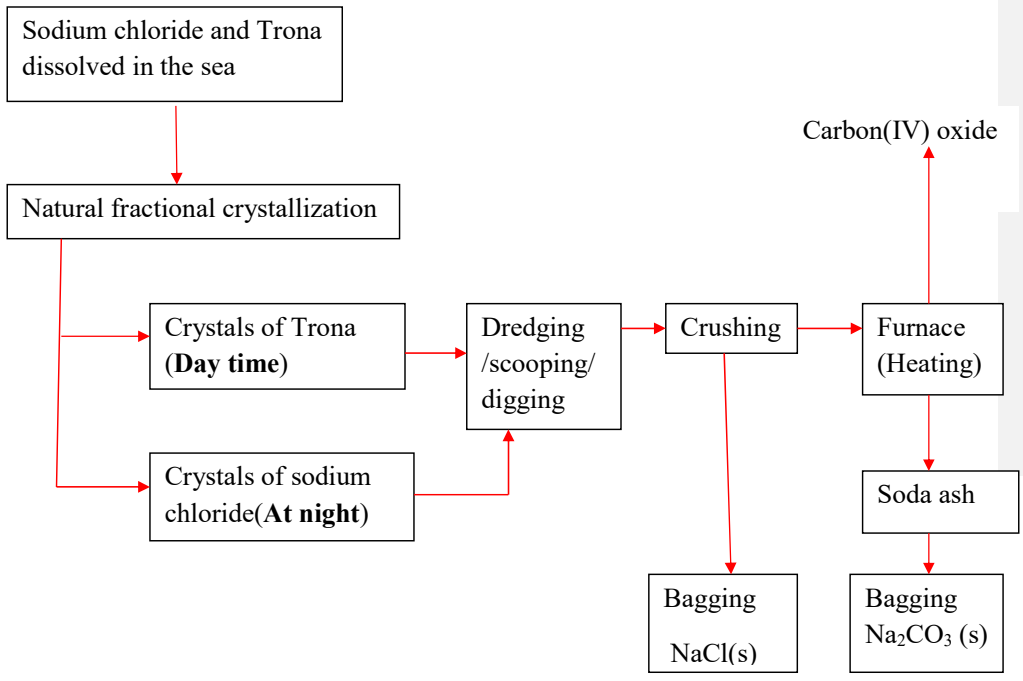
(ii)in softening hard water

(iii)in the manufacture of soapless detergents.

(iv)Swimming pool "pH increaser"

Sodium chloride is also found dissolved in the lake. Solubility of sodium chloride decrease with decreases in temperature/ sodium chloride has lower solubility at lower temperatures. When temperatures decrease at night it crystallize out .The crystals are then mechanically dug/dredged /scooped then packed for sale as animal/cattle feeds and seasoning food.

Summary flow diagram showing the extraction of Soda ash from Trona



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b)The Solvay process for industrial manufacture of sodium carbonate(IV)
(i)Raw materials.

-**Brine** /Concentrated Sodium chloride from salty seas/lakes.
-**Ammonia** gas from Haber.
-**Limestone** /Calcium carbonate from chalk /limestone rich rocks.
-**Water** from rivers/lakes.

(ii)Chemical processes

Ammonia gas is passed **up** to meet a **downward** flow of sodium chloride solution / brine to form **ammoniated** brine/**ammoniacal** brine **mixture** in the **ammoniated brine chamber**

The ammoniated brine mixture is then pumped up, atop the carbonator/ solvay tower.

In the carbonator/ solvay tower, ammoniated brine/ammoniacal brine mixture slowly trickle down to meet an upward flow of carbon(IV)oxide gas.

The carbonator is shelved /packed with quartz/broken glass to

(i) reduce the rate of flow of ammoniated brine/ammoniacal brine mixture.
(ii)increase surface area of the liquid mixture to ensure a lot of ammoniated brine/ammoniacal brine mixture react with carbon(IV)oxide gas.

Insoluble sodium hydrogen carbonate and soluble ammonium chloride are formed from the reaction.

Chemical equation

$CO_2(g) + H_2O(l) + NaCl\ (aq) + NH_3(g) \rightarrow NaHCO_3(s) + NH_4Cl(aq)$

The products are then filtered. **Insoluble** sodium hydrogen carbonate forms the **residue** while soluble ammonium chloride forms the **filtrate**.

Sodium hydrogen carbonate itself can be used:

(i) as baking powder and preservation of some soft drinks.
(ii) as a buffer agent and antacid in animal feeds to improve fibre digestion.
(iii) making dry chemical fire extinguishers.

In the Solvay process Sodium hydrogen carbonate is then heated to form Sodium carbonate/soda ash, water and carbon (IV) oxide gas.

Chemical equation

$2NaHCO_3\ (s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(l)$

Sodium carbonate is stored ready for use in:

(i) during making glass/lowering the melting point of mixture of sand/SiO_2 from 1650°C and CaO from 2500°C to around 1500°C

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(ii) in softening hard water
(iii) in the manufacture of soapless detergents.
(iv) swimming pool "pH increaser".

Water and carbon(IV)oxide gas are recycled back to the ammoniated brine/ammoniacal brine chamber.

More carbon(IV)oxide is produced in the kiln/furnace. Limestone is heated to decompose into Calcium oxide and carbon(IV)oxide.

Chemical equation

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

Carbon(IV)oxide is recycled to the carbonator/solvay tower. Carbon (IV)oxide is added water in the **slaker** to form Calcium hydroxide. This process is called **slaking**.

Chemical equation

$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$

Calcium hydroxide is mixed with ammonium chloride from the carbonator/solvay tower in the **ammonia regeneration chamber** to form Calcium chloride, water and more ammonia gas.

Chemical equation

$Ca(OH)_2(aq) + 2NH_4Cl(aq) \rightarrow CaCl_2(s) + 2NH_3(g) + H_2O(l)$

$NH_3(g)$ and $H_2O(l)$ are recycled.

Calcium chloride may be used:

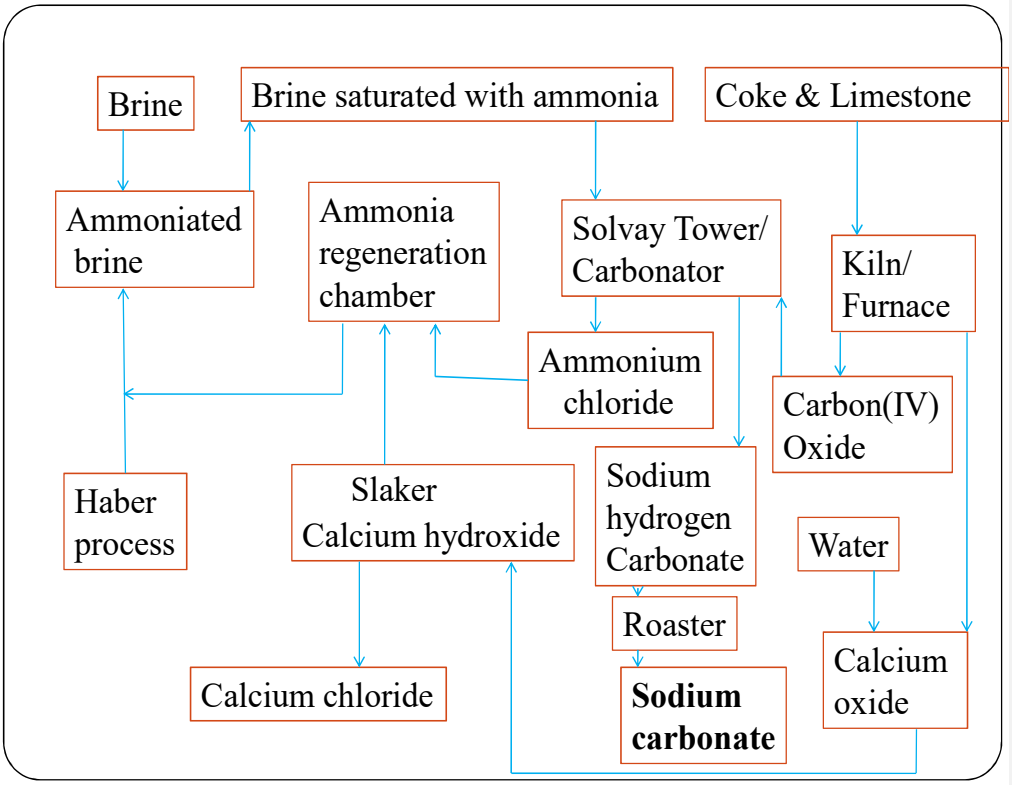
(i)as drying agent in the school laboratory during gas preparation (except ammonia gas)

(ii)to lower the melting point of solid sodium chloride / rock salt salts during the Downs process for industrial extraction of sodium metal.

Detailed Summary flow diagram of Solvay Process

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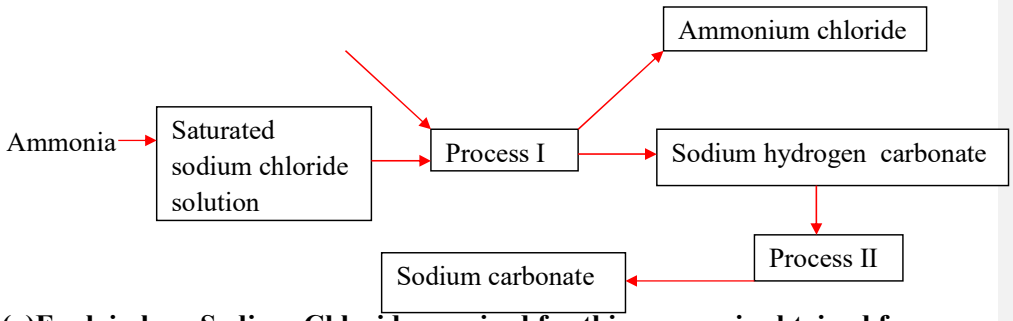
Practice

1. The diagram below shows part of the Solvay process used in manufacturing sodium carbonate. Use it to answer the questions that follow.

Carbon (IV)oxide

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(a)Explain how Sodium Chloride required for this process is obtained from the sea.

Sea water is pumped /scooped into shallow pods. Evaporation of most of the water takes place leaving a very concentrated solution.

(b)(i) Name process:

I. Filtration

II. Decomposition

(ii) Write the equation for the reaction in process:

Process I

Chemical equation

$CO_2(g) + H_2O(l) + NaCl\ (aq) + NH_3(g) \rightarrow NaHCO_3(s) + NH_4Cl(aq)$

Process II

Chemical equation

$2NaHCO_3\ (s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(l)$

(c)(i) Name two substances recycled in the solvay process

Ammonia gas , Carbon(IV)Oxide and Water.

(ii)Which is the by-product of this process?

Calcium(II)Chloride /$CaCl_2$

(iii)State two uses that the by-product can be used for:

1. As a drying agent in the school laboratory preparation of gases.
2. In the Downs cell/process for extraction of Sodium to lower the melting point of rock salt.

(iv)Write the chemical equation for the formation of the by-products in the Solvay process.

Chemical equation

$Ca(OH)_2\ (aq) + 2NH_4Cl\ (aq) \rightarrow CaCl_2(s) + 2NH_3(g) + H_2O(l)$

(d)In an experiment to determine the % purity of Sodium carbonate produced in the Solvay process, 2.15g of the sample reacted with exactly 40.0cm3 of 0.5M Sulphuric (VI)acid.

(i)Calculate the number of moles of sodium carbonate that reacted.

Chemical equation

$Na_2CO_3\ (aq) + H_2SO_4\ (aq) \rightarrow Na_2SO_4\ (aq) + CO_2(g) + H_2O(l)$

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Mole ratio Na_2CO_3 :H_2SO_4 => 1:1

Moles H_2SO_4 = $\frac{\text{Molarity x Volume}}{1000}$ => $\frac{0.5 \times 40.0}{1000}$ = **0.02 Moles**

Moles of Na_2CO_3 = **0.02 Moles**

(ii)Determine the % of sodium carbonate in the sample.

Molar mass of Na_2CO_3 = **106g**

Mass of Na_2CO_3 = moles x Molar mass => 0.02 x 106 = **2.12 g**

% of Na_2CO_3 = $\frac{(2.12 \text{ g} \times 100)}{2.15}$ = **98.6047%**

(e) State two uses of soda ash.

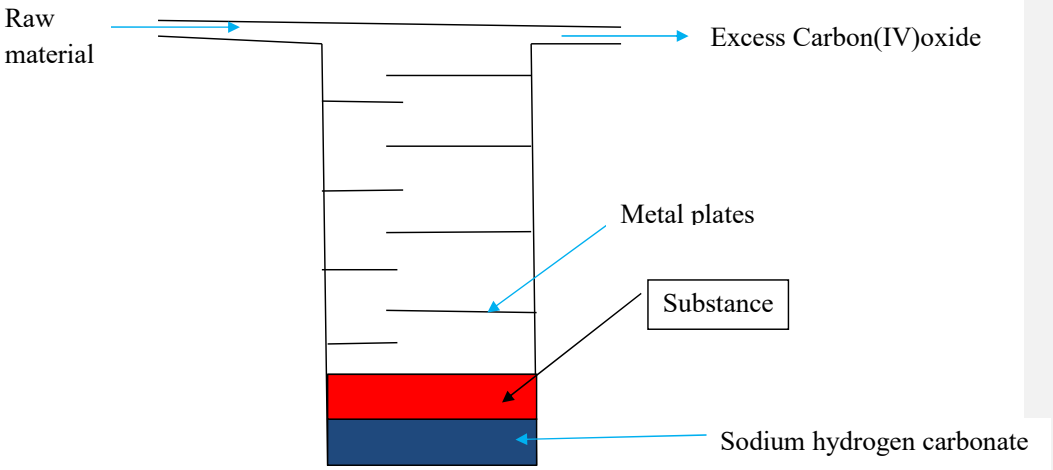
(i) during making glass/lowering the melting point of mixture of sand/SiO_2 from 1650ºC and CaO from 2500ºC to around 1500ºC

(ii) in softening hard water

(iii) in the manufacture of soapless detergents.

(iv) swimming pool "pH increaser".

(f)The diagram below shows a simple ammonia soda tower used in manufacturing sodium carbonate .Use it to answer the questions that follow:



(i)Name the raw materials needed in the above process

-Ammonia
-Water
-Carbon(IV)oxide
-Limestone
-Brine/ Concentrated sodium chloride

(ii)Identify substance A

Ammonium chloride /NH_4Cl

(iii) Write the equation for the reaction taking place in:

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I.Tower.

Chemical equation

$CO_2(g) + NaCl\ (aq) + H_2O(l) + NH_3(g) \rightarrow NaHCO_3(s) + NH_4Cl(aq)$

II. Production of excess carbon (IV)oxide.

Chemical equation

$CaCO_3\ (s) \rightarrow CaO(s) + CO_2(g)$

III. The regeneration of ammonia

Chemical equation

$Ca(OH)_2\ (aq) + 2NH_4Cl\ (aq) \rightarrow CaCl_2(s) + 2NH_3(g) + H_2O(l)$

(iv)Give a reason for having the circular metal plates in the tower.

-To slow the downward flow of brine.

-To increase the rate of dissolving of ammonia.

-To increase the surface area for dissolution

(v)Name the gases recycled in the process illustrated above.

Ammonia gas , Carbon(IV)Oxide and Water.

2. Describe how you would differentiate between carbon (IV)oxide and carbon(II)oxide using chemical method.

Method I

-Bubble both gases in lime water/$Ca(OH)_2$

-white precipitate is formed if the gas is carbon (IV) oxide

- **No** white precipitate is formed if the gas is carbon (II) oxide

Method II

-ignite both gases

- Carbon (IV) oxide does **not** burn/ignite

- Carbon (II) oxide burn with a blue non-sooty flame.

Method III

-Lower a burning splint into a gas containing each gas separately.

-burning splint is extinguished if the gas is carbon (IV) oxide

-burning splint is **not** extinguished if the gas is carbon (II) oxide.

3.Using Magnesium sulphate(VI)solution ,describe how you can differentiate between a solution of sodium carbonate from a solution of sodium hydrogen carbonate

-Add Magnesium sulphate(VI) solution to separate portions of a solution of sodium carbonate and sodium hydrogen carbonate in separate test tubes

-White precipitate is formed in test tube containing sodium carbonate

-No white precipitate is formed in test tube containing sodium hydrogen carbonate.

Chemical equation

$Na_2CO_3\ (aq) + MgSO_4\ (aq) \rightarrow Na_2SO_4\ (aq) + MgCO_3(s)$ (white ppt)

Ionic equation

$CO_3^{2-}\ (aq) + Mg^{2+}\ (aq) \rightarrow MgCO_3(s)$

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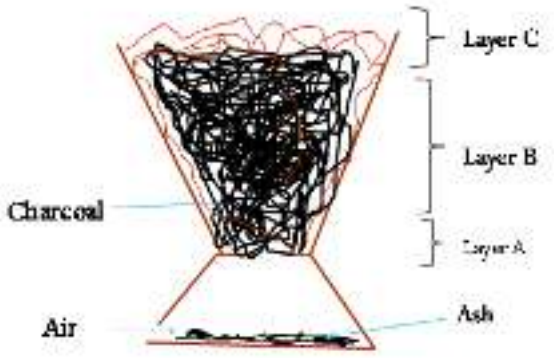
(white ppt)

Chemical equation

$2NaHCO_3\ (aq) + MgSO_4\ (aq) \rightarrow Na_2SO_4\ (aq) + Mg(HCO_3)_2\ (aq)$

(colourless solution)

4. The diagram below shows a common charcoal burner .Assume the burning take place in a room with sufficient supply of air.



(a)Explain what happens around:

(i)Layer A

Sufficient/excess air /oxygen enter through the air holes into the burner .It reacts with/oxidizes Carbon to carbon(IV)oxide

Chemical equation

$C(s) + O_2(g) \rightarrow CO_2\ (g)$

(ii)Layer B

Hot carbon(IV)oxide rises up and is reduced by more carbon/charcoal to carbon (II)oxide.

Chemical equation

$C(s) + CO_2(g) \rightarrow 2CO\ (g)$

(ii)Layer C

Hot carbon(II)oxide rises up and burns with a blue flame to be oxidized by the excess air to form carbon(IV)oxide.

$2CO\ (g) + O_2(g) \rightarrow 2CO_2(g)$

(b)State and explain what would happen if the burner is put in an enclosed room.

The hot poisonous /toxic carbon(II)oxide rising up will not be oxidized to Carbon(IV)oxide.

(c)Using a chemical test, describe how you would differentiate two unlabelled black solids suspected to be charcoal and copper(II)oxide.

Method I

-Burn/Ignite the two substances separately.

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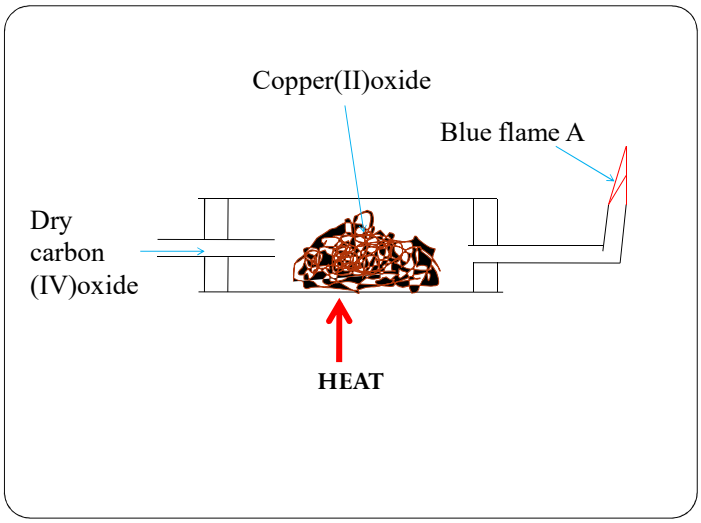
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-Charcoal burns with a blue flame
- Copper(II)oxide does not burn

Method II

-Add dilute sulphuric(VI)acid/Nitric(V)acid/Hydrochloric acid separately.
-Charcoal does not dissolve.
- Copper(II)oxide dissolves to form a colourless solution.

5. Excess Carbon(II)oxide was passed over heated copper(II)oxide as in the set up shown below for five minutes.



(a)State and explain the observations made in the combustion tube.

Observation

Colour change from black to brown

Explanation

Carbon (II)oxide reduces black copper(II)oxide to brown copper metal itself oxidized to Carbon(IV)oxide.

Chemical equation

$CO(g) + CuO(s) \rightarrow Cu(s) + CO_2(g)$

(black) (brown)

(b) (i)Name the gas producing flame A

Carbon(II)oxide

(ii)Why should the gas be burnt?

It is toxic/poisonous

(iii)Write the chemical equation for the production of flame A

$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$

(c)State and explain what happens when carbon(IV)oxide is prepared using Barium carbonate and dilute sulphuric(VI)acid.

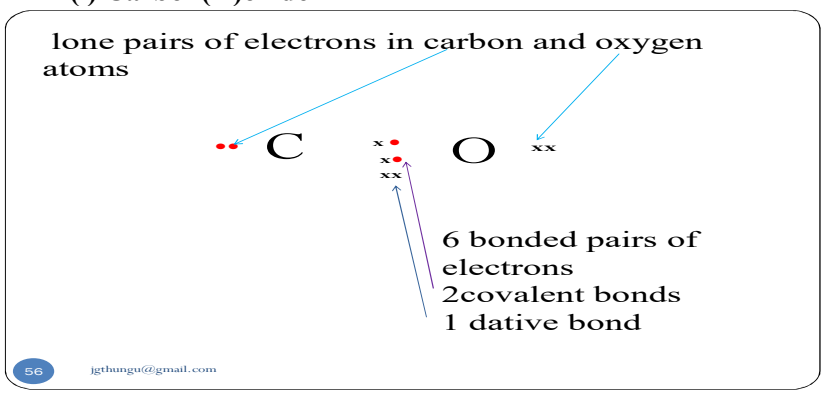
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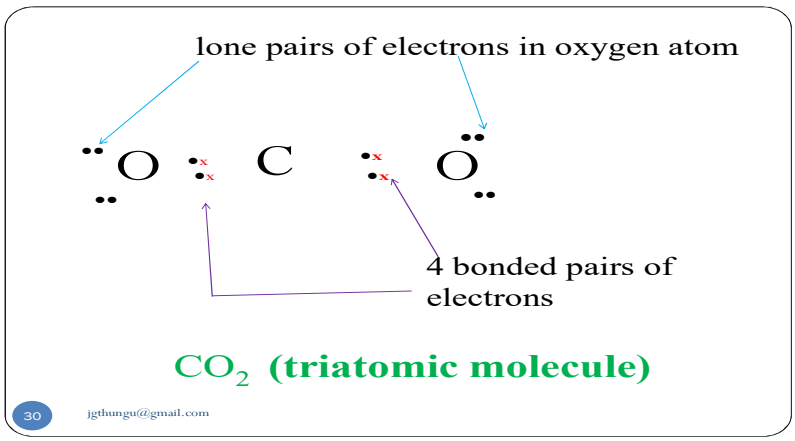
Reaction starts then stops after sometime producing small/little quantity of carbon(IV)oxide gas.
Barium carbonate react with dilute sulphuric(VI)acid to form insoluble Barium sulphate(VI) that cover/coat unreacted Barium carbonate stopping further reaction to produce more Carbon(IV)oxide.

(d) Using dot () and cross(x) to represent electrons show the bonding in a molecule of :

(i) Carbon(II)oxide



(ii) Carbon(IV)Oxide.



(e) Carbon (IV)oxide is an environmental pollutant of global concern. Explain.

-It is a green house gas thus causes global warming.

-It dissolves in water to form acidic carbonic acid which causes "acid rain"

(f)Explain using chemical equation why lime water is used to test for the presence of Carbon (IV) oxide instead of sodium hydroxide.

Using lime water/calcium hydroxide:

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- a visible white precipitate of calcium carbonate is formed that dissolves on bubbling excess Carbon (IV) oxide gas

Chemical equation

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

(white precipitate)

$CaCO_3(aq) + H_2O(l) + CO_2(g) \rightarrow Ca(HCO_3)_2(aq)$

Using sodium hydroxide:

- **No** precipitate of sodium carbonate is formed Both sodium carbonate and sodium hydrogen carbonate are soluble salts/dissolves.

Chemical equation

$2NaOH(aq) + CO_2(g) \rightarrow Na_2CO_3(s) + H_2O(l)$

(No white precipitate)

$Na_2CO_3(s) + H_2O(l) + CO_2(g) \rightarrow 2NaHCO_3(s)$

(g)Ethan-1,2-dioic acid and methanoic acid may be used to prepare small amount of carbon(II)oxide in a school laboratory.

(i) Explain the modification in the set up when using one over the other.

Before carbon(II)oxide is collected:

-when using methanoic acid, **no** concentrated sodium/potassium hydroxide is needed to absorb Carbon(IV)oxide.

-when using ethan-1,2-dioic acid, concentrated sodium/potassium hydroxide is needed to absorb Carbon(IV)oxide.

(ii)Write the equation for the reaction for the formation of carbon(II)oxide from:

I.Methanoic acid.

Chemical equation $HCOOH(aq) \rightarrow CO(g) + H_2O(l)$

II. Ethan-1,2-dioic acid

Chemical equation $HOOCCOOH(aq) \rightarrow CO_2(g)+CO(g)+H_2O(l)$

(h)Both carbon(II)oxide and carbon(IV)oxide affect the environment. Explain why carbon(II)oxide is more toxic/poisonous.

-Both gases are colourless,denser than water and odourless.

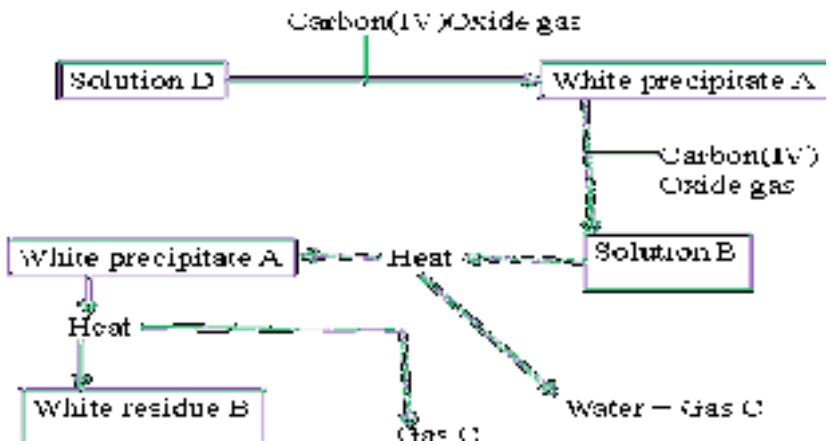
-Carbon(II)oxide is preferentially absorbed by human/mammalian haemoglobin when inhaled forming stable carboxyhaemoglobin instead of oxyhaemoglobin.This reduces the free haemoglobin in the blood leading to suffocation and quick death. **--** Carbon(IV)oxide is a green house gas that increases global warming.

-Carbon(II)oxide is readily oxidized to carbon(IV)oxide

6.Study the flow chart below and use it to answer the questions that follow.

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(a)Name:

(i)the white precipitate A

Calcium carbonate

(ii) solution B

Calcium hydrogen carbonate

(iii) gas C

Carbon(IV)oxide

(iv) white residue B

Calcium oxide

(v) solution D

Calcium hydroxide/lime water

(b)Write a balanced chemical equation for the reaction for the formation of:

(i) the white precipitate A from solution D

Chemical equation

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

(ii) the white precipitate A from solution B

Chemical equation

$Ca(HCO_3)_2(aq) \rightarrow CO_2(g) + CaCO_3(s) + H_2O(l)$

(iii) solution B from the white precipitate A

Chemical equation

$CO_2(g) + CaCO_3(s) + H_2O(l) \rightarrow Ca(HCO_3)_2(aq)$

(iv) white residue B from the white precipitate A

Chemical equation

$CaCO_3(s) \rightarrow CO_2(g) + CaO(s)$

(iv) reaction of white residue B with water

Chemical equation

$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$

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CHEMISTRY OF CHLORINE

A.CHLORINE

Chlorine is a non-metallic element in group VII (Group 17) of the periodic table. It has electronic configuration 2:8:7. It gains one valence election to form stable Cl^- ion, it belongs to the chemical family of halogens.

Occurrence

-As Brine-concentration sodium chloride solution dissolved in salty seas water, oceans and lakes e.g. Lake Magadi in Kenya is very salty.

-As rock-salt solid sodium chloride crystals in the earths crust all over the world.

B) Preparation

Chlorine gas may be prepared in the school laboratory from the following:

a)Heating solid Manganese (iv) Oxide and Concentrated Hydrochloric acid.

b) Heating Lead (IV) Oxide and concentrated hydrochloric acid.

c)Reacting Potassium Manganate (VII) with concentrated Hydrochloric acid

d)Reacting Potassium /sodium Dichromate (VI) Acid with Concentrated Hydrochloric acid.

Set up of school laboratory preparation of chlorine.



Preparation of Chlorine from Conc. HCl and Oxidizing agent

c) Properties of chlorine. (Questions)

1. What is the colour of chlorine?

Pale green.

2. Describe the smell of chlorine.

Pungent irritating smell.

3. What method is used in collection of chlorine gas explain.

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-Downward delivery.

-Chlorine is $1^1/_2$ denser than air.

4.(i) What is the purpose of concentrated sulphuric (VI) acid.

-To dry the gas.

(ii) Name two other substances that can be used in place of concentrated sulphuric (VI) acid.

-Calcium chloride

-Silica gel

(iii) Name a substance that cannot be used in place of concentrated sulphuric (VI) acid explain.

-Calcium oxide reacts with chlorine.

5.(a)Write three possible reactions between concentrated hydrochloric acid and the oxidizing agents.

1. **$2KMnO_4(s) + 16HCl(aq) \rightarrow 2KCl(aq) + 2MnCl_2(aq) + 8H_2O(l) + 5Cl_2(g)$**
2. **$K_2Cr_2O_7(s) + 14HCl(aq) \rightarrow 2KCl(aq) + 2CrCl_3(aq) + 7H_2O(l) + 3Cl_2(g)$**
3. **$Na_2Cr_2O_7(s) + 14HCl(aq) \rightarrow 2NaCl(aq) + CrCl_3(aq) + 7H_2O(l) + 3Cl_2(g)$**
4. **$PbO_2(s) + 4HCl(aq) \rightarrow PbCl_2(aq) + Cl_2(g) + 2H_2O(l)$**
5. **$MnO_2(s) + 4HCl(aq) \rightarrow MnCl_2(aq) + Cl_2(g) + 2H_2O(l)$**

(b) Why is Hydrochloric acid used in all the above cases?

Oxidizing agents $KMnO_4/PbO_2/MnO_2/K_2Cr_2O/Na_2Cr_2O_7$ readily oxidize hydrochloric acid to chlorine themselves reduced to their chlorides.

Generally:

$2HCl\ (aq) + [O] \rightarrow Cl_2\ (g) + H_2O\ (l)$

(From oxidizing agent)

6. State and explain the observation made when chlorine is bubbled in water.

Observation

-Pale yellow colour of chlorine fades.

-yellow solution formed.

Explanation

Chlorine dissolves then reacts with water to form yellow chlorine water.

Chlorine water is chemically a mixture of hydrochloric acid and chloric(I)acid (hypochlorous acid).

A mixture of hydrochloric acid and chloric(I)acid (hypochlorous acid) is commonly called Chlorine water

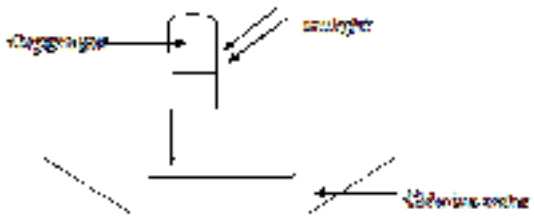
Chemical equation:

$Cl_2(g) + H_2O(l) \rightarrow HCl(aq) + HClO(aq)$

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7. Chlorine water in a boiling tube inverted into a trough was exposed to sunlight for two hours. Using a well labeled diagram show and explain the observations made.



Chlorine (I) acid is an unstable compound.

After two hours the chloric (I) acid in chlorine water decomposes to hydrochloric acid and releases oxygen gas. This reaction takes place in sunlight.

Chemical equation

$$\mathbf{2HOCl(aq) \rightarrow 2HCl(aq) + O_2(g)}$$

8. State and explain the observation made when chlorine gas is bubbled in gas jar containing damp/wet/moist litmus papers.

Observation

The blue litmus turns red then both the red/blue litmus papers are bleached/decolourized.

Explanation

Chlorine reacts with water in the litmus papers to form acidic hydrochloric acid and chloric (l) acid that turns blue litmus papers red.

Chemical Equation

$$\mathbf{Cl_2(g) + H_2O(l) \rightarrow HCl(aq) + HClO(aq)}$$

Explanation

Unstable chloric (I) acid oxidizes the dye/colured litmus paper to colourless material

Chemical Equation

$$\mathbf{HClO(aq) + \underset{(coloured)}{dye} \rightarrow HCl(aq) + \underset{(colourless)}{(dye + O)}}$$

Or:

$$\mathbf{HClO(aq) + \underset{(coloured)}{dye\text{-}O} \rightarrow HCl(aq) + \underset{(colourless)}{dye}}$$

NB Chlorine does not therefore bleach/decolourize dry litmus paper/dye because chloric(I) acid cannot be formed in absence of water.

9. Blue litmus papers were put in a flask containing cold dilute sodium hydroxide. Chlorine gas was bubbled into the solution. State and explain the observations made.

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Observation

blue litmus papers were bleached /decolorized.

Pale green colour of chlorine fades.

Explanation

-Sodium hydroxide reacts with chlorine to form sodium chloride and sodium hypochlorite. Sodium hypochlorite bleaches dyes by oxidation.

Chemical Equation

$$Cl_2 + 2NaOH(aq) \rightarrow NaCl(aq) + NaClO(aq) + H_2O$$

$$NaClO(aq) + \underset{\text{(coloured)}}{\text{dye}} \rightarrow NaCl(aq) + \underset{\text{(Colourless)}}{(\text{dye} + O)}$$

$$NaClO(aq) + \underset{\text{(Coloured)}}{(\text{dye-O})} \rightarrow NaCl(aq) + \underset{\text{(Colourless)}}{\text{dye}}$$

10.Blue litmus papers were put in flask containing hot concentrated sodium hydroxide. Chlorine gas was bubbled into the solution. State and explain the observations made.

Observation.

blue litmus papers were bleached.

Pale green colour of chlorine fades.

Explanation

Hot concentrated sodium hydroxide reacts with chlorine to form sodium chloride and sodium chloride (V).Sodium chlorate (V) bleaches by oxidation.

Chemical equation

$$2Cl_2(g) + 4NaOH(aq) \rightarrow 3NaCl(aq) + NaClO_3(aq) + H_2O(l)$$

$$NaClO_3(aq) + 3(\text{dyes}) \rightarrow NaCl(aq) + 3(\text{dye} + O)$$

$$NaClO_3(aq) + 3(\text{dyes-O}) \rightarrow NaCl(aq) + 3\ \text{dyes}$$

$NaClO_3$ is also a weed killer

11. State three main use of chlorine gas.

-Manufacture of polyvinyl chloride (P.V.C) // polychloroethene pipes.

-Manufacture of hydrochloric acid used in “Pickling” of metals.

-Manufacture of bleaching agents

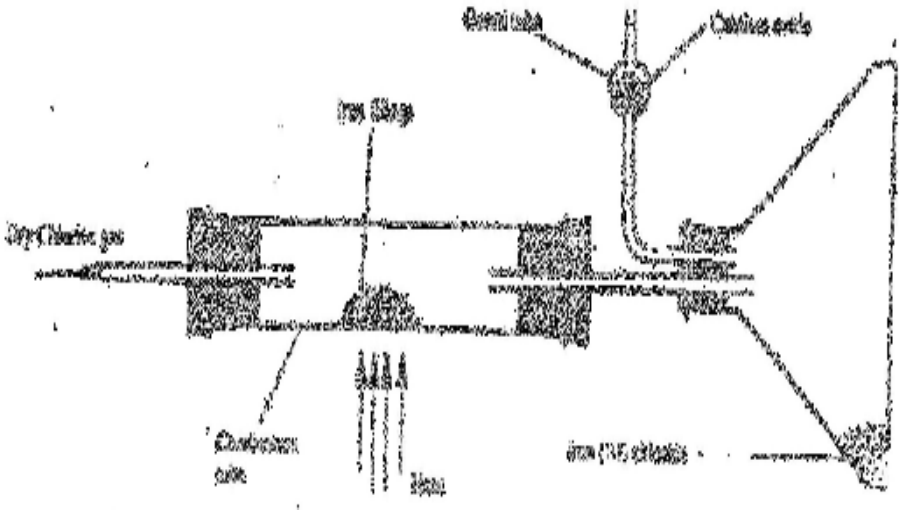
-Chlorination of water to kill germs.

12. The diagram below shows the effect of chlorine on heated iron wool.

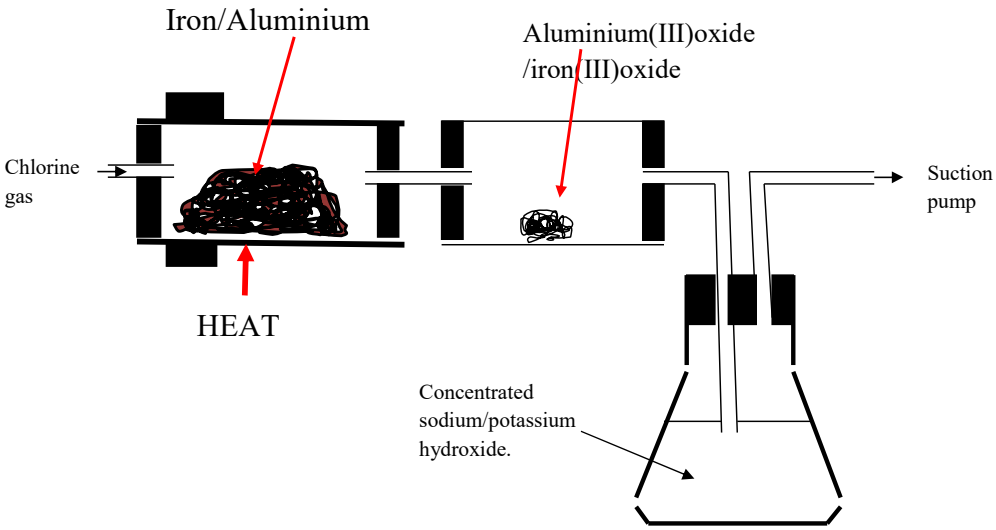
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Method I



Method II



a) Identify a suitable drying agent to dry chlorine gas.

-Conc. H_2SO_4 / Concentrated sulphuric (VI) acid.

-Anhydrous Calcium (II) Chloride.

-Silica gel

b) State and explain the observations made in combustion tube in method I and II

Observation

Iron glows red hot

Brown crystals are formed

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Explanation

Iron reacts with chlorine to form dark brown crystals of iron (III) Chloride.

This reaction is exothermic and requires no farther heating once started. Iron (III) Chloride sublimes away ensuring the unreacted Iron completely reacts with chlorine gas.

Chemical equation

$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(g)$

c) (i) Why is the brown solid collected at the point as shown in method I and II.

-Heated iron (III) Chloride crystals sublime to gas and solidify on the cooler parts.

(ii) Name another metal that can be used in place of iron to react with chlorine and collected at similar point on heating explain.

Metal Aluminum

Explanation

Aluminum reacts with chlorine to form a while sublimate of aluminum (III) chloride at the cooler parts

Chemical equation

$2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s/g)$

d) What is the purpose of suction pump?

To pull the gaseous products into the set up.

e) What is the function of:

(i) Sodium hydroxide in method II. Explain.

To absorb poisonous/toxic excess unreacted chlorine gas.

Sodium hydroxide reacts with chlorine to form sodium chloride, Sodium hypochlorite and water.

Chemical equation:

$2NaOH(aq) + Cl_2(g) \rightarrow NaCl(aq) + NaClO(aq) + H_2O(l)$

$2KOH(aq) + Cl_2(g) \rightarrow KCl(aq) + KClO(aq) + H_2O(l)$

(ii) Anhydrous calcium chloride/calcium oxide in method I. Explain.

To absorb moisture/water in the set up to prevent it from hydrolyzing iron (III) chloride/aluminium oxide.

Explanation

Iron (III) chloride and Aluminium chloride fumes and reacts with small traces of water to form a solution of iron (III) hydroxide/aluminium hydroxide and hydrogen chloride gas.

Chemical equation

$FeCl_3(s) + 3HCl(aq) \rightarrow Fe(OH)_3(aq) + 3HCl(g)$

$AlCl_3(s) + 3HCl(aq) \rightarrow Al(OH)_3(aq) + 3HCl(g)$

f) Based on e (i) and (ii) above what precaution should be made in:

(i) method II to ensure correct results.

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-Tube B should be completely dry to prevent hydrolysis of iron (III) Chloride to iron (III) hydroxide.

(ii) Carrying out method I

-Should be done in a fume chamber or in the open because chlorine gas is poisonous/toxic.

(g) Name another substance that can be used place of Sodium hydroxide in method I

Potassium hydroxide

(h) Calcium oxide cannot be used in place of calcium chloride during preparation of chlorine. Explain.

Calcium oxide is a base. It reacts /absorbs water to form calcium hydroxide solution.

Calcium hydroxide reacts with chlorine to form a mixture of calcium chloride and calcium hypochlorite.

Chemical equation

$2Ca(OH)_2(aq) + 2Cl_2(g) \rightarrow CaCl_2(aq) + CaOCl_2(aq) + H_2O(l)$

13. (a)State and explain the observation made when a piece of burning magnesium ribbon is lowered in a gas jar containing chlorine gas.

-Magnesium ribbon continues burning with a bright flame.

-White solid formed.

-Pale yellow colour of chlorine fades

Explanation**:**

Magnesium reacts with chlorine forming a white solid of magnesium chloride.

Chemical equation

$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$

(b) Write the equation for the reaction that takes place if zinc is used.

$Zn(s) + Cl_2(g) \rightarrow ZnCl_2(s)$

14. Burning phosphorus was lowered in a gas jar containing chlorine gas.

a) State the observations made.

-Phosphorus continues to burn.

-Dense white fumes formed.

-Pale green colour of chlorine fades.

b) Write two possible equations that take place.

$P_4(s) + 6Cl_2(g) \rightarrow 4PCl_3(s)$

$P_4(s) + 10Cl_2(g) \rightarrow 4PCl_3(s)$

(c) State two reasons why the deflagrating spoon with rid/cover should be used.

-Chlorine in the gas jar is poisonous/toxic.

-Burning phosphorus produces poisonous/toxic phosphorus (III) chloride // phosphorus (V) chloride.

-Ensure the reaction is not affected by air/oxygen from the atmosphere.

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(d) After the reaction is complete, 2cm^3 of distilled water were added. The solution formed was tested with both blue and red litmus papers.

(i) State the observations made.

-Blue litmus paper turns red

-Red litmus paper remain red

(ii) Explain the observation made in d(i) above

-Phosphoric (V) Chloride hydrolyze in water to phosphoric (V) acid and produce hydrogen chloride gas. Both hydrogen chloride and phosphoric (V) acid are acidic.

Chemical equation

$PCl_5(l) + 4H_2O(l) \rightarrow H_3PO_4(aq) + 5HCl(g)$

15. State and explain the observations made when gas jar containing chlorine is inverted over another containing hydrogen sulphide gas.

Observation

Yellow solid formed.

Pale colour of chlorine fades

Explanation

Chlorine oxidizes hydrogen sulphide to sulphur itself reduced to hydrogen chloride gas. A little water catalyzes the reaction.

Chemical equation

$H_2S(g) + Cl_2(g) \rightarrow S(s) + HCl(g)$

(yellow solid) (White Fume)

16. Chlorine was bubbled in aqueous ammonia solution in a beaker state and explain the observation made.

Observation:

White fumes evolved.

Pale green colour of chlorine fades.

Explanation

Chlorine reacts with ammonia gas to form a dense white fume of ammonia chloride and Nitrogen gas is produced.

Chemical equation

$8NH_3(g) + 3Cl_2(g) \rightarrow 6NH_4Cl(s) + N_2(g)$

17. (a) Dry gas was bubbled in cold dilute sodium hydroxide solution. Explain the observations made:

Observation

Pale green colour of chlorine fades.

Pale yellow solution is formed.

Explanation

Chlorine reacts withhot concentrated sodiumsodium hydroxide / Potassium hydroxide solution to form pale yellow solution of metal chlorate (V) and chlorides of the metal

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Chemical equation

$Cl_2(g)$ + 2NaOH → NaClO(aq) + NaCl(aq) + $H_2O(l)$

(sodium hydroxide) (Sodium Chlorate (I))

$Cl_2(g)$ + 2KOH → KClO(aq) + NaCl(aq) + $H_2O(l)$

(Potassium hydroxide) (Potassium Chlorate (I))

(b)The experiment in 17(a) was repeated with hot concentrated sodium hydroxide solution. Explain the observation made.

Observation

Pale green colour of chlorine fades.

Pale yellow solution is formed.

Explanation

Chlorine reacts with hot concentrated Sodium hydroxide/Potassium hydroxide solution to form pale yellow solution of metal chlorate (v) and chlorides of metals.

Chemical equation

$3Cl_2(g) + 6NaOH(aq) \rightarrow NaClO_3(aq) + 5NaCl(aq) + 3H_2O(l)$

(Sodium hydroxide) (Sodium Chlorate (V))

$3Cl_2(g) + 6KOH(aq) \rightarrow KClO_3(aq) + 5KCl(aq) + 3H_2O(l)$

(Potassium hydroxide) (Potassium Chlorate (V))

The products formed when chlorine reacts with alkalis depend thus on temperature and the concentration of alkalis.

(c) (i) Write the equation for the formation of calcium chlorite (I) and calcium chlorate (V).

$2Ca(OH)_2(aq) + 2Cl_2(g) \rightarrow CaCl_2(aq) + CaOCl_2(aq) + H_2O(l)$

(Calcium hydroxide) (Calcium Chlorate(I))

(Cold/dilute)

$Ca(OH)_2(aq) + Cl_2(g) \rightarrow CaCl_2(aq) + Ca(ClO_3)_2(aq) + H_2O(l)$

(Calcium Chlorate(V))

B: THE HALOGENS

a) What are halogens?

These are elements in group VII of the periodic table. They include:

Element	Symbol	Atomic number	Electric configuration	Charge of ion	Valency	State at Room Temperature
Fluorine	F	9	2:7	F^-	1	Pale yellow gas
Chlorine	Cl	17	2:8:7	Cl^-	1	Pale green gas
Bromine	Br	35	2:8:18:7	Br^-	1	Red liquid

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Iodine	**I**	**53**	**2:8:18:18:7**	**I^-**	**1**	**Grey Solid**
Astatine	**At**	**85**	**2:8:18:32:18:7**	**At^-**	**1**	**Radioactive**

b) Compare the atomic radius and ionic radius of chloride ion and chlorine. Explain.

The radius of chlorine is smaller than the ionic radius o the chloride ion.
Effective nucleus attraction on outer energy level in chloride ion is less than chlorine atom because of extra gained electron gained electron that repelled thus causes the outer energy level to expand/increase.

c) Compare the atomic radius of chlorine and fluorine Explain.

Atomic radius of Fluorine is smaller than that of chlorine.
Chlorine has more energy levels than fluorine occupied by more electrons.

d) Chlorine is a gas, Bromine is a liquid, Iodine is a solid. Explain the above observations.

-Bromine, Chlorine and iodine exists as diatomic molecules bonded by strong covalent bond. Each molecule is joined to the other by weak intermolecular forces/ Van-der-waals forces.
-The strength of intermolecular/Van-der-waals forces of attraction increase with increase in molecular size/atomic radius.Iodine has therefore the largest atomic radius and thus strongest intermolecular forces to make it a solid.

e) (i) What is electronegativity?

Electronegativity is the tendency/ease of acquiring /gaining electrons by an element during chemical reaction.
It is measured using Pauling's scale.
Fluorine with Pauling scale 4.0 is the most electronegative element in the periodic table and thus the highest tendency to acquire/gain extra electron.

(ii) The table below shows the electronegativity of the halogens.

Halogen	F	Cl	Br	I	At
Electronegativity (Pauling's scale)	**4.0**	**3.0**	**2.8**	**2.5**	**2.2**

Explain the trend in electronegativity of the halogens.

Decrease down the group from fluorine to Astatine
Atomic radius increase down the group decreasing electron – attracting power down the group from fluorine to astatine.

(f) (i)What is electron affinity

Electron affinity is the energy required to gain an electron in an atom of an element in its gaseous state.

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(ii) Study the table below showing the election affinity of halogens for the process

$x + e \rightarrow x^-$

Halogen	F	Cl	Br	I
Electron affinity $kJmole^{-1}$	-333	-364	-342	-295

(iii) Explain the trend in electron affinity of the halogens.

-Decrease down the group

-Atomic radius of halogens increase down the group thus incoming/gained electron is attracted less strongly by the progressively larger atoms with a decreasing effective nuclear charge on outer energy level

(iv) Which is a move stable ion Cl^- or Br^- explain?

-Cl^- ion.

-Has a more negative/exothermic electron affinity than Br^-

(v) Differentiate between electron affinity and:

I. Ionization energy.

Ionization energy is the energy required to lose /donate an electron in an atom of an element in its gaseous state while electron affinity is the energy required to gain/acquire extra electron by an atom of an element in its gaseous state. Both are measured in kilojoules per mole.

II. Electronegativity.

-Electron affinity is the energy required to gain an electron in an atom of an element in gaseous state. It involves the process:

$X(g) + e \rightarrow X^-(g)$

Electronegativity is the ease/tendency of gaining/ acquiring electrons by an element during chemical reactions.

It does not involve use of energy but theoretical arbitrary Pauling's scale of measurements.

(g) (i) $5cm^3$ of sodium chloride, Sodium bromide and Sodium iodide solutions were put separately in test tubes. 5 drops of chlorine water was added to each test tube: state and explain the observation made.

Observation

Yellow colour of chlorine water fades in all test tubes expect with sodium chloride.

-Coloured Solution formed.

Explanation

Chlorine is more electronegative than bromine and iodine. On adding chlorine water, bromine and Iodine are displaced from their solutions by chlorine.

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(ii) The experiment in g (i) was repeated with 5 drops of bromine water instead of chlorine water .explain the observation made.

Observation

Yellow colour of bromine water fades in test tube containing sodium iodide.

Brown solution formed in test tube containing sodium iodide

Explanation

Bromine is more electronegative than iodide but less 6than chlorine.

On adding Bromine water, iodide displaced from its solution but not chlorine.

(iii) Using the knowledge in g(i) and (ii) above,

I. Complete the table below using (X) to show no reaction and (√) to show a reaction.

Halide ion / Halogen ion in solution Halogen	F^-	Cl^-	Br^-	I^-
F_2	X	√	√	√
Cl_2	X	X	√	√
Br_2	X	X	X	√
I_2	X	X	X	√

Write an ionic equation for the reaction where there is (V)

$F_{2\,(g)} + 2Cl^-_{(aq)} \rightarrow 2F^-_{(aq)} + Cl_{2(g)}$

$F_{2\,(g)} + 2Br^-_{(aq)} \rightarrow 2F^-_{(aq)} + Br_{2(aq)}$

$F_{2\,(g)} + 2I^-_{(aq)} \rightarrow 2F^-_{(aq)} + I_{2(aq)}$

$Cl_{2\,(g)} + 2Br^-_{(aq)} \rightarrow 2Cl^-_{(aq)} + Br_{2(aq)}$

$Cl_{2\,(g)} + 2I^-_{(aq)} \rightarrow 2Cl^-_{(aq)} + I_{2(aq)}$

$Br_{2\,(aq)} + 2I^-_{(aq)} \rightarrow 2Br^-_{(aq)} + I_{2(aq)}$

(h) State one uses of:

(iv) Fluorine

Manufacture of P.T.F.E (Poly tetra fluoroethene) synthetic fiber.

Reduce tooth decay when added in small amounts/equations in tooth paste.

Note: large small quantities of fluorine /fluoride ions in water cause browning of teeth/flourosis.

Hydrogen fluoride is used to engrave word pictures in glass.

(v) Bromine

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Silver bromide is used to make light sensitive photographic paper/films.

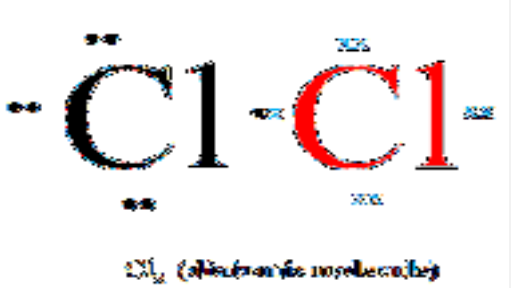
(vi) Iodide

Iodine dissolved in alcohol is used as medicine to kill bacteria in skin cuts. It is called tincture of iodine.

(ii) The table below to show some compounds of halogens.

Element / Halogen	H	Na	Mg	Al	Si	C	P
F	**HF**	**NaF**	**MgH_2**	**AlF_3**	**SiF_4**	**CF_4**	**PF_3**
Cl	**HCl**	**NaCl**	**MgCl**	**$AlCl_3$**	**$SiCl_3$**	**CCl_4**	**PCl_3**
Br	**HBr**	**NaBr**	**$MgBr_2$**	**$AlBo_3$**	**$SiBr_4$**	**CBr_4**	**PBr_3**
I	**Hl**	**Nal**	**Mgl_2**	**All_3**	**SiL_4**	**Cl_2**	**Pb_3**

(j) (i) Using dot (.) and Cross (x) to represen



(iii)Name the type of bond formed.

Covalent.

(iv) Below is the table showing the bond energy of four halogens.

Bond	Bond energy k J $mole^{-1}$
Cl-Cl	242
Br-Br	193
I-I	151

III. What do you understand by the term "bond energy"

Bond energy is the energy required to break/ form one mole of chemical bond

IV. Explain the trend in bond Energy of the halogens above:

-Decrease down the group from chlorine to Iodine

-Atomic radius increase down the group decreasing the energy required to break the covalent bonds between the larger atom with reduced effective nuclear charge an outer energy level that take part in bonding.

(k) Some compounds of chlorine are in the table below the oxidation state of chlorine in each compound.

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Compound	Oxidation state	Name of compound
$NaClO_3$	**+5**	**Sodium chlorate (V)**
ClO_2	**+4**	**Chloric (IV) oxide**
$KClO_2$	**+3**	**Potassium chlorate (III)**
NaClO	**+1**	**Sodium Chlorite (I)**
Cl_2	**0**	**Chlorine Molecule**
NaCl	**-1**	**Sodium Chloride (I)**
$MgCl_2$	**-1**	**Magnesium Chloride (I)**

C. HYDROGEN CHLORIDE

a) Occurrence

Hydrogen Chloride does not occur free in the atmosphere or in nature

b) Preparation

Hydrogen chloride may be prepared in the school laboratory by reacting solid sodium/potassium chloride crystals with concentrated sulphuric (Vi) acid as in the set up below.

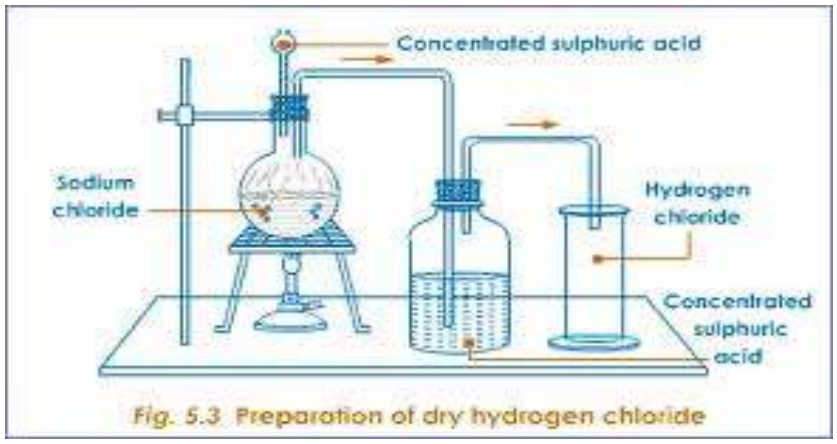


Fig. 5.3 Preparation of dry hydrogen chloride

c) Properties of hydrogen chloride gas(questions)

1. What precautions should be taken when handling concentrated sulphuric acid? Explain.

-Wear protective clothing/gloves to avoid accidental contact with skin.

-Concentrated sulphuric (VI) acid is highly corrosive-it causes painful wounds when in contact with skin.

2. What method of gas collection is used? Explain.

-Downward delivery// upward displacement of water

-Hydrogen chloride is denser than air.

3. a) Write the equation for the reaction that takes place.

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$NaCl(s) + H_2SO_4(l) \rightarrow NaHSO_4(aq) + HCl(g)$
$KCl(s) + H_2SO_4(l) \rightarrow KHSO_4(aq) + HCl(g)$
NaCl is commonly used because it is cheaper than KCl

b) What property of concentrated sulphuric (VI) acid is used during the above reaction
-is the least volatile mineral acid, thus displace the more volatile hydrogen chloride from its salt (KCl/NaCl)
d)i)What is the purpose of concentrated sulphuric (VI) acid.
-Drying agent / to dry the gas.
ii) What property of concentrated sulphuric (VI) acid is used during the above use.
-Is hygroscopic – absorbs water but do not form solution.
iii) Name another substance which can be used to dry chlorine gas.
-anhydrous Calcium chloride
- silica gel
iv)Using a chemical equation, explain why anhydrous calcium oxide cannot be used in flask B
-Calcium oxide reacts with water /moisture to form calcium hydroxide. The calcium hydroxide formed reacts with chlorine to form calcium hypochlorite.
Chemical equations:
$CaO(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2O(l)$
$Ca(OH)_2(aq) + Cl_2(g) \rightarrow CaOCl_2(aq) + H_2O(l)$
This reduces the amount of Chlorine produced.
d)Blue and red litmus papers were dipped in the hydrogen chloride prepared above. The Procedure was repeated with damp/wet/moist litmus papers. Explain the differences in observations made.
-Dry blue litmus papers remain blue
-Dry red litmus papers remain red
-Damp/moist/wet blue litmus papers turn red
-Damp/moist/wet red litmus paper turns red.
-Dry hydrogen chloride is a molecular compound that is joined by covalent bonds between the atoms. The gas is polar thus dissolves in water and ionize completely to free H^+ that are responsible to turning blue litmus paper red.
e) Dry hydrogen chloride gas was bubbled in two separately beakers containing water and in methylbenzene.
(i) Classify the two solvents as either "polar" or "non-polar"
Water – polar
Methylbenzene – non-polar
(ii) State and explain the observations made in the beaker containing:
(i)Methylbenzene

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Colour of litmus solution remain.
Hydrogen chloride is a molecular substance. When dissolved in non-polar solvent, it does not dissociate / ionize to release H^+ ions that changes the colour of litmus solution.

(ii)Water
Colour of litmus solution change to red.
Hydrogen chloride is a molecular substance. When dissolved in polar solvent like water, it dissociate/ionize to release H^+ ions that changes litmus solution to red.
(iii)Why should an inverted filter funnel be used to dissolve hydrogen chloride.
- The filter funnel is dipped just below the water surface to increase the surface area of dissolving the gas and prevent suck back.
(iv)Name the solution formed when hydrogen chloride dissolves in water.
Hydrochloric acid
(f) Describe the test for presence of hydrogen chloride gas.
-Dip a glass rod in ammonia. Bring it to the mouth of a gas jar containing a gas suspected to be hydrogen chloride
-White fumes of ammonia chloride are formed.
(g) Place 5cm^3 of dilute hydrochloric acid into a four separate test tubes. To separate test tube add zinc, magnesium iron and copper metals. State and explain the observations made.
Observation
– Effervescence/bubbles/fizzing in all cases except copper
- Colourless solution formed with zinc and magnesium.
- Green solution formed with ion.
- Gas produced that extinguishes splint with explosion.
Explanation
Metals above hydrogen in reactivity series react with hydrochloric and liberating hydrogen gas.
Chemical Equation:
Concentrated hydrochloric acid is a weak oxidizing agent than other concentrated acids i.eSulphuric (VI) acid and nitric (V) acid that react with all metals even those lower in the reactivity series.
(h) Place 5cm^3 of dilute hydrochloric acid into five separate test tubes. To separate test tubes, add calcium carbonate, silver carbonate, copper carbonate, iron (II) carbonate and Sodium hydrogen carbonate. Explain the observations made.
Observation
Effervescence/bubbles/fizzing vigorously except in silver carbonate and lead (II) carbonate that stop later.

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- **Colourless solution formed except with iron (II) carbonate and copper (II) carbonate**
- **Green solution formed with iron (II) carbonate**
- **Blue solution formed with copper (II) carbonate**

Explanation.

Carbonates and hydrogen carbonate react with dilute hydrochloric acid to produce carbon (IV) oxide, water and form chlorides.

All chlorides formed are soluble Except Lead (II) Chloride (soluble on heating/warming) and silver chloride.

Chemical equation:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

(Colourless solution)

Chemical equation:

$$Ag_2CO_3(s) + 2HCl(aq) \rightarrow AgCl(s) + H_2O(l) + CO_2(g)$$

(Coats/Cover Ag_2CO_3)

Chemical equation:

$$CuCO_3(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2O(l) + CO_2(g)$$

(Blue Solution)

Chemical equation:

$$FeCO_3(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2O(l) + CO_2(g)$$

Chemical equation:

$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

Place 5cm^3 of dilute sodium hydroxide, Potassium hydroxide and aqueous ammonia solution into three separate test tubes. Add one drop of phenolphthalein indicator drop wise, add dilute hydrochloric acid. Explain the observations made.

Observation

Colour of Phenophthalein indicator change from pink to colourless.

Explanation

Hydrochloric acid neutralizes alkalis to salt and water

When all the alkali has reacted with the acid, An extra slight excess acid turns the indicator used to colourless.

Chemical equation:

$$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$$

Chemical equation:

$$KOH(aq) + HCl(aq) \rightarrow KCl(aq) + H_2O(l)$$

Chemical equation:

$$NH_4OH(aq) + HCl(aq) \rightarrow NHaCl(aq) + H_2O(l)$$

(j) Place 5cm^3 of hydrochloric acid into four separate test tube tubes Separately add about 1g of each of copper (II) Oxide, Zinc (II) Oxide, Lead (II) Oxide< Calcium (II) Oxide. What happens to each test tube? Explain.

Observation:

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All Solid dissolves except Lead (II) Oxide

Colourless solution formed with zinc Oxide and calcium (II) Oxide blue solution formed with copper (II) Oxide.

Explanation:

Metal oxides dissolves in dilute hydrochloric acid to form water and chloride salt Insoluble Lead (II) chloride and silver chloride once formed cover/coat unreacted oxides stopping further reaction.

Chemical equation: **$CuO(s) + HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l)$**

Chemical equation: $CaO(s) + HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l)$

Chemical equation: $PbO(s) + 2HCl(aq) \rightarrow PbCl_2(aq) + H_2O(l)$

Chemical equation: $ZnO(s) + HCl(aq) \rightarrow ZCl_2(aq) + H_2O(l)$

(k) Manufacture of Hydrochloric acid.

(i) Raw Materials

1. Hydrogen

(i) During **electrolysis** of Brine from the flowing mercury-cathode cell during the manufacture of sodium hydroxide solution.

(ii)From **water gas** by passing steam in heated charcoal.

$$C(s) + H_2O \rightarrow CO(g) + H_2(g)$$

(iii)From partial oxidation of natural gas/methane

$$CH_4(g) + O_2(g) \rightarrow CO(g) + 3H_2(g)$$

2.Chlorine

(i)From electrolysis of **fused**/solid sodium chloride in the downs process during extraction of sodium

(ii)From electrolysis of **brine**/concentrated sodium chloride solution in the flowing mercury-cathode during the manufacture of sodium hydroxide solution.

(ii)Chemical processes.

Hydrogen and chlorine gases are separately passed through concentrated sulphuric(VI) acid to act as a **drying** agent.

Small amount of pure hydrogen is continuously ignited in a chamber with **continous** supply of pure dry chlorine.

Large amount of hydrogen explodes.

Hydrogen burns in chlorine to form hydrogen chloride gas.

Chemical Equation

$$H_2(g) + Cl(g) \rightarrow 2HCl(g)$$

The hydrogen chloride produced is then passed up to meet a downward flow of water in the **absorbtion** chambers. Hydrogen chloride is very soluble in water and dissolves to form 35% concentrated hydrochloric acid.

Chemical Equation

$$HCl(\mathbf{g}) + (aq) \rightarrow HCl(\mathbf{aq})$$

The absorption chamber is shelved and packed with broken glass beads to

(i)Slow down the downward flow of water.

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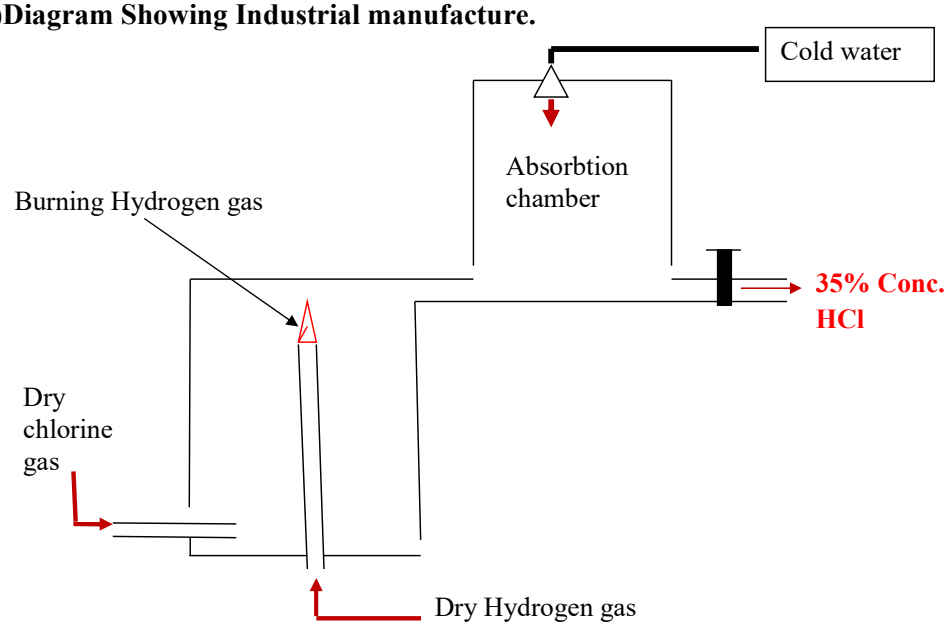
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(ii)Increase surface area over which the water dissolves
The hydrochloric acid is then transported in steel tanks lined with rubber for market

(iii)Uses of Hydrochloric Acid

- To standardize the pH of (alcohol and wines)
- Regenerating ion-exchange resin during removal of hardness of water.
- Pickling of metals to remove oside layers on their surfaces.
- In the manufacture of dyes and drugs.
- Making zinc chloride for making dry cells.

(iv)Diagram Showing Industrial manufacture.



(ii)Environmental effects of manufacturing HCl

- Hydrochloric acid is acidic. Any leakage from a manufacturing plant to nearby rivers/lake causes exess acidity that lowers pH of water killing marine life.
- Hydrogen chloride leakage into atmosphere dissolves to form "acidic rain" that accelerate corrosion in buildings, Breathing problems to human beings and kill fauna and flora around the paint.
- Chlorine leakage causes breathing and sight problems to human being. It accelerates bleaching of dyed metals.
- Hydrogen leakage can cause an explosion because impure hydrogen explodes on ignition.

(ii) Factors considered in setting hydrochloric acid manufacturing plant.

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1. Nearness to the manufacturing of sodium hydroxide because the byproducts of electrolysis of brine are the raw materials for hydrochloric acid plant.
2. Availability of natural gas for extraction of hydrogen.
3. Nearness/Availability of water to dissolve the hydrogen chloride gas.
4. Availability of labour, market, capital and good means of transport.

D: CHLORIDE (Cl^-) SALTS

(a)Occurrence.

1. Chlorides are salts derived from hydrochloric acid. Hydrochloric acid is a monobasic (HX) salt with only one ionazable/replaceable "H" in its molecule. All chlorides are therefore normal salts.
2. All metals exist as chloride salt except platinum and gold as below

Metal	**K**	**Na**	**Li**	**Mg**	**Ca**	**Al**	**Zn**	**Fe**	**Pb**	**H.**	**Cu**	**Ag**	**Hg**
Formula of chloride	KCl	NaCl	LiCl	$MgCl_2$	$CaCl_2$	$AlCl_3$	$ZnCl_2$	$FeCl_2$ $FeCl_3$	PbCl $PbCl_4$	HCl	CuCl $CuCl_2$	AgCl	Hg_2Cl_2 $HgCl_2$

(**i**)Both $FeCl_2$ and $FeCl_3$ exists but $FeCl_2$ is readily oxidized to $FeCl_3$ because it is more stable.

(**ii**)$PbCl_2$ and $PbCl_4$ exist but $PbCl_4$ is only oxidized to form $PbCl_2$ by using excess chlorine. It is less stable.

(**iii**)CuCl and $CuCl_2$ exists but $CuCl_2$ is (thermodynamically) more stable than CuCl. CuCl disproportionate to Cu and $CuCl_2$..

(**iv**)HgCl and $HgCl_2$exists as molecular compounds.

3. All chlorides are soluble/dissolves in water **except** silver chloride(AgCl), Copper (I) chloride CuCl, Mercury (I) Chloride Hg_2Cl_2 and Lead (II) Chloride $PbCl_2$ that dissolves in **warm** water.
4. Most chlorides are very stable compounds. They **do not** decompose on gentle or strong bunsen burner heating in a school laboratory except **Ammonium Chloride**.

5. Heating ammonium chloride

Place about 2g of solid ammonium chloride crystals in a clean dry boiling tube.Heat gently then strongly.

Observation

-red **litmus paper turn blue**

-blue litmus paper remains blue

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Then later:

-both blue litmus papers turn red

Explanation:

Ammonium chloride on heating decomposes through **chemical sublimation** to ammonia and hydrogen chloride gas. Ammonia gas is less dense than hydrogen chloride. It is a **basic** gas and diffuses out **faster** to turn red litmus paper to **blue**. Hydrogen chloride is an **acidic** gas .It is denser than ammonia gas and thus diffuses **slower** than ammonia gas to turn the already both blue litmus paper to **red**.

Chemical equation

$$NH_4Cl(s) \rightarrow HCl(g) + NH_3(g)$$

(**acidic** gas) (**basic/alkaline** gas)

(b) Test for Cl^- ions

1. The following experiment shows the test for the presence of Cl^- ions in **solids** chloride salts.

   (a)Procedure:

   Place about 1g of sodium chloride, Zinc chloride and copper (II) chloride in separate boiling tubes. Place moist blue and red litmus papers on the mouth of the test tube. Carefully, add three drops of concentrated sulphuric (VI) acid. Dip a glass rod in aqueous ammonia solution then bring it to the mouth of the boiling tube.

observation	inference
-red litmus paper remain red **-blue litmus paper turn red**	**H^+ ions**
-vigorous effervescence/fizzing /bubbling	**Cl^- ions**
-white fumes produced on	**HCl gas suspected**

(b)Explanation:

Concentrated sulphuric (VI) acid is the less volatile mineral acid.

It vigorously displaces chlorine in metal chlorides to evolve acidic hydrogen chloride gas fumes.

Chemical equation

$$NaCl(s) + H_2SO_4(l) \rightarrow NaHSO_4(aq) + HCl(g)$$
$$KCl(s) + H_2SO_4(l) \rightarrow KHSO_4(aq) + HCl(g)$$
$$CuCl_2(s) + H_2SO_4(l) \rightarrow CuSO_4(aq) + 2HCl(g)$$
$$ZnCl_2(s) + H_2SO_4(l) \rightarrow ZnSO_4(aq) + 2HCl(g)$$

Hydrogen chloride and ammonia gases react and form **white fumes** of ammonium chloride that confirms presence of **Cl^-** ions in the **solid** substance.

Chemical equation

$$NH_3(\mathbf{g}) + HCl(\mathbf{g}) \rightarrow NH_4Cl(\mathbf{s})$$

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2. The following experiment shows the test for the presence of **Cl^- ions** in **solution /aqueous** chloride salts.

(i)Using aqueous Lead (II) nitrate(V)

(a)Procedure:

I.Place about 5cm3 of sodium chloride, Iron (III) chloride and copper (II) chloride in separate boiling tubes. Add four drops of Lead (II) nitrate(V) solution to each. Preserve.

Observation	**Inference**
White precipitate/ppt	**SO_4^{2-}, SO_3^{2-}, Cl^-,CO_3^{2-}**

II.To the preserved sample, add six drops of nitric (V) acid. Preserve.

Observation	**Inference**
White precipitate/ppt persist	**SO_4^{2-}, Cl^-**

III. To the preserved sample, heat the mixture to boil

Observation	**Inference**
White precipitate/ppt dissolves on boiling/warming	**Cl^-**

Explanation:

I.When Lead(II) nitrate(V) solution is added to an unknown salt , a **white precipitate/ppt** of Lead(II) sulphate(VI) Lead(II) carbonate(IV) Lead(II) sulphate(IV) Lead(II) chloride(I) are formed.

Ionic equation:

$Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$

$Pb^{2+}(aq) + SO_3^{2-}(aq) \rightarrow PbSO_3(s)$

$Pb^{2+}(aq) + CO_3^{2-}(aq) \rightarrow PbCO_3(s)$

$Pb^{2+}(aq) + Cl^-(aq) \rightarrow PbCl_2(s)$

II.When the white precipitate/ppt formed is acidified with dilute **nitric(V) acid**, the white precipitate of Lead(II) sulphate(VI) and Lead(II) chloride(I) **persist/remain** while that of Lead(II) carbonate(IV) and Lead(II) sulphate(IV) **dissolves**.

III.On **heating /warming** Lead (II) chloride (I) **dissolves** but on cooling it **recrystallizes**.This shows the presence of **Cl^- ions** in **aqueous** solutions

. **(ii)Using aqueous silver (I) nitrate(V)**

Procedure

I. Place about 5cm3 of sodium chloride, Iron (III) chloride and copper (II) chloride in separate boiling tubes. Add four drops of silver(I) nitrate(V) solution to each. Preserve.

Observation	**Inference**

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White precipitate/ppt	Cl^-, CO_3^{2-}

II. To the preserved sample, add six drops of nitric (V) acid. Preserve.

Observation	Inference
White precipitate/ppt persist	Cl^-

Explanation:

I.When silver(I) nitrate(V) solution is added to an unknown salt , a **white precipitate /ppt** of silver(I) carbonate(IV) and silver(I) chloride(I) are formed.

Ionic equation:

$2Ag^+(aq) + CO_3^{2-}(aq) \rightarrow Ag_2CO_3(s)$

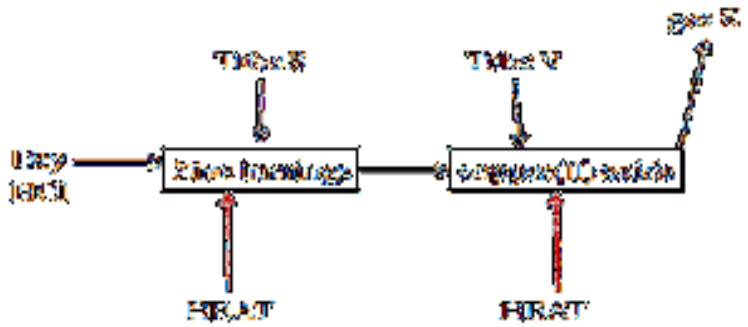
$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

II. When the white precipitate/ppt formed is acidified with dilute **nitric (V) acid**, the white precipitate of silver (I) chloride (I) **persist/remain**. This shows the presence of **Cl^-**ions in **aqueous solutions**.

Silver (I) carbonate (IV) **dissolves** when reacted with nitric (V) acid.

COMPREHENSIVE REVISION QUESTIONS

1. In an experiment, dry hydrogen chloride gas was passed through heated zinc turnings as in the set up below. The gas produced was the passed through copper(II) oxide



a) Write the equation for the reaction :

(i)For the school laboratory preparation of hydrogen chloride gas.

$NaCl(s) + H_2SO_4(l) \rightarrow NaHSO_4(aq) + HCl(g)$

(ii)in tube S

$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

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b)State and explain the observation made in tube V.

Observations-**colour of solid changes from black to brown**

-colourless liquid forms on the cooler parts of tube V

Explanation-**Hydrogen produced in tube S reduces black copper(II) oxide to brown copper metal and the gas oxidized to water vapour that condense on cooler parts..**

Chemical equation.

$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$

(c)How would the total mass of tube S and tube V and their contents compare before and after the experiment.

Tube S- **Mass increase/rise because Zinc combine with chlorine to form heavier Zinc Chloride.**

Tube V- **Mass decrease/falls/lowers because copper (II) oxide is reduced to lighter copper and oxygen combine with hydrogen to form water vapour that escape.**

2. Chlorine is prepared by using solid sodium chloride, concentrated sulphuric(VI) acid and potassium manganate(VII)

a)What is the role of the following in the reaction;

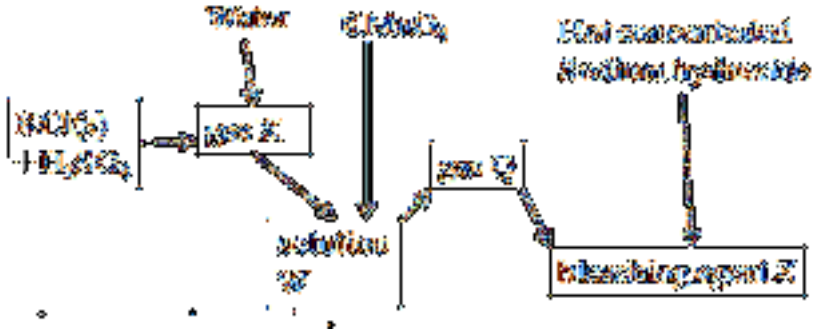
(i) concentrated sulphuric(VI)

To produce hydrogen chloride gas by reacting with the solid sodium chloride.

(ii) Potassium manganate(VII)

To oxidize hydrogen chloride gas to chlorine

3.Use the flow chart below to answer the questions that follow.



a)(i) Name:

gas X **Hydrogen chloride**

solution W **hydrochloric acid**

gas Q **chlorine**

bleaching agent Z**sodium chlorate(V)**

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b)Write the chemical equation for the formation of :

(i) gas X

$NaCl(s) + H_2SO_4(l) \rightarrow NaHSO_4(aq) + HCl(g)$

(ii)solution W

$HCl(g) + (aq) \rightarrow HCl(aq)$

(iii)gas Q

$2KMnO_4 + 16HCl(aq) \rightarrow 2KCl(aq) + 2MnCl_2(aq) + 8H_2O(l) + 5Cl_2(g)$

(iv)bleaching agent Z

$6NaOH(aq) + 3Cl_2(g) \rightarrow NaCl(aq) + NaClO_3(aq) + 3H_2O(l)$

c)State and explain the following observations;

(i) a glass rod dipped in aqueous ammonia is brought near gas X

Observation: Dense white fumes

Explanation:Ammonia gas reacts with hydrogen chloride gas to form dense white fumes of ammonium chloride.

Chemical equation: $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$

(ii)Wet blue and red litmus papers were dipped into gas Q

Observations: Blue litmus paper turned red the both are bleached /decolorized.

Explanations: chlorine reacts with water to form both acidic hydrochloric and chloric (I) acids that turn blue litmus paper red. Unstable chloric (I) acid oxidizes the dye in the papers to colourless.

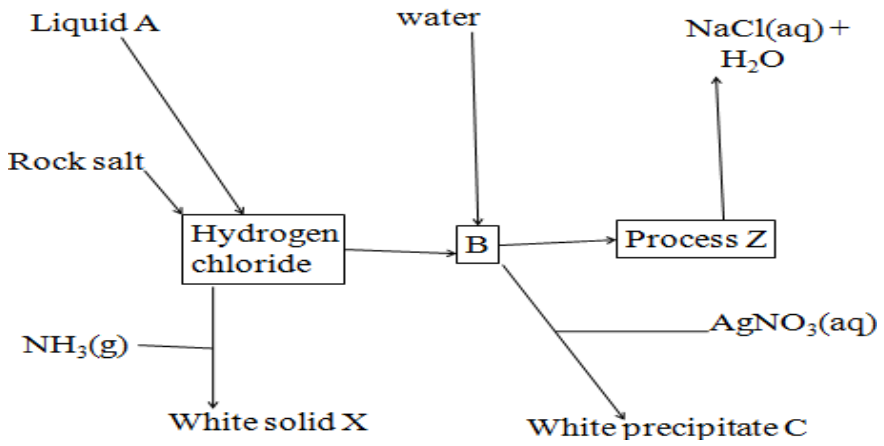
Chemical equations

$Cl_2(g) + HCl(aq) \rightarrow HCl(aq) + HClO(aq)$

Coloured dye $+HClO(aq) \rightarrow HCl(aq) +$ (Colourless dye +O)//

(Coloured dye-O) $+ HClO(aq) \rightarrow HCl(aq) +$ Colourless dye

4.Use the flow chart below to answer the questions that follow



a) Name

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Liquid A **Concentrated sulphuric(VI) acid**
Process Z **Neutralization**
White solid X **Ammonium chloride**

b)Write the equation for the formation of:

(i) Hydrogen chloride

$NaCl(s) + H_2SO_4(l) \rightarrow NaHSO_4(aq) + HCl(g)$

(ii) B

$HCl(g) + (aq) \rightarrow HCl(aq)$

(iii)process Z (using ionic equation)

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

(iv)C (using ionic equation)

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

c)Describe how solution B is obtained.

Bubbling hydrogen chloride gas through inverted funnel into distilled water until no more dissolve.

5 The results obtained when halogens are bubbled into test tubes containing solutions of halide A,B and C is as in the table below. Tick(v) means a reaction took place. Cross(x) means no reaction took place.

Halogens	Halide ions in solution		
	A	B	C
I_2	x	-	x
Br_2	x	v	-
Cl_2	-	v	v

a)Identify the halide ions represented by letter

A **Cl^-**
B **I^-**
C **Br^-**

b)Write the ionic equation for the reaction that take place with halide:

(i) C

$Cl_2(g) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)$

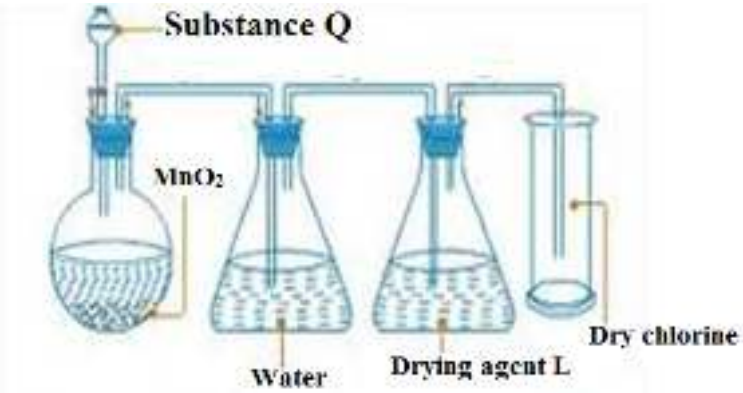
(ii) B

$Cl_2(g) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)$
$Cl_2(g) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$

6.The diagram below shows a set up of apparatus for the school laboratory collection of dry chlorine gas.

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a)Name:

(i) Substance Q

Concentrated hydrochloric acid

(ii) Suitable drying agent L

-Concentrated sulphuric (VI) acid

-anhydrous calcium chloride

-silica gel

b) State a missing condition for the reaction to take place faster.

-Heat/Heating

c) Red and blue litmus papers were dipped into the chlorine gas from the above set up .State and explain the observations made.

Observation: Blue litmus paper remains blue. Red litmus paper remain red.

Explanation: Dry chlorine has no effect on dry litmus papers.

d) Write the equation for the reaction taking place in the conical flask

MnO_4 (s) + 4HCl(aq) -> $MnCl_2$(aq) + $2H_2O$(l) + Cl_2(g)

e) Name two other substances that can be used in place of MnO_2

Lead(IV) oxide (PbO_2)

Potassium manganate(VI)($KMnO_4$)

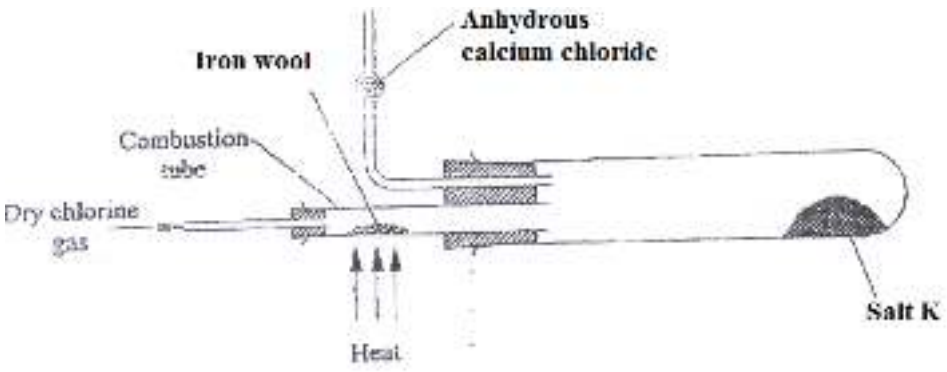
Potassium dichromate($K_2Cr_2O_4$)

Bleaching powder($CaOCl_2$)

7. The set up below shows the apparatus used to prepare and collect anhydrous iron(III) chloride.

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a)Name salt K

Iron(III)cchloride

b) Write the equation for the reaction for the formation of salt K

$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s/g)$

State and explain the following

(i)Small amount of water is added to iron (II) chloride in a test tube then shaken

Solid dissolves to form a green solution. Iron(II) chloride is soluble in water

(ii)I.Three drops of aqueous sodium hydroxide is added to aqueous iron(II) chloride and then added excess of the alkali.

Observation:

Green precipitate is formed that persist/remain /insoluble in excess akali.

Explanation**:**

Iron(II) chloride reacts with aqueous sodium hydroxide to form a green precipitate of iron(II) hydroxide.

Ionic equation:

$Fe^{2+}(aq) + OH^-(aq) \rightarrow Fe(OH)_2(s)$

II.Six drops of hydrogen peroxide is added to the mixture in d(ii) above.

Observation:

Effervescence/bubbling/fizzing take place and the green precipitate dissolve to form a yellow/brown solution.

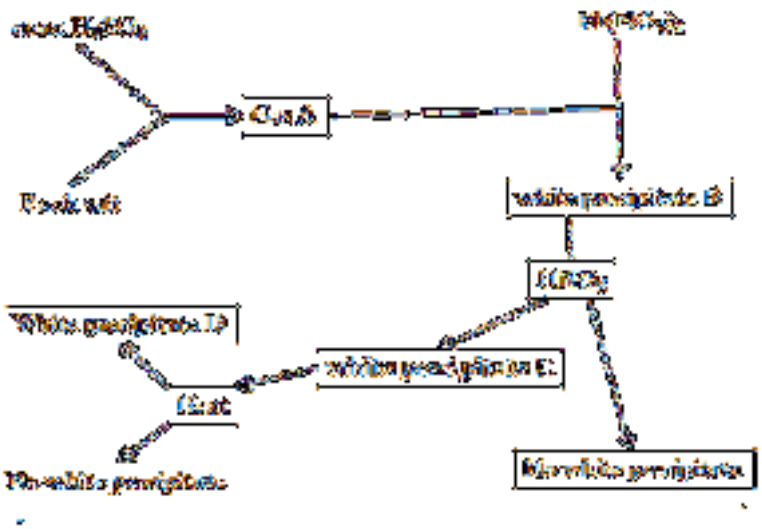
Explanation**:**

hydrogen peroxide oxidizes green $Fe^{2+}$to yellow/ brown Fe^{3+}solution.

9.Use the flow chart below to answer the questions that follow.

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a) Write the chemical equation for the formation of gas A

$NaCl(s) + H_2SO_4(l) \rightarrow NaHSO_4(aq) + HCl(g)$

b) Identify:

(i) four possible ions that can produce white precipitate B

SO_4^{2-},SO_3^{2-}, CO_3^{2-}, Cl^-

(ii)two possible ions that can produce;

I.White precipitate C

SO_4^{2-},Cl^-

II.colourless solution D

SO_3^{2-}, CO_3^{2-}

(iii)possible ions present in

I.White precipitate E

SO_4^{2-}

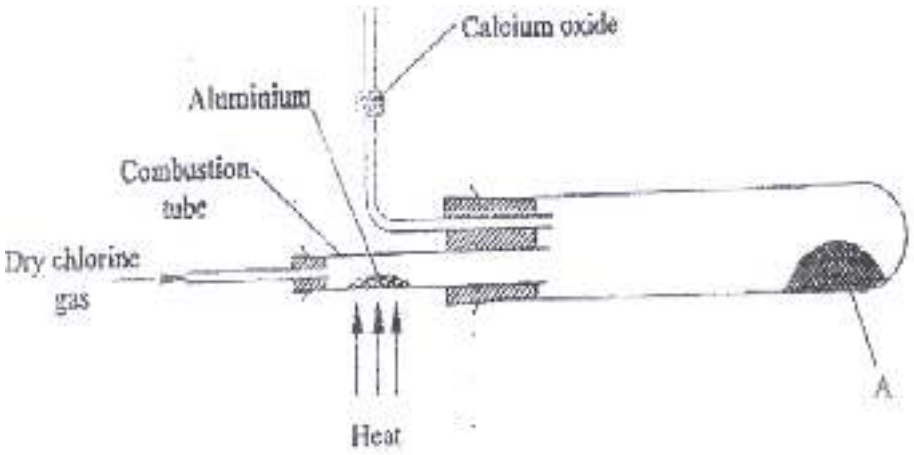
II.colourless solution F

Cl^-

11. Below is a set up in the preparation of a particular salt. Study it and answer the questions that follow.

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State the observation made when aluminium wool is heated.

Glows red hot.

b)(i) Identify salt A

aluminium(III) chloride// $AlCl_3$

(ii)Write the equation for the formation of salt A

$2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s/g)$

(iii)What property of salt A is exhibited as shown in the experiment.

It sublimes//sublimation.

(iv)Calculate the minimum volume of chlorine required to form 700kg of iron(III) chloride at room temperature.(Fe= 56.0, Cl=35.5, 1 mole of a gas =24000cm3, 1000g = 1kg)

Mole ratio Fe : Cl_2 = 2: 3 molar mass $FeCl_3$ = 162.5g

Method 1

2 x 162.5 g $FeCl_3$ -> 3x 22400 cm3 Cl_2

700 x1000 g$FeCl_3$ -> (700 x1000 x3 x22400)/(2 x 162.5)

$=1.4474 \times 10^{-8}$ cm3

Method 2

Moles of $FeCl_3$= mass/ molar mass

=> (700 x 1000) / 162.5 = 4307.6923 moles

Moles of Cl_2= 3/2 moles of $FeCl_3$

=>3/2 x 4307.6923 = 6461.5385 moles

Volume of chlorine= moles x molar gas volume

=>6461.5385 x 24000 = 1.5508×10^{-8} cm3

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c) Name another metal that can produce similar results as salt K.

Iron

d)(i) What is the purpose of anhydrous calcium chloride.

-ensure the apparatus are water free.

-prevent water from the atmosphere from entering and altering//hydrolysing salt A

(ii) Write the equation for the reaction that take place if anhydrous calcium chloride is not used in the above set up.

$AlCl_3(s) + 3H_2O(l) \rightarrow Al(OH)_3(aq) + 3HCl(g)$

(iii) Write the equation for the reaction that take place when Iron metal is reacted with dry hydrogen chloride gas.

$Fe(s) + 2HCl(g) \rightarrow FeCl_2(s) + H_2(g)$

(iv)Calculate the mass of Iron(II)chloride formed when 60cm3 of hydrogen chloride at r.t.p is completely reacted. (1 mole of a gas =24dm3 at r.t.p, Fe = 56.O, Cl= 35.5)

Chemical equation $Fe(s) + 2HCl(g) \rightarrow FeCl_2(s) + Cl_2(g)$

Mole ratio HCl: $FeCl_2$ = 1:1

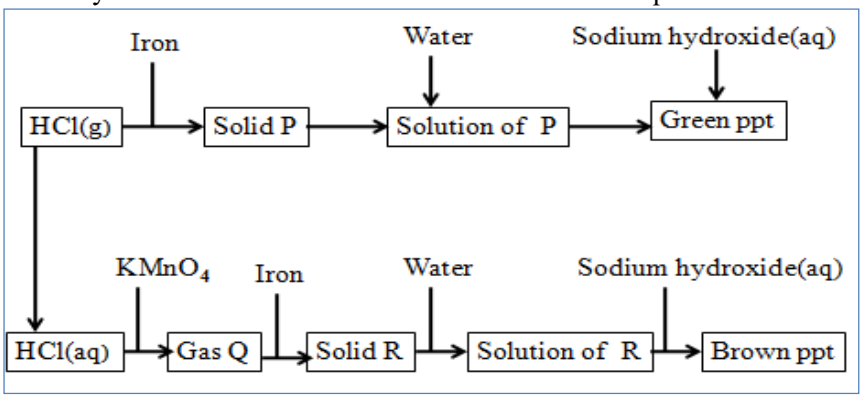
Molar mass $FeCl_2$ = 127g

Moles of HCl used = 60cm3 /24000cm3 = 2.5×10^{-3} moles

Moles of $FeCl_2$ = Moles of HCl $\Rightarrow 2.5 \times 10^{-3}$ moles

Mass of $FeCl_2$ = moles x molar mass $\Rightarrow 2.5 \times 10^{-3} \times 127 = 0.3175g$

12.Study the flow chart below and use it to answer the questions that follow



a)Identify substance:

P **Iron(II) chloride//$FeCl_2$**

Q **Chlorine // Cl_2**

R **Iron(III) chloride//$FeCl_3$**

b)Write the equation for the reaction for the formation of:

(i) gas Q

$2KMnO_4(s) + 16HCl(aq) \rightarrow 2KCl(aq) + 2MnCl_2(aq) + 8H_2O(l) + 5Cl_2(g)$

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(ii) the green precipitate (using ionic equation)

Ionic equation:

$Fe^{2+}(aq)$ + $2OH^{-}(aq)$ -> $Fe(OH)_2(s)$

(ii) the brown precipitate (using ionic equation)

Ionic equation:

$Fe^{3+}(aq)$ + $3OH^{-}(aq)$ -> $Fe(OH)_3(s)$

c)A glass rod was dipped in aqueous ammonia. The rod was then brought near hydrogen chloride. State and explain the observation made.

Observation:

White fumes

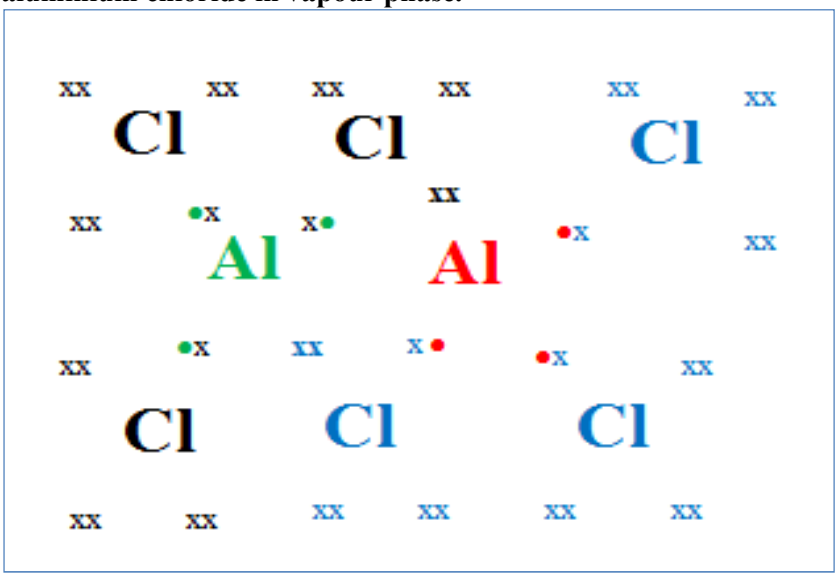
Explanation:

Ammonia gas reacts with hydrogen chloride gas to form white fumes of ammonium chloride.

Chemical equation:

$NH_3(g)$ + $HCl(g)$ -> $NH_4Cl(s)$

13. Using dot(.)and cross(x)to represent electrons,show the bonding in aluminium chloride in vapour phase.



(b)How many electrons in :

(i)aluminium atoms are used in bonding.

Six electrons(three valence electrons in each aluminium atom)

(ii)chlorine atoms atoms are used in dativebonding.

four electrons(two lone pairs of valence electrons in two chlorine atoms)

(iii)the molecule are used in bonding.

Sixteen electrons

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-six valence electrons from aluminium atom through covalent bond

-six valence electrons from chlorine atoms through covalent bond.

- four valence electrons from chlorine atoms through dative bond

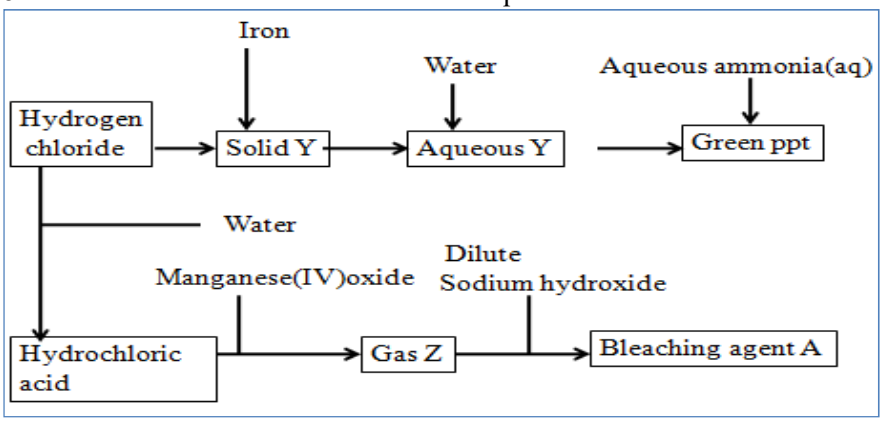
(c)How many lone pair of electrons do not take part in bonding within the molecule.

Sixteen(16) lone pairs from six chlorine atoms(32 electrons)

(d)Aluminium chloride does not conduct electricity in molten state but Magnesium chloride conduct.

Aluminium chloride is a molecular compound that has no free mobile Al^{3+} and Cl^- ions which are responsible for conducting electricity. Magnesium chloride has free mobile Mg^{2+} and Cl^- ions because it is an ionic compound.

8. Use the flow chart below to answer the questions that follow:



a)Write an equation for the school laboratory formation of hydrogen chloride gas

$NaCl(s) + H_2SO_4(l) \rightarrow NaHSO_4(aq) + HCl(g)$

$KCl(s) + H_2SO_4(l) \rightarrow KHSO_4(aq) + HCl(g)$

b)Name:

I. solid Y **Iron (II) chloride ($FeCl_2$)**

II green precipitate **Iron (II) hydroxide ($Fe(OH)_2$**

III Gas Y **Chlorine (Cl_2)**

IV. Bleaching agent A **Sodium hypochlorite (NaOCl)**

c)Blue and red litmus papers were dipped into bleaching agent A. Write the equation for the reaction that takes place.

Coloured dye +NaOCl(aq) ->NaCl(aq) + (Colourless dye + O)//

(Coloured dye-O) + NaOCl(aq) ->NaCl(aq) + Colourless dye

d)State four uses of gas Z

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1. **Bleaching agent**
2. **Manufacture of hydrochloric acid**
3. **Chlorination of water to kill germs**
4. **Manufacture of PVC pipes**

CHEMISTRY OF NITROGEN

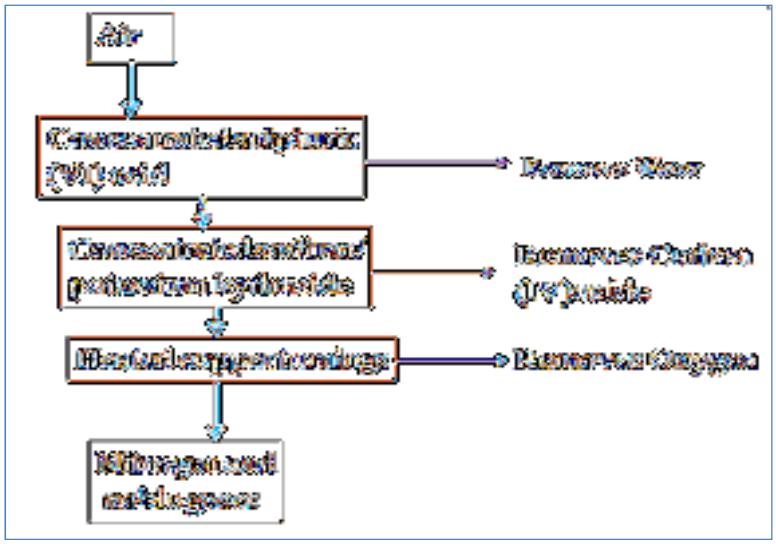
A.NITROGEN

a) Occurrence:

Nitrogen is found in the atmosphere occupying about 78% by volume of air. Proteins, amino acids, polypeptides in living things contain nitrogen.

b) Isolation of nitrogen from the air.

Nitrogen can be isolated from other gases present in air like oxygen, water (vapour), carbon (IV) oxide and noble gases as in the school laboratory as in the flow chart below:



Water is added slowly into an "empty flask" which forces the air out into another flask containing concentrated sulphuric (VI) acid. Concentrated sulphuric (VI) acid is hygroscopic. It therefore absorb/remove water present in the air sample.

More water forces the air into the flask containing either concentrated sodium hydroxide or potassium hydroxide solution. These alkalis react with carbon IV) oxide to form the carbonates and thus absorbs/remove carbon IV) oxide present in the air sample.

Chemical equation $2NaOH\ (aq) + CO_2\ (g) \rightarrow Na_2CO_3\ (aq) + H_2O(l)$

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Chemical equation $2KOH\ (aq) + CO_2\ (g) \rightarrow K_2CO_3\ (aq) + H_2O(l)$

More water forces the air through a glass tube packed with copper turnings. Heated brown copper turnings react with oxygen to form black copper (II) oxide.

Chemical equation $2Cu\ (s) + O_2\ (g) \rightarrow CuO\ (s)$

(brown) (black)

The remaining gas mixture is collected by upward delivery/downward displacement of water/over water. It contains about 99% nitrogen and 1% noble gases.

On a large scale for industrial purposes, nitrogen is got from fractional distillation of air.

c) Nitrogen from fractional distillation of air.

For commercial purposes nitrogen is got from the fractional of air.

Air is first passed through a dust precipitator/filter to remove dust particles.

The air is then bubbled through either concentrated sodium hydroxide or potassium hydroxide solution to remove/absorb Carbon(IV) oxide gas.

Chemical equation $2NaOH\ (aq) + CO_2\ (g) \rightarrow Na_2CO_3\ (aq) + H_2O(l)$

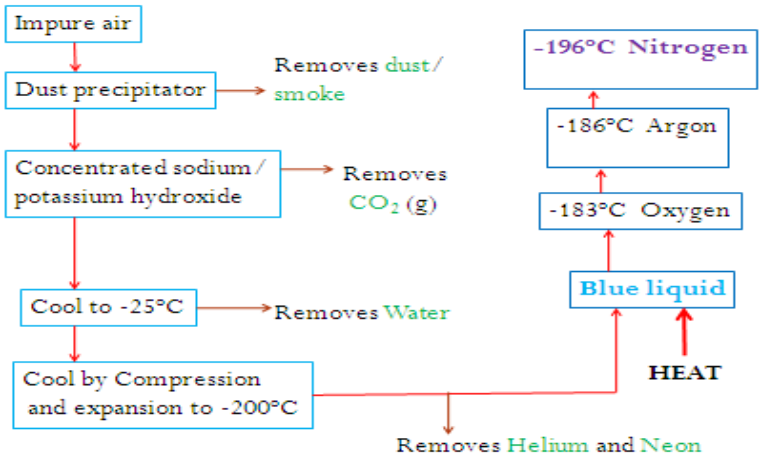
Chemical equation $2KOH\ (aq) + CO_2\ (g) \rightarrow K_2CO_3\ (aq) + H_2O(l)$

Air mixture is the cooled to -25°C.At this temperature, water (vapour) liquidifies and then solidify to ice and thus removed.

The air is further cooled to -200°C during which it forms a blue liquid.

The liquid is then heated. Nitrogen with a boiling point of -196°C distils first then Argon at-186°C and then finally Oxygen at -183°C boils last.

Flow diagram showing fractional distillation of air/industrial production of Nitrogen/Oxygen/Argon

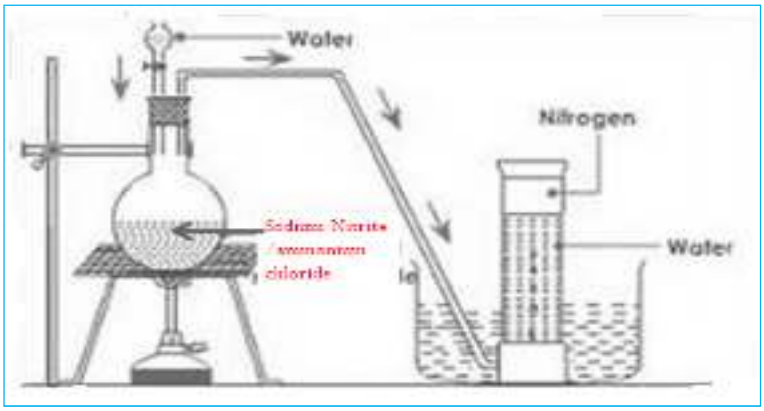


c) School laboratory preparation of Nitrogen.

The diagram below shows the set up of the school laboratory preparation of nitrogen gas.

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d.Properties of Nitrogen gas(Questions)

1.Write the equation for the reaction for the school laboratory preparation of nitrogen gas.

Chemical equation $NH_4Cl\ (s) + NaNO_2(s) \rightarrow NaCl\ (g) + NH_4NO_2\ (s)$

Chemical equation $NH_4NO_2\ (s) \rightarrow N_2\ (g) + H_2O\ (l)$

2. State three physical properties of nitrogen gas.

- colourless, odourless, less dense than air ,neutral and slightly soluble in water

3. State and explain the observation made when a burning magnesium ribbon is lowered in a gas jar containing nitrogen gas.

Observation; It continues burning with a blight blindening flame forming white ash.

Explanation

Magnesium burns to produce enough heat /energy to reacts with nitrogen to form white magnesium nitride.

Chemical equation $3Mg\ (s) + N_2\ (g) \rightarrow Mg_3N_2\ (s)$ (white ash/solid)

4. State two main uses of nitrogen gas

-manufacture of ammonia from Haber process

- As a refrigerant in storage of semen for Artificial insemination.

B. OXIDES OF NITROGEN

Nitrogen forms three main oxides:

i)Nitrogen(I) oxide(N_2O)

ii) Nitrogen(II) oxide (NO)

iii) Nitrogen (IV) oxide(NO_2)

i) Nitrogen (I) oxide(N_2O)

a) Occurrence

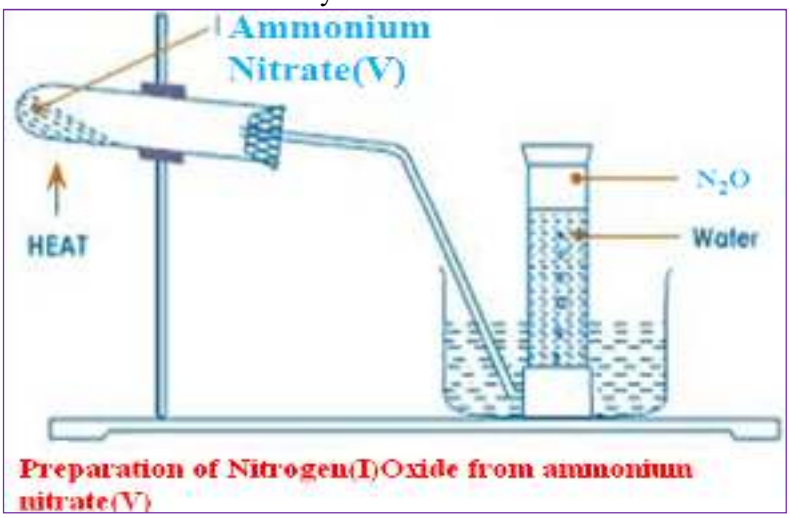
Nitrogen (I) oxide does not occur naturally but prepared in a laboratory.

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b)Preparation

The set up below shows the set up of apparatus that can be used to prepare Nitrogen (I) oxide in a school laboratory.



Preparation of Nitrogen(I)Oxide from ammonium nitrate(V)

c) Properties of nitrogen (I) oxide (Questions)

1. Write the equation for the reaction for the school laboratory preparation of Nitrogen (I) oxide.

Chemical equation $NH_4NO_2(s)$ -> H_2O (l) + N_2O (g)

2.a) State and explain three errors made in the above set up

-Oxygen is being generated instead of Nitrogen (I) oxide.

Ammonium Nitrate(V) should be used instead of potassium manganate(VI) and manganese(IV)oxide.

b) State three physical properties of Nitrogen (I) oxide.

-slightly soluble in water.
-colourless
-odourless
-less dense than air
-slightly sweet smell

3. State and explain the observation made when a burning magnesium ribbon is lowered in a gas jar containing Nitrogen (I) oxide.

Observation - Continues to burn with a bright flame
-White solid/residue is formed

Explanation-Magnesium burns in air to produce enough heat/energy split/break Nitrogen (I) oxide gas into free Nitrogen and oxygen then continues to burn in oxygen to form white solid/ash of Magnesium oxide.

Chemical equation

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$Mg(s) + N_2O(g) \rightarrow MgO(s) + N_2(g)$

4. State and explain the observation made when the following non metals are burnt then lowered in a gas jar containing Nitrogen (I) oxide.

a) Carbon/charcoal

Observation - Continues to burn with an orange glow

-colorless gas is formed that forms white precipitate with lime water.

Explanation-Carbon/charcoal burns in air to produce enough heat/energy split/break Nitrogen (I) oxide gas into free Nitrogen and oxygen then continues to burn in oxygen to form carbon (IV) oxide gas. Carbon (IV) oxide gas reacts to form a white precipitate with lime water.

Chemical equation $C(s) + 2N_2O(g) \rightarrow CO_2(g) + 2N_2(g)$

b) Sulphur powder

Observation - Continues to burn with a blue flame

-colorless gas is formed that turn orange acidified potassium dichromate (VI) to green.

Explanation-Sulphur burns in air to produce enough heat/energy split/break Nitrogen (I) oxide gas into free Nitrogen and oxygen then continues to burn in oxygen to form sulphur (IV) oxide gas. Sulphur (IV) oxide gas turns orange acidified potassium dichromate (VI) to green.

Chemical equation $S(s) + 2N_2O(g) \rightarrow SO_2(g) + 2N_2(g)$

5. State two uses of nitrogen (I) oxide

-As laughing gas because as anesthesia the patient regain consciousness laughing hysterically after surgery.

-improves engine efficiency.

6. State three differences between nitrogen (I) oxide and oxygen

-Oxygen is odourless while nitrogen (I) oxide has faint sweet smell

-Both relight/rekindle a glowing wooden splint but Oxygen can relight a feeble glowing splint while nitrogen (I) oxide relights well lit splint.

-Both are slightly soluble in water but nitrogen (I) oxide is more soluble.

ii) Nitrogen (II) oxide (NO)

a) Occurrence

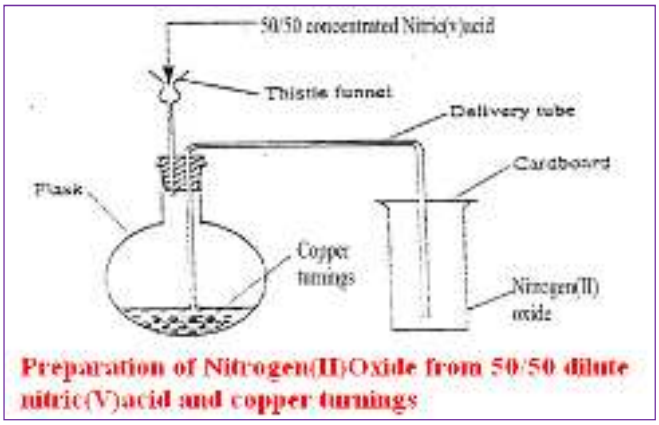
Nitrogen (II) oxide does not occur naturally but prepared in a laboratory.

b)Preparation

The set up below shows the set up of apparatus that can be used to prepare Nitrogen (II) oxide in a school laboratory.

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Preparation of Nitrogen(II)Oxide from 50/50 dilute nitric(V)acid and copper turnings

c) Properties of nitrogen (II) oxide (Questions)

1. **Write the equation for the reaction for the school laboratory preparation of Nitrogen (II) oxide.**

Chemical equation $3Cu(s) + 8HNO_3(aq) \rightarrow 4H_2O(l) + 2NO(g) + 2Cu(NO_3)_2(aq)$
Chemical equation $3Zn(s) + 8HNO_3(aq) \rightarrow 4H_2O(l) + 2NO(g) + 2Zn(NO_3)_2(aq)$
Chemical equation $3Mg(s) + 8HNO_3(aq) \rightarrow 4H_2O(l) + 2NO(g) + 2Mg(NO_3)_2(aq)$

2. State three physical properties of Nitrogen (II) oxide.

-insoluble in water.
-colourless
-odourless
-denser dense than air
-has no effect on both blue and red litmus papers

2. **State and explain the observation made when a burning magnesium ribbon is lowered in a gas jar containing Nitrogen (II) oxide.**

Observation - Continues to burn with a bright flame
-White solid/residue is formed
Explanation-Magnesium burns in air to produce enough heat/energy split/break Nitrogen (II) oxide gas into free Nitrogen and oxygen then continues to burn in oxygen to form white solid/ash of Magnesium oxide.
Chemical equation $2Mg(s) + 2NO(g) \rightarrow 2MgO(s) + N_2(g)$

3. **State and explain the observation made when the following non metals are burnt then lowered in a gas jar containing Nitrogen (II) oxide.**
   a) Carbon/charcoal

Observation - Continues to burn with an orange glow
-colorless gas is formed that forms white precipitate with lime water.

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Explanation-Carbon/charcoal burns in air to produce enough heat/energy split/break Nitrogen (II) oxide gas into free Nitrogen and oxygen then continues to burn in oxygen to form carbon (IV) oxide gas.Carbon (IV) oxide gas reacts to form a white precipitate with lime water.

Chemical equation $C(s) + 2NO(g) \rightarrow CO_2(g) + N_2(g)$

b) Sulphur powder

Observation - Continues to burn with a blue flame

-colorless gas is formed that turn orange acidified potassium dichromate (VI) to green.

Explanation-Sulphur burns in air to produce enough heat/energy split/break Nitrogen (II) oxide gas into free Nitrogen and oxygen then continues to burn in oxygen to form sulphur (IV) oxide gas.Sulphur (IV) oxide gas turns orange acidified potassium dichromate (VI) to green.

Chemical equation $S(s) + N_2O(g) \rightarrow SO_2(g) + N_2(g)$

c) Phosphorus

Observation - Continues to produce dense white fumes

Explanation-Phosphorus burns in air to produce enough heat/energy split/break Nitrogen (II) oxide gas into free Nitrogen and oxygen then continues to burn in oxygen to form dense white fumes of phosphorus (V) oxide gas.

Chemical equation $4P(s) + 10NO(g) \rightarrow 2P_2O_5(g) + 5N_2(g)$

5. State one use of nitrogen (II) oxide

As an intermediate gas in the Ostwalds process for manufacture of nitric(V) gas.

6. State and explain the observation made when nitrogen (II) oxide is exposed to the atmosphere.

Observation–brown fumes produced/evolved that turn blue litmus paper red.

Explanation- Nitrogen (II) oxide gas on exposure to air is quickly oxidized by the air/ oxygen to brown nitrogen (IV) oxide gas. Nitrogen (IV) oxide gas is an acidic gas.

Chemical equation $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

(colorless) (brown)

ii) Nitrogen (IV) oxide (NO_2)

a) Occurrence

Nitrogen (IV) oxide occurs -naturally from active volcanic areas.

-formed from incomplete combustion of the internal combustion engine of motor vehicle exhaust fumes.

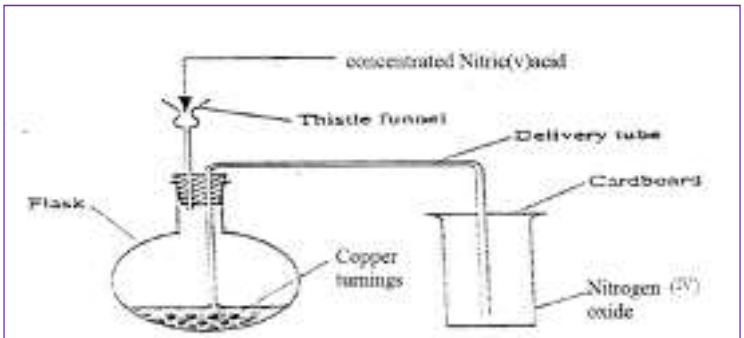
-from lightening

b)Preparation

The set up below shows the set up of apparatus that can be used to prepare Nitrogen (IV) oxide in a school laboratory.

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Preparation of Nitrogen(IV)oxide from Conc. Nitric(V)acid and Copper turnings

c) Properties of nitrogen (IV)oxide (Questions)

1. Write the equation for the reaction for the school laboratory preparation of Nitrogen (II) oxide.

Chemical equation $Cu(s) + 4HNO_3(aq) \rightarrow 2H_2O(l) + 2NO_2(g) + Cu(NO_3)_2(aq)$

Chemical equation $Zn(s) + 4HNO_3(aq) \rightarrow 2H_2O(l) + 2NO_2(g) + Zn(NO_3)_2(aq)$

Chemical equation $Fe(s) + 4HNO_3(aq) \rightarrow 2H_2O(l) + 2NO_2(g) + Fe(NO_3)_2(aq)$

2. State three physical properties of Nitrogen (IV) oxide.

-soluble/dissolves in water.
-brown in colour
-has pungent irritating poisonous odour/smell
-denser dense than air
-turns blue litmus papers to red

3. State and explain the observation made when Nitrogen (IV) oxidegas is bubbled in water.

Observation–The gas dissolves and thus brown colour of the gas fades
-A colourless solution is formed
-solution formed turns blue litmus papers to red
-solution formed has no effect on red

Explanation-Magnesium burns in air to produce enough heat/energy split/break Nitrogen (IV) oxide gas dissolves then react with water to form an acidic mixture of nitric(V) acid andnitric(III) acid.

Chemical equation $H_2O(l) + 2NO_2(g) \rightarrow HNO_3(aq) + HNO_2(aq)$
(nitric(V) acid) (nitric(III) acid)

4. State and explain the observation made when a test tube containing Nitrogen (IV) oxide is cooled then heated gently then strongly.

Observation on cooling
-Brown colour fades
-Yellow liquid formed

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Observation on gentle heating
-Brown colour reappears
-Yellow liquid formed changes to brown fumes/gas
Observation on gentle heating
-Brown colour fades
-brown fumes/gas changes to a colourless gas
Explanation-Brown nitrogen (IV) oxide gas easily liquefies to yellow dinitrogen tetraoxide liquid.When the yellow dinitrogen tetraoxide liquid is gently heated it changes back to the brown nitrogen (IV) oxidegas.When the brown nitrogen (IV) oxide gas is strongly heated it decomposes to colourless mixture of Nitrogen (II) oxide gas and Oxygen.
Chemical equation $O_2(s) + 2NO(g) \text{ ===== } 2NO_2(g) \text{ ===== } N_2O_4(l)$
(colourless gases) (brown gas) (yellow liquid)

5. State and explain the observation made when a burning magnesium ribbon is lowered in a gas jar containing Nitrogen (IV) oxide.
Observation - Continues to burn with a bright flame
-White solid/residue is formed
-Brown fumes/colour fades
Explanation-Magnesium burns in air to produce enough heat/energy split/break brown Nitrogen (IV) oxide gas into free colourless Nitrogen and oxygen then continues to burn in oxygen to form white solid/ash of Magnesium oxide.
Chemical equation $4Mg(s) + 2NO_2(g) \rightarrow 4MgO(s) + N_2(g)$

4. State and explain the observation made when the following non metals are burnt then lowered in a gas jar containing Nitrogen (IV) oxide.

a) Carbon/charcoal

Observation - Continues to burn with an orange glow
-Brown fumes/colour fades
-colorless gas is formed that forms white precipitate with lime water.
Explanation-Carbon/charcoal burns in air to produce enough heat/energy split/break brown Nitrogen (IV) oxide gas into free colourless Nitrogen and oxygen then continues to burn in oxygen to form carbon (IV) oxide gas.Carbon (IV) oxide gas reacts to form a white precipitate with lime water.
Chemical equation $2C(s) + 2NO_2(g) \rightarrow 2CO_2(g) + N_2(g)$

b) sulphur powder

Observation - Continues to burn with a blue flame
-Brown fumes/colour fades
-colorless gas is formed that turn orange acidified potassium dichromate (VI) to green.

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Explanation-Sulphur burns in air to produce enough heat/energy split/break brown Nitrogen (IV) oxide gas into free colourless Nitrogen and oxygen then continues to burn in oxygen to form sulphur (IV) oxide gas.Sulphur (IV) oxide gas turns orange acidified potassium dichromate (VI) to green.

Chemical equation $2S(s) + 2NO_2(g) \rightarrow 2SO_2(g) + N_2(g)$

c) Phosphorus

Observation- Continues to produce dense white fumes
-Brown fumes/colour fades

Explanation-Phosphorus burns in air to produce enough heat/energy split/break brown Nitrogen (IV) oxide gas into free colourless Nitrogen and oxygen then continues to burn in oxygen to form dense white fumes of phosphorus (V) oxide gas.

Chemical equation $8P(s) + 10NO_2(g) \rightarrow 4P_2O_5(g) + 5N_2(g)$

5. State two uses of nitrogen (IV) oxide

-In theOstwald process for industrial manufacture of nitric (V) gas.
-In the manufacture of T.N.T explosives

6. State and explain the observation made when nitrogen (II) oxide is exposed to the atmosphere.

Observation–brown fumes produced/evolved that turn blue litmus paper red.

Explanation- Nitrogen (II) oxide gas on exposure to air is quickly oxidized by the air/ oxygen to brown nitrogen (IV) oxide gas. Nitrogen (IV) oxide gas is an acidic gas.

Chemical equation $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$
(colourless) (brown)

C. AMMONIA (NH_3)

Ammonia is a compound of nitrogen and hydrogen only. It is therefore a hydride of nitrogen.

a) Occurrence

Ammonia gas occurs -naturally from urine of mammals and excretion of birds
-formed in the kidney of human beings

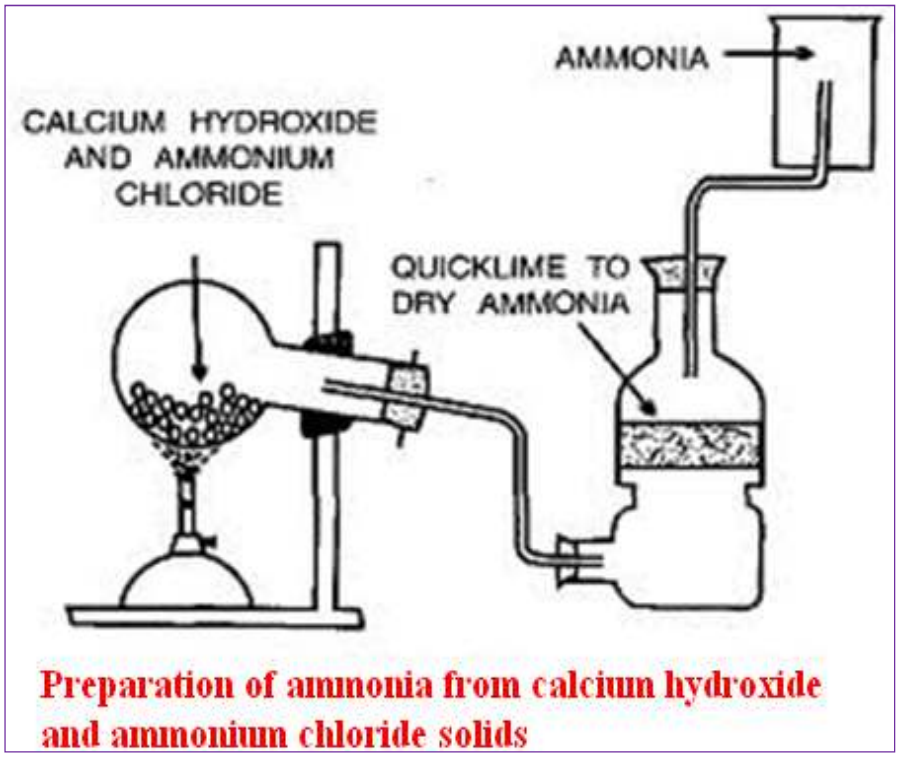
b)Preparation

The set up below shows the set up of apparatus that can be used to prepare dry Ammonia gas in a school laboratory.

Set up method 1

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Preparation of ammonia from calcium hydroxide and ammonium chloride solids

1. Write the equation for the reaction taking place in:

a) Method 1

Chemical equation

$Ca(OH)_2(s) + NH_4Cl(s) \rightarrow CaCl_2(aq) + H_2O(l) + 2NH_3(g)$

b)Method 2

Chemical equation

$NaOH(aq) + NH_4Cl(aq) \rightarrow NaCl(aq) + H_2O(l) + NH_3(g)$

2. State three physical properties of ammonia.

-has a pungent choking smell of urine
-Colourless
-Less dense than air hence collected by upward delivery
-Turns blue litmus paper blue thus is the only naturally occurring basic gas (at this level)

3. Calcium oxide is used as the drying agent. Explain why calcium chloride and concentrated sulphuric(VI) acid cannot be used to dry the gas.

-Calcium chloride reacts with ammonia forming the complex compound $CaCl_2.8H_2O$.

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Chemical equation $CaCl_2$ (s) + $8NH_3$(g) -> $CaCl_2$.$8NH_3$(g)

-Concentrated sulphuric(VI) acid reacts with ammonia forming ammonium sulphate(VI) salt compound

Chemical equation $2NH_3$(g) +H_2SO_4(l) ->$(NH_4)_2SO_4$(aq)

4. Describe the test for the presence of ammonia gas.

Using litmus paper:

Dip moist/damp/wet blue and red litmus papers in a gas jar containing a gas suspected to be ammonia.The blue litmus paper remain blue and the red litmus paper turns blue.Ammonia is the only basic gas.(At this level)

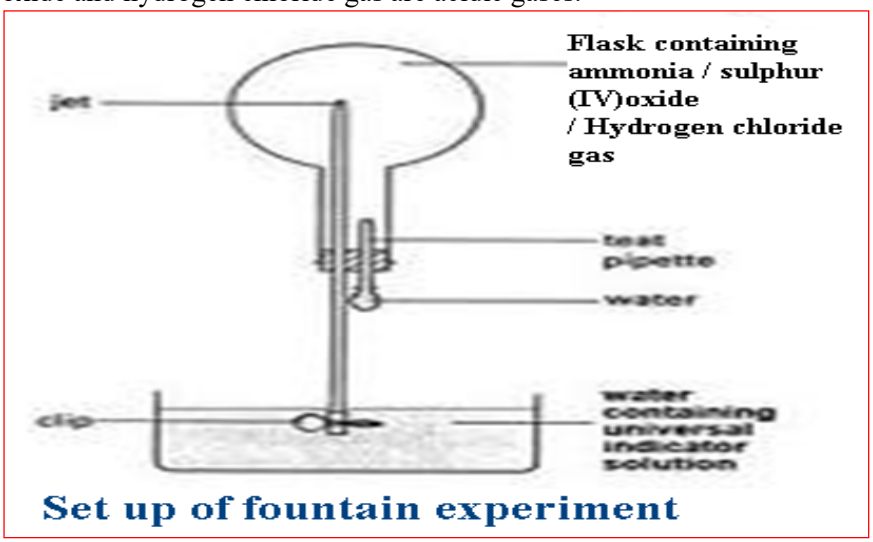
Using hydrogen chloride gas

Dip a glass rod in concentrated hydrochloric acid. Bring the glass rod near the mouth of a gas jar suspected to be ammonia. White fumes (of ammonium chloride)are produced/evolved.

5. Describe the fountain experiment to show the solubility of ammonia.

Ammonia is very soluble in water.

When a drop of water is introduced into flask containing ammonia, it dissolves all the ammonia in the flask. If water is subsequently allowed into the flask through a small inlet, atmospheric pressure forces it very fast to occupy the vacuum forming a fountain. If the water contains three/few drops of litmus solution, the litmus solution turns blue because ammonia is an alkaline/basic gas. If the water contains three/few drops of phenolphthalein indicator, the indicator turns pink because ammonia is an alkaline/basic gas. Sulphur(IV)oxide and hydrogen chloride gas are also capable of the fountain experiment . If the water contains three/few drops of phenolphthalein indicator, the indicator turns colourless because both Sulphur(IV) oxide and hydrogen chloride gas are acidic gases.



Set up of fountain experiment

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6. State and explain the observation made when hot platinum /nichrome wire is placed over concentrated ammonia solution with Oxygen gas bubbled into the mixture.

Observations

Hot platinum /nichrome wire continues to glow red hot.

Brown fumes of a gas are produced.

Explanation

Ammonia reacts with Oxygen on the surface of the wire .This reaction is exothermic producing a lot of heat/energy that enables platinum wire to glow red hot. Ammonia is oxidized to Nitrogen(II)oxide gas and water. Hot platinum /nichrome wire acts as catalyst to speed up the reaction. Nitrogen(II)oxide gas is further oxidized to brown Nitrogen(IV)oxide gas on exposure to air.

Chemical equation

(i)$4NH_3(g) + 5O_2(g) \xrightarrow{-Pt-} 4NO(g) + 6H_2O(l)$

(ii)$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

7. Ammonia gas was ignited in air enriched with Oxygen gas. State and explain the observations made

Observations

- Ammonia gas burns with a green flame
- Colourless gas produced

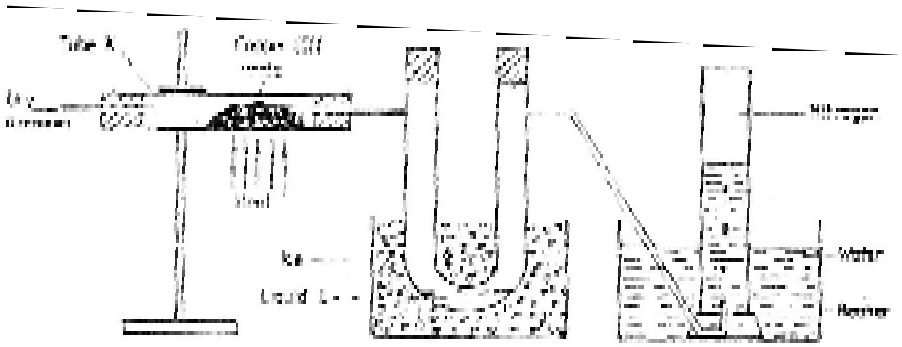
Explanation

Ammonia gas burns with a green flame in air enriched with Oxygen to from Nitrogen gas and water.

Chemical equation

$2NH_3(g) + O_2(g) \rightarrow N_2(g) + 3H_2O(l)$

8. Dry ammonia was passed through heated copper(II)Oxide as in the set up below.



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(a)State the observations made in tube K

-Colour changes from black to brown

-Colourless liquid droplet form on the cooler parts of tube K

(b)(i)Identify liquid L.

-Water/ $H_2O(l)$

(ii)Explain a chemical and physical test that can be used to identify liquid L.

Chemical test

(i) Add three/few drops of liquid L into anhydrous copper(II)sulphate(VI).
Colour changes from white to blue.
Explanation-Water changes white anhydrous copper(II)sulphate(VI) to blue hydrated copper(II)sulphate(VI)

(ii) Add three/few drops of liquid L into anhydrous cobalt(II)Chloride.
Colour changes from blue to pink.
Explanation-Water changes blue anhydrous cobalt(II)Chloride to pink hydrated cobalt(II)Chloride.

Physical test

(i)Heat the liquid. It boils at 100ºC at sea level (1atmosphere pressure/760mmHg pressure, 101300Pa,101300Nm^{-2}).

(ii)Cool the liquid. It freezes at 0.0ºC .

(iii)Determine the density. It is 1.0gcm^{-3}

(c)Write the equation for the reaction that take place.

$2NH_3(g) + 3CuO(s) \rightarrow N_2(g) + 3H_2O(l) + 3Cu(s)$
(black) (brown)

$2NH_3(g) + 3PbO(s) \rightarrow N_2(g) + 3H_2O(l) + 3Pb(s)$
(brown when hot) (grey)

8.(a)What is aqueous ammonia

Aqueous ammonia is formed when ammonia gas is dissolved in water.

$NH_3(g) + (aq) \rightarrow NH_3(aq)$

A little $NH_3(aq)$ reacts with ammonia water to form ammonia solution(NH_4OH)

$NH_3(aq) + H_2O(l) \rightleftharpoons OH^-(aq) + NH_4^+(aq)$

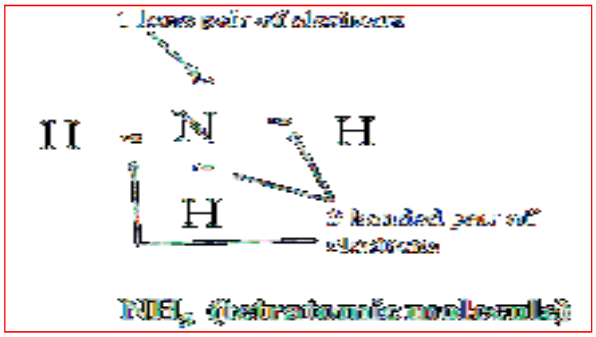
This makes a solution of aqueous ammonia is a weak base /alkali unlike other two alkalis.

9.Using dot and cross to represent outer electrons show the bonding in:

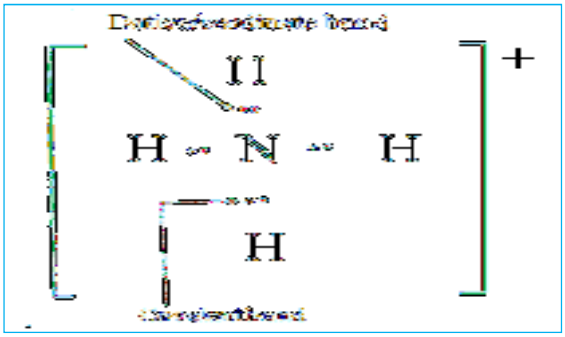
(a) NH_3

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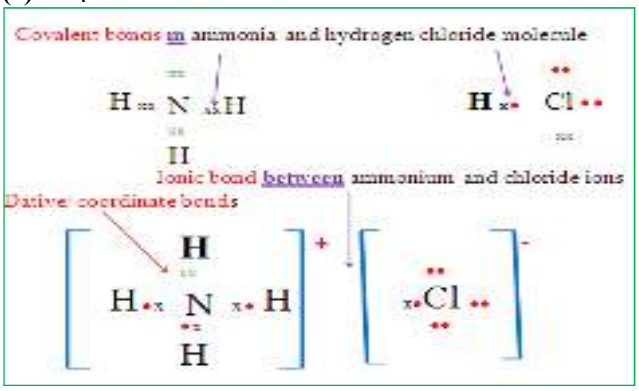
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(b) NH_4^+



(c)NH_4Cl



10.Name four uses of ammonia

(i)In the manufacture of nitrogenous fertilizers.
(ii) In the manufacture of nitric(V)acid from Ostwalds process.
(iii)As a refrigerant in ships and warehouses.
(iv)In softening hard water.
(v)In the solvay process for the manufacture of sodium carbonate.
(vi)In the removal of grease and stains.

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11.(a)Calculate the percentage of Nitrogen in the following fertilizers:

(i) $(NH_4)_2SO_4$

Molar mass of $(NH_4)_2SO_4$ = 132g

Mass of N in $(NH_4)_2SO_4$= 28g

% of N => $\frac{28}{132} \times 100$ = **21.2121%**

(ii) $(NH_4)_3PO_4$

Molar mass of $(NH_4)_3PO_4$ = 149g

Mass of N in $(NH_4)_3PO_4$= 42g

% of N => $\frac{42}{149} \times 100$ = **28.1879%**

(b)State two advantages of fertilizer a (i) over a (ii) above.

(i)Has higher % of Nitrogen

(ii)Has phosphorus which is necessary for plant growth.

(c) Calculate the mass of Nitrogen in a 50kg bag of:

(i) **$(NH_4)_2SO_4$**

% of N in $(NH_4)_2SO_4$ = **21.2121%**

Mass of N in 50 kg $(NH_4)_2SO_4$= $\frac{21.2121 \times 50}{100}$ = **10.6 kg**

(ii) **NH_4NO_3**

Molar mass of NH_4NO_3 = 80g

Mass of N in $(NH_4)_3PO_4$= 28g

% of N => $\frac{28}{80} \times 100$ = **35%**

% of N in **NH_4NO_3** = **35%**

Mass of N in 50 kg $(NH_4)_2SO_4$= $\frac{35 \times 50}{100}$ = **17.5 kg**

NH_4NO_3 therefore has a higher mass of Nitrogen than $(NH_4)_2SO_4$

d).Manufacture of Ammonia /Haber process

Most of the Ammonia produced for industrial purposes uses the Haber process developed by the German Scientist Fitz Haber.

(i)Raw materials

The raw materials include:

(i)Nitrogen from Fractional distillation of air from the atmosphere.

(ii)Hydrogen from:

I. Water gas-passing steam through heated charcoal

$C(s) + H_2O(l) \rightarrow CO(g) + H_2(g)$

II .Passing natural gas /methane through steam.

$CH_4(g) + H_2O(l) \rightarrow CO(g) + 3H_2(g)$

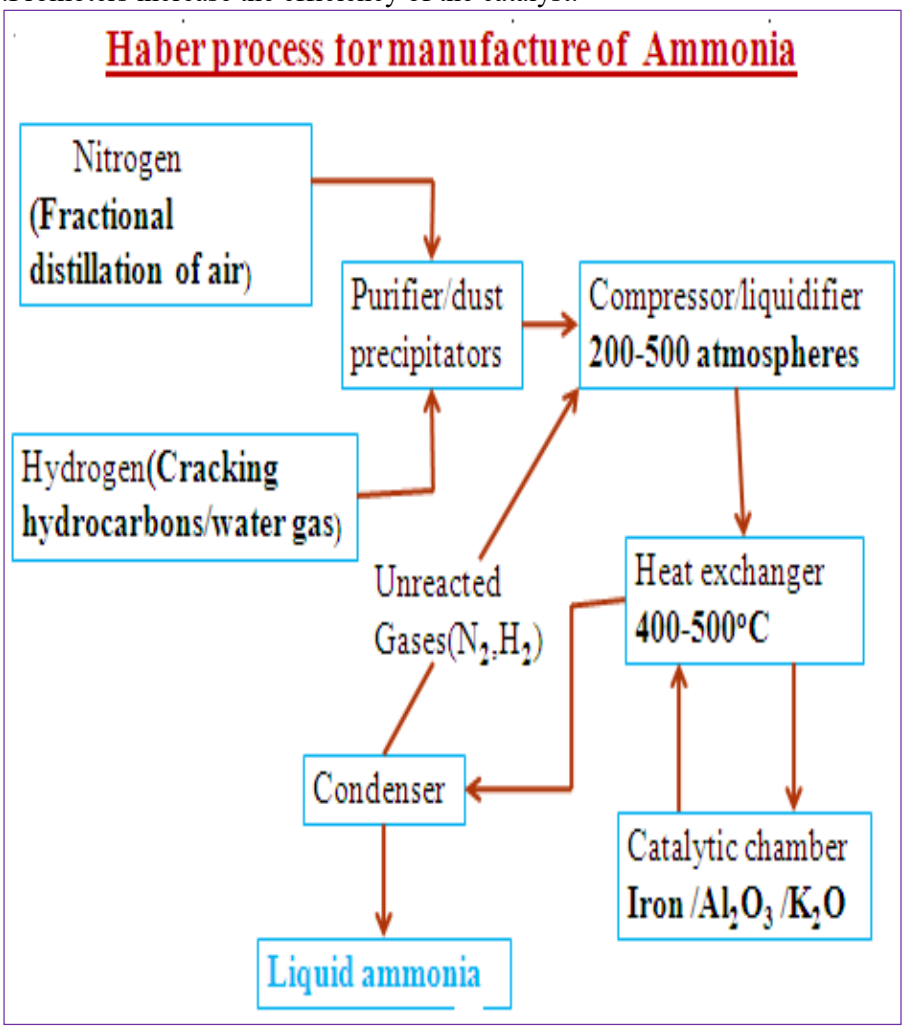
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(ii)Chemical process

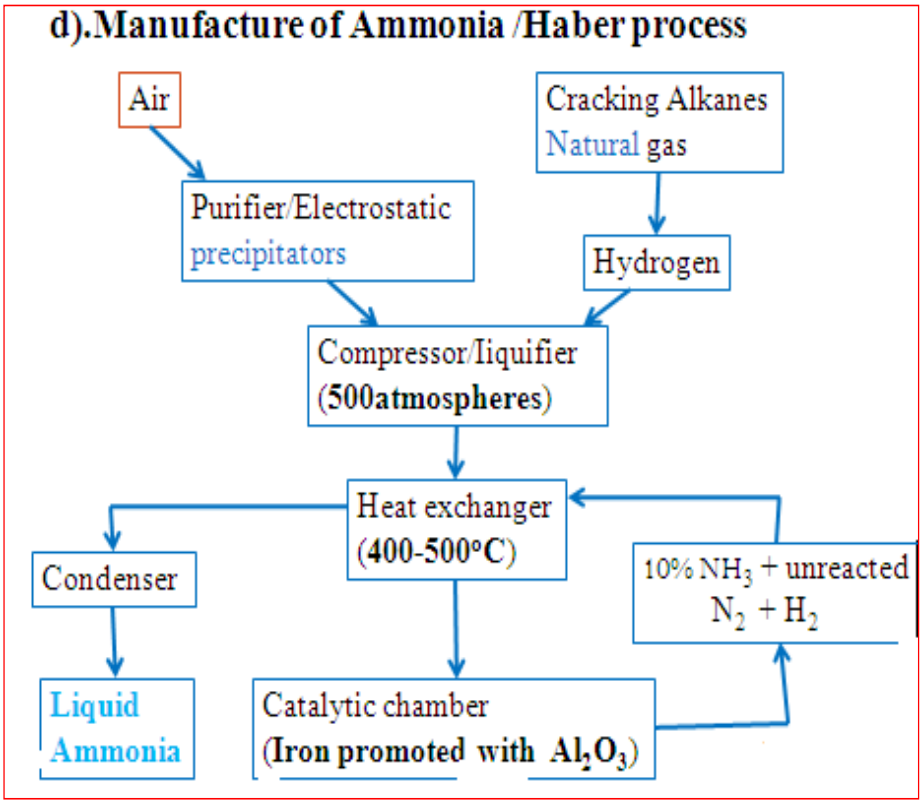
Hydrogen and Nitrogen are passed through a purifier to remove unwanted gases like Carbon(IV)oxide,Oxygen,sulphur(IV)oxide, dust, smoke which would poison the catalyst.

Hydrogen and Nitrogen are then mixed in the ratio of 3:1 respectively. The mixture is compressed to 200-250atmoshere pressure to liquidify. The liquid mixture is then heated to 400- 450°C.The hot compressed gases are then passed over finely divided Iron catalyst promoted/impregnated with Al_2O_3 /K_2O .Promoters increase the efficiency of the catalyst.



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Optimum conditions in Haber processs

Chemical equation

$N_2(g) + 3H_2(g) \xrightarrow{Fe/Pt} 2NH_3(g) \quad \Delta H = -92kJ$

Equilibrium/Reaction rate considerations

(i)**Removing ammonia** gas once formed shift the equilibrium forward to the right to replace the ammonia. More/higher yield of ammonia is attained.

(ii)**Increase in pressure** shift the equilibrium forward to the right where there is less volume/molecules. **More/higher yield** of ammonia is attained. Very **high** pressures raise the **cost** of production because they are **expensive** to produce and maintain. An optimum pressure of about 200atmospheres is normally used.

(iii)**Increase in temperature** shift the equilibrium backward to the left because the reaction is exothermic ($\Delta H = -92kJ$) . Ammonia formed **decomposes** back to Nitrogen and Hydrogen to remove excess heat therefore a **less** yield of ammonia is attained. Very **low** temperature decreases the collision frequency of Nitrogen and Hydrogen and thus the rate of reaction **too slow** and **uneconomical**.

An optimum temperature of about 450°C is normally used.

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(iv)Iron and platinum can be used as catalyst. **Platinum** is a **better** catalyst but more **expensive** and easily **poisoned** by impurities than Iron. Iron is promoted /impregnated with AluminiumOxide(Al_2O_3) to increase its surface area/area of contact with reactants and thus efficiency. The catalyst does not increase the yield of ammonia but it speed up its rate of formation.

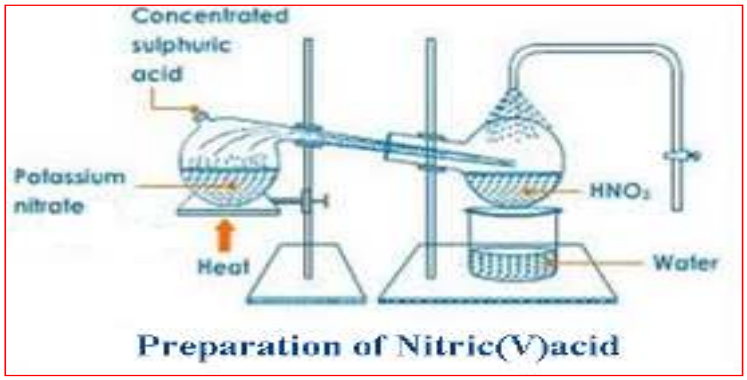
e) Nitric(V)acid (HNO_3)

a)Introduction.

Nitric(V)acid is one of the mineral acids .There are three mineral acids; Nitric(V)acid, sulphuric(VI)acid and hydrochloric acid. Mineral acids do not occur naturally but are prepared in a school laboratory and manufactured at industrial level.

b) School laboratory preparation

Nitric(V)acid is prepared in a school laboratory from the reaction of Concentrated sulphuric(VI)acid and potassium nitrate(V) below.



Preparation of Nitric(V)acid

(c)Properties of Concentrated Nitric (V)acid(Questions)

1.Write an equation for the school laboratory preparation of nitric(V)acid.

$KNO_3(s) + H_2SO_4(l) \rightarrow KHSO_4(s) + HNO_3(l)$

2.Sodium nitrate(V)can also be used to prepare nitric(V)acid. State two reasons why potassium nitrate(V) is preferred over Sodium nitrate(V).

(i) Potassium nitrate(V) is more volatile than sodium nitrate(V) and therefore readily displaced from the less volatile concentrated sulphuric(VI)acid

(ii) Sodium nitrate(V) is hygroscopic and thus absorb water . Concentrated sulphuric(VI)acid dissolves in water. The dissolution is a highly exothermic process.

3. An all glass apparatus /retort is used during the preparation of nitric(V) acid. Explain.

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Hot concentrated nitric(V) acid vapour is highly corrosive and attacks rubber cork apparatus if used.

4. Concentrated nitric(V) acid is colourless . Explain why the prepared sample in the school laboratory appears yellow.

Hot concentrated nitric(V) acid **decomposes** to brown nitrogen(IV)oxide and Oxygen gases.

$4HNO_3(l/g) \rightarrow 4NO_2(g) + H_2O(l) + O_2(g)$

Once formed the brown nitrogen(IV)oxide **dissolves** in the acid forming a yellow solution .

5. State and explain the observation made when concentrated nitric (V) acid is heated.

Observation

Brown fumes are produced.

Colourless gas that relights/rekindles glowing splint

Explanation

Hot concentrated nitric(V) acid **decomposes** to water, brown nitrogen(IV)oxide and Oxygen gases. Oxygen gas is not visible in the brown fumes of nitrogen (IV) oxide.

$4HNO_3(g) \rightarrow 4NO_2(g) + H_2O(l) + O_2(g)$

6. Explain the observations made when:

(a) About 2cm3 of Iron(II)sulphate(VI) solution is added about 5 drops of concentrated nitric(V) acid and the mixture then heated/warmed in a test tube.

Observation

(i)Colour changes from green to brown.

(ii)brown fumes /gas produced on the upper parts of the test tube.

Explanation

Concentrated nitric(V) acid is a powerful/strong **oxidizing** agent. It oxidizes green Fe^{2+} ions in $FeSO_4$ to brown/yellow Fe^{3+} .The acid is reduced to colourless Nitrogen(II)oxide.

Chemical equation:

$6FeSO_4(aq) + 3H_2SO_4(aq) + 2HNO_3(aq) \rightarrow 3Fe_2(SO_4)_3(aq) + 4H_2O + 2NO(g)$

Colourless Nitrogen(II)oxide is rapidly further oxidized to brown Nitrogen(IV)oxide by atmospheric oxygen.

Chemical equation:

$2NO(g) + O(g) \rightarrow 2NO_2(g)$

(colourless) (brown)

(b) A spatula full of sulphur powder in a clean dry beaker was added to 10cm3 concentrated nitric (V) acid and then heated gently/warmed.

Observation

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(i)Yellow colour of sulphur fades.
(ii) Brown fumes /gas produced.

Explanation

Concentrated nitric(V) acid is a powerful/strong **oxidizing** agent. It oxidizes yellow sulphur to colourless concentrated sulphuric(VI)acid. The acid is reduced to brown Nitrogen(IV)oxide gas.

Chemical equation:

$S(s) + 6HNO_3(l) \rightarrow 4NO_2(g) + H_2O(l) + H_2SO_4(l)$

(c) A few/about 1.0g pieces of copper turnings/Zinc granules/ Magnesium ribbon are added 10cm3 of concentrated nitric(V) acid in a beaker.

Observation

(i) Brown fumes /gas produced.
(ii) Blue solution formed with copper turnings
(iii) Colourless solution formed with Zinc granules/Magnesium ribbon

Explanation

Concentrated nitric (V) acid is a powerful/strong **oxidizing** agent. It oxidizes metals to their metal nitrate (VI) salts. The acid is reduced to brown Nitrogen (IV) oxide gas.

Chemical equation:

$Cu(s) + 4HNO_3(l) \rightarrow 2NO_2(g) + H_2O(l) + Cu(NO_3)_2(aq)$
$Zn(s) + 4HNO_3(l) \rightarrow 2NO_2(g) + H_2O(l) + Zn(NO_3)_2(aq)$
$Mg(s) + 4HNO_3(l) \rightarrow 2NO_2(g) + H_2O(l) + Mg(NO_3)_2(aq)$
$Pb(s) + 4HNO_3(l) \rightarrow 2NO_2(g) + H_2O(l) + Pb(NO_3)_2(aq)$
$Ag(s) + 2HNO_3(l) \rightarrow NO_2(g) + H_2O(l) + AgNO_3(aq)$

(d)Properties of Dilute Nitric (V)acid(Questions)

(i)What is dilute nitric(v)acid

When concentrated nitric(v)acid is added to over half portion of water ,it is relatively said to be dilute. A dilute solution is one which has more solvent/water than solute/acid. The number of moles of the acid are present in a large amount/volume of the solvent.This makes the molarity /number of moles present in one cubic decimeter of the solution to be low e.g. 0.02M.

If more water is added to the acid until the acid is too dilute to be diluted further then an **infinite** dilute solution if formed.

(ii))1cm length of polished Magnesium ribbon was put is a test tube containing 0.2M dilute nitric(v)acid. State and explain the observation made.

Observation

-Effervescence/bubbling/fizzing
-Colourless gas produced that extinguish burning splint with an explosion/pop sound
-Colourless solution formed
-Magnesium ribbon dissolves/decrease in size

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Explanation
Dilute dilute nitric(v)acid reacts with Magnesium to form hydrogen gas.
$Mg(s) + 2HNO_3(aq) \rightarrow H_2(g) + Mg(NO_3)_2(aq)$
With other reactive heavy metals, the hydrogen gas produced is rapidly oxidized to water.
Chemical equation $3Pb(s) + 8HNO_3(aq) \rightarrow 4H_2O(l) + 2NO(g) + 2Pb(NO_3)_2(aq)$
Chemical equation $3Zn(s) + 8HNO_3(aq) \rightarrow 4H_2O(l) + 2NO(g) + 2Zn(NO_3)_2(aq)$
Chemical equation $3Fe(s) + 8HNO_3(aq) \rightarrow 4H_2O(l) + 2NO(g) + 2Fe(NO_3)_2(aq)$
Hydrogen gas therefore is usually not prepared in a school laboratory using dilute nitric (v)acid.
(iii)A half spatula full of sodium hydrogen carbonate and Copper(II) carbonate were separately to separate test tubes containing 10cm3 of 0.2M dilute nitric (V) acid.
Observation
-Effervescence/bubbling/fizzing
-Colourless gas produced that forms a white precipitate with lime water.
-Colourless solution formed with sodium hydrogen carbonate.
- Blue solution formed with Copper(II) carbonate.
Explanation
Dilute dilute nitric (v)acid reacts with Carbonates and hydrogen carbonates to form Carbon(IV)oxide, water and nitrate(V)salt
$CuCO_3(s) + 2HNO_3(aq) \rightarrow H_2O(l) + Cu(NO_3)_2(aq) + CO_2(g)$
$ZnCO_3(s) + 2HNO_3(aq) \rightarrow H_2O(l) + Zn(NO_3)_2(aq) + CO_2(g)$
$CaCO_3(s) + 2HNO_3(aq) \rightarrow H_2O(l) + Ca(NO_3)_2(aq) + CO_2(g)$
$PbCO_3(s) + 2HNO_3(aq) \rightarrow H_2O(l) + Pb(NO_3)_2(aq) + CO_2(g)$
$FeCO_3(s) + 2HNO_3(aq) \rightarrow H_2O(l) + Fe(NO_3)_2(aq) + CO_2(g)$
$NaHCO_3(s) + HNO_3(aq) \rightarrow H_2O(l) + NaNO_3(aq) + CO_2(g)$
$KHCO_3(s) + HNO_3(aq) \rightarrow H_2O(l) + KNO_3(aq) + CO_2(g)$
$NH_4HCO_3(aq) + HNO_3(aq) \rightarrow H_2O(l) + NH_4NO_3(aq) + CO_2(g)$
$Ca(HCO_3)_2(aq) + 2HNO_3(aq) \rightarrow 2H_2O(l) + Ca(NO_3)_2(aq) + 2CO_2(g)$
$Mg(HCO_3)_2(aq) + 2HNO_3(aq) \rightarrow 2H_2O(l) + Mg(NO_3)_2(aq) + 2CO_2(g)$
(iii) 25.0cm3 of 0.1M Nitric(V) acid was titrated with excess 0.2M sodium hydroxide solution using phenolphthalein indicator.

I. State the colour change at the end point

Colourless

II. What was the pH of the solution at the end point. Explain.

pH 1/2/3

A little of the acid when added to the base changes the colour of the indicator to show the end point. The end point therefore is acidic with low pH of Nitric(V) acid. Nitric(V) acid is a strong acid with pH 1/2/3.

III. Calculate the number of moles of acid used.

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Number of moles = $\frac{\text{molarity x volume}}{1000}$ => $\frac{0.1 \times 25}{1000}$ = **2.5×10^{-3}**moles

IV. Calculate the volume of sodium hydroxide used

Volume of sodium hydroxide in cm3

= 1000 x Number of moles => $\frac{\textbf{1000x 2.5 x } 10^{-3}}{\text{Molarity}}$ = $\frac{\textbf{12.5cm3}}{\textbf{0.2}}$

(e)Industrial large scale manufacture of Nitric (V) acid

(i)Raw materials

1. Air/Oxygen

Oxygen is got from fractional distillation of air

Ammonia from Haber process.

2**. Chemical processes**

Air from the atmosphere is passes through electrostatic precipitators/filters to remove unwanted gases like Nitrogen, Carbon (IV) oxide, dust, smoke which may poison the catalyst. The ammonia -air mixture is compressed to 9 atmospheres to reduce the distance between reacting gases.

The mixture is passed through the heat exchangers where a temperature of 850ºC-900ºC is maintained.

The first reaction takes place in the catalytic chamber where Ammonia reacts with the air to form Nitrogen (**II**) Oxide and water.

Optimum condition in Ostwald's process

Chemical equation

$4NH_3$ (g) + $5O_2$ (g) === Pt/Rh === 4NO (g) + $6H_2O$ (g) ΔH = -950kJ

The reaction is reversible and exists in dynamic equilibrium where the products reform back the reactants. The following factors are used to increase the yield/amount of Nitrogen (II) oxide:

(i)**Removing Nitrogen (II) oxide** gas once formed shift the equilibrium forward to the right to replace the Nitrogen (II) oxide.

More/higher yield of Nitrogen (II) oxide is attained as reactants try to return the equilibrium balance.

(ii)**Increase in pressure** shift the equilibrium backward to the left where there are less volume/molecules.

Less/lower yield of Nitrogen (II) oxide is attained.

Very **low** pressures increases the **distance** between reacting NH_3 and O_2 molecules.

An **optimum** pressure of about **9 atmospheres** is normally used.

Cooling the mixture condenses the water vapour to liquid water

(iii)**Increase in temperature** shift the equilibrium backward to the left because the reaction is exothermic (ΔH = -950kJ).

Nitrogen (II) oxide and water vapour formed **decomposes** back to Ammonia and Oxygen to remove excess heat therefore a **less** yield of Nitrogen (II) oxide is attained.

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Very **low** temperature decreases the collision frequency of Ammonia and Oxygen and thus the rate of reaction **too slow** and **uneconomical**.
An optimum temperature of about 900ºC is normally used.
(iv)**Platinum** can be used as **catalyst**.
Platinum is very **expensive.** It is:
-promoted with Rhodium to increase the surface area/area of contact.
-added/coated on the surface of asbestos to form Platonized –asbestos to reduce the amount/quantity used.
The catalyst does not increase the yield of Nitrogen (II) Oxide but it speed up its rate of formation.
Nitrogen (II) oxide formed is passed through an oxidation reaction chamber where more air oxidizes the Nitrogen (II) Oxide to Nitrogen **(IV) Oxide** gas**.**
Chemical equation

$$2NO\ (g) + O_2\ (g) \rightarrow 2NO_2\ (g)$$

Nitrogen (IV) Oxide gas is passed up to meet a downward flow of water in the absorption chamber. The gas reacts with water to form a mixture of Nitric (V) and Nitric (III) acids
Chemical equation.

$$2NO_2\ (g) + H_2O\ (l) \rightarrow HNO_2\ (as) + HNO_3\ (as)$$

Excess air is bubbled through the mixture to oxidize Nitric (III)/ HNO_2 (as) to Nitric (V)/HNO_3 (as)
Chemical equation.

$$O_2\ (g) + 2HNO_2\ (as) \rightarrow 2HNO_3\ (as)$$

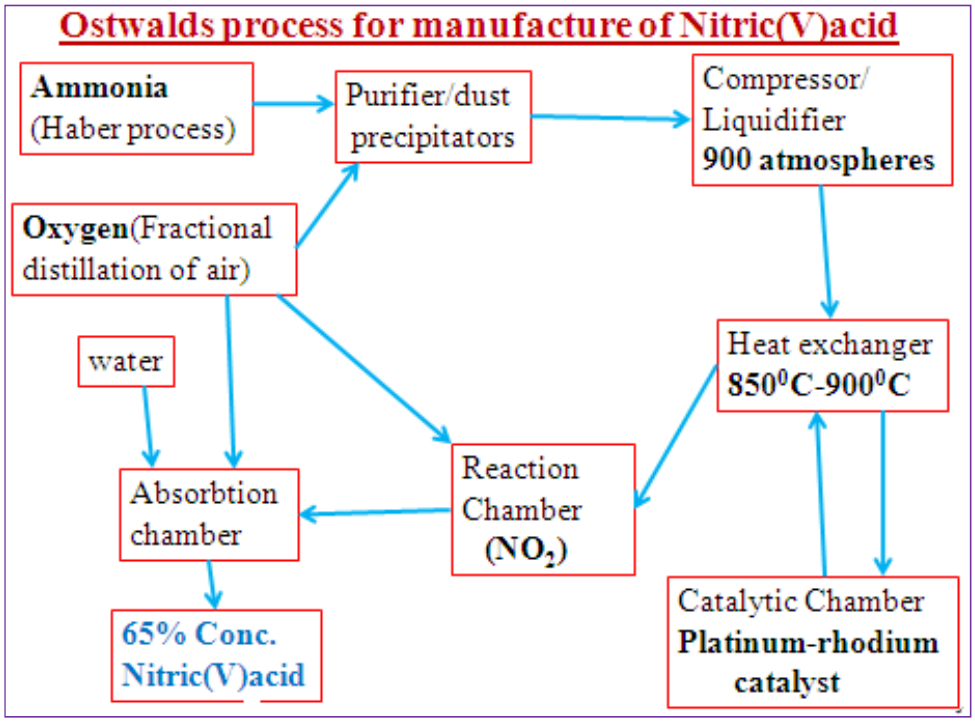
Overall chemical equation in the absorption chamber.

$$O_2\ (g) + 4NO_2\ (g) + 2H_2O\ (l) \rightarrow 4HNO_3\ (as)$$

The acid is 65% concentrated. It is made 100% concentrated by either:
(i) fractional distillation or
(ii) added to concentrated sulphuric (VI) acid to remove the 35% of water.

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A factory uses 63.0 kg of 68% pure nitric (V) acid per day to produce an ammonium fertilizer for an agricultural county. If the density of the acid is 1.42 gcm^{-3}, calculate:

(i) the concentration of the acid used in moles per litre.

Molar mass HNO_3 = 63

Method 1

Moles of HNO_3 in 1cm3 = Mass in 1cm3 / Molar mass HNO_3 = 1.42 / 63 = **0.0225 moles**

Molarity = Moles x 1000 / 1 cm3 => 0.0225 moles x10000 = **22.5molesdm^{-3}/M**

100% = 22.5molesdm^{-3}/M

68% = 68 x 22.5 / 100 = 15.3M/ **molesdm^{-3}**

Method 2

Moles of HNO_3 in 1000cm3 = Mass in 1000cm3 / Molar mass HNO_3 =>1.42 x1000 / 63

=**22.5397 molesdm^{-3}/M**

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100% = 22.5397 molesdm^{-3}/M

68% = $\frac{68 \times 22.5397}{100}$ = 15.327 **molesdm^{-3}**

(ii) the volume of ammonia gas at r.t.p used. (H=1.0, N=14.0, O=16.0, one mole of gas = 24 dm^{-3} at r.t.p)

Chemical equation

HNO_3 (as) + NH_3 (g) -> NH_4NO_3 (as)

Mole ratio HNO_3 (as): NH_3 (g) = 1: 1

1 mole HNO3 (as) -> 24dm3 NH_3 (g)

15.327 mole HNO_3 (as) ->$\frac{15.327 \text{ mole} \times 24 \text{ dm3}}{1 \text{dm3}}$ = **367.848dm3**

(iii) the number of crops which can be applied the fertilizer if each crop requires 4.0g.

HNO_3 (aq) + NH_3 (g) -> NH_4NO_3 (aq)

Molar mass NH_4NO_3 =80 g

Mole ratio HNO_3: NH_4NO_3 = **1: 1**

Mass of HNO_3 in 63.0 kg = 68% x 63 =**42.84kg**

1 mole HNO_3 (aq) =63g -> 80g NH_4NO_3

(42.84x1000) g HNO_3 (aq) -> $\frac{(42.84 \times 1000) \text{ g} \times 80}{63}$

= **54400g**

Mass of fertilizer = $\frac{54400\text{g}}{\text{Mass per crop } 4.0}$ = 13600 **crops**

E. NITRATE (V) NO_3^- and NITRATE (III) NO_2^- Salts

Nitrate (V) /NO_3^- and Nitrate (III) /NO_2^- are salts derived from Nitric (V)/HNO_3 and Nitric (III)/HNO_2 acids. Both HNO_3 and HNO_2 are monobasic acids with only **one** ignitable hydrogen in a molecule.

Only KNO_2, NaNO2 and NH_4NO_2 exist. All metallic nitrate (V) salts exist.

All Nitrate (V) /NO_3^- and Nitrate (III) /NO_2^- are soluble/dissolve in water.

(a)Effect of heat on Nitrate (V) /NO_3^- and Nitrate (III) /NO_2^- salts (Test for presence of Nitrate (V) /NO_3^- ions in solid state)

1. All Nitrate (III) /NO_2^- salts are not affected by gentle or strong heating **except** ammonium nitrate (III) NH_4NO_2.

Ammonium nitrate (III) NH_4NO_2 is a colourless solid that decompose to form Nitrogen gas and water.

Chemical equation

NH_4NO_2 (s) -> H_2O (l) + N_2 (g)

This reaction is used to prepare small amounts of Nitrogen in a school laboratory.

2. All Nitrate (V) /NO_3^- salts decompose on strong heating:

Experiment

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Put ½ spatula full of sodium nitrate (V) into a test tube. Place moist blue/red litmus papers on the mouth of the test tube. Heat strongly when test tube is slanted. Test the gases produced using glowing splint

Caution (i) Wear safety gas mask and hand gloves

(ii)Lead (II) nitrate (V) decomposes to Lead (II) oxide that reacts and fuses with the test tube permanently.

Repeat with potassium nitrate(V), copper(II) nitrate(V), Lead(II)nitrate(V), silver nitrate(V), Zinc nitrate(V), Magnesium nitrate(V) and Ammonium nitrate(V).

Observations

Cracking sound

Brown fumes/gas produced except in potassium nitrate (V) and Sodium nitrate (V)

Glowing splint relights/rekindles but feebly in Ammonium nitrate(V).

Black solid residue with copper(II) nitrate(V)

White residue/solid with sodium nitrate(V), potassium nitrate(V),silver nitrate(V), Magnesium nitrate(V)

Yellow residue/solid when hot but white on cooling with Zinc nitrate(V)

Brown residue/solid when hot but yellow on cooling with Lead(II)nitrate(V)

Explanation

1. Potassium nitrate(V) and Sodium nitrate(V) decomposes on strong heating to form potassium nitrate(III) and Sodium nitrate(III) producing Oxygen gas. Oxygen gas relights/rekindles a glowing splint.

Chemical equation.

$2KNO_3(s) \rightarrow 2KNO_2(s) + O_2(g)$

$2NaNO_3(s) \rightarrow 2NaNO_2(s) + O_2(g)$

2.Heavy metal nitrate(V)salts decomposes to form the oxide, brown nitrogen (IV) oxide and Oxygen gas.

Copper(II)oxide is black.Zinc oxide is yellow when hot and white when cool/cold.

Lead(II)oxide is yellow when cold/cool and brown when hot/heated.

Hydrated copper(II)nitrate is blue. On heating it melts and dissolves in its water of crystallization to form a green solution. When all the water of crystallization has evaporated,the nitrate(V)salt decomposes to black Copper(II)oxide and a mixture of brown nitrogen(IV)oxide gas and colourless Oxygen gas.

Chemical equation

$2Cu(NO_3)_2(s) \rightarrow 2CuO(s) + 4NO_2(g) + O_2(s)$

$2Ca(NO_3)_2(s) \rightarrow 2CaO(s) + 4NO_2(g) + O_2(s)$

$2Zn(NO_3)_2(s) \rightarrow 2ZnO(s) + 4NO_2(g) + O_2(s)$

$2Mg(NO_3)_2(s) \rightarrow 2MgO(s) + 4NO_2(g) + O_2(s)$

$2Pb(NO_3)_2(s) \rightarrow 2PbO(s) + 4NO_2(g) + O_2(s)$

$2Fe(NO_3)_2(s) \rightarrow 2FeO(s) + 4NO_2(g) + O_2(s)$

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Silver nitrate(V)and Mercury(II)nitrate decomposes to the corresponding metal and a mixture of brown nitrogen(IV)oxide gas and colourless Oxygen gas.
Chemical equation

$2AgNO_3(s) \rightarrow 2Ag(s) + 2NO_2(g) + O_2(s)$

$Hg(NO_3)_2(s) \rightarrow Hg(l) + 2NO_2(g) + O_2(s)$

The production/evolution of brown fumes of Nitrogen(IV)oxide gas on heating a salt is a confirmatory test for presence of NO_3^- ions of heavy metals

(b)Brown ring test (Test for presence of Nitrate(V) /NO_3^- ions in aqueous/ solution state)

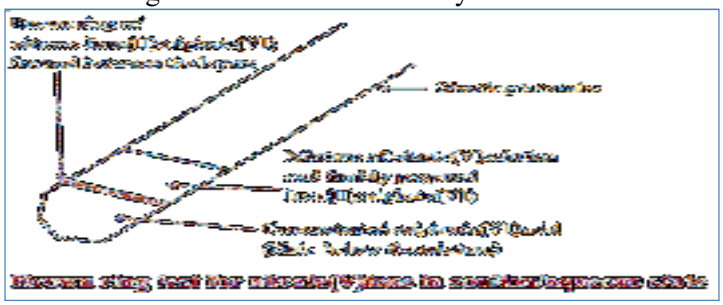
Experiment

Place 5cm3 of Potassium nitrate(V)solution onto a clean test tube. Add 8 drops of freshly prepared Iron(II)sulphate(VI)solution. Swirl/ shake.
Using a test tube holder to firmly slant and hold the test tube, carefully add 5cm3 of Concentrated sulphuric (VI) acid down along the side of test tube.Do not shake the test tube contents.

Caution: Concentrated sulphuric (VI) acid is highly corrosive.

Observation.

Both Potassium nitrate(V)and freshly prepared Iron(II)sulphate (VI)do not form layers
On adding Concentrated Sulphuric(VI)acid,two layers are formed.
A brown ring is formed between the layers.



Explanation

All nitrate(V)salts are soluble. They form a miscible mixture when added freshly prepared Iron(II)sulphate(VI)solution. Concentrated sulphuric(VI)acid is denser than the miscible mixture thus settle at the bottom.
At the junction of the layers, the acid reacts with nitrate(V)salts to form Nitric(V)acid/HNO_3. Nitric(V)acid/HNO_3 is reduced to Nitrogen (II)oxide by the Iron(II)sulphate(VI) salt to form the complex compound Nitroso-iron(II)sulphate(VI)/$FeSO_4.NO$. Nitroso-iron(II)sulphate(VI) is brown in colour.It forms a thin layer at the junction between concentrated sulphuric (VI)acid and the miscible mixture of freshly prepared Iron(II) sulphate(VI) and the nitrate(V)salts as a brown ring.

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Chemical equation

$FeSO_4(aq) + NO(g) \rightarrow$ **$FeSO_4.NO$**(aq)

(**Nitroso-iron(II)sulphate(VI)complex**)

The brown ring disappear if shaken because concentrated sulphuric (VI)acid mixes with the aqueous solution generating a lot of heat which decomposes Nitroso-iron(II)sulphate(VI)/$FeSO_4$.NO to iron(II)sulphate(VI) and Nitrogen(II)oxide.

Chemical equation

$FeSO_4.NO$ $(aq) \rightarrow FeSO_4(aq) + NO(g) \rightarrow$

Iron(**II**)sulphate(VI) solution is easily/readily oxidized to iron(**III**)sulphate(VI) on exposure to air/oxygen. The brown ring test thus require **freshly** prepared Iron(**II**) sulphate(VI) solution

(c)Devardas alloy test (Test for presence of Nitrate(V) /NO_3^- ions in aqueous/ solution state)

Experiment

Place 5cm3 of Potassium nitrate(V)solution onto a clean test tube. Add 5 drops of sodium hydroxide solution. Swirl/ shake. Add a piece of aluminium foil to the mixture.Heat.Test any gases produced using both blue and red litmus papers.

Observation.	Inference
Effervescence/bubbles/fizzing colourless gas that has a pungent smell of urine Blue limus paper remain blue Red litmus paper turn red.	NO_3^-

Explanation

The Devardas alloy test for NO_3^- ions in solution was developed by the Italian scientist Artulo Devarda(1859-1944)

When a NO_3^-salt is added sodium hydroxide and aluminium foil, effervescence of ammonia gas is a confirmatory test for NO_3^- ions.

CHEMISTRY OF SULPHUR

A.SULPHUR (S)

Sulphur is an element in Group VI Group 16)of the Periodic table . It has atomic number 16 and electronic configuration 16 and valency 2 /divalent and thus forms the ion S^{2-}

A. Occurrence.

Sulphur mainly occurs:

(i) as **free** element in Texas and Louisiana in USA and Sicily in Italy.

(ii)**Hydrogen sulphide** gas in active volcanic areas e.g. Olkaria near Naivasha in Kenya

(iii)as **copper pyrites(**$CuFeS_2$**)** ,**Galena** (PbS,**Zinc blende**(ZnS))and **iron pyrites**(FeS_2) in other parts of the world.

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B. Extraction of Sulphur from **Fraschs** process

Suphur occurs about 200 metres underground. The soil structure in these areas is usually **weak** and can easily **cave** in.

Digging of tunnels is thus discouraged in trying to extract the mineral.

Sulphur is extracted by drilling three concentric /round pipes of diameter of ratios 2:8: 18 centimeters.

Superheated water at 170ºC and 10atmosphere pressure is forced through the outermost pipe.

The high pressures ensure the water remains as liquid at high temperatures instead of vapour of vapour /gas.

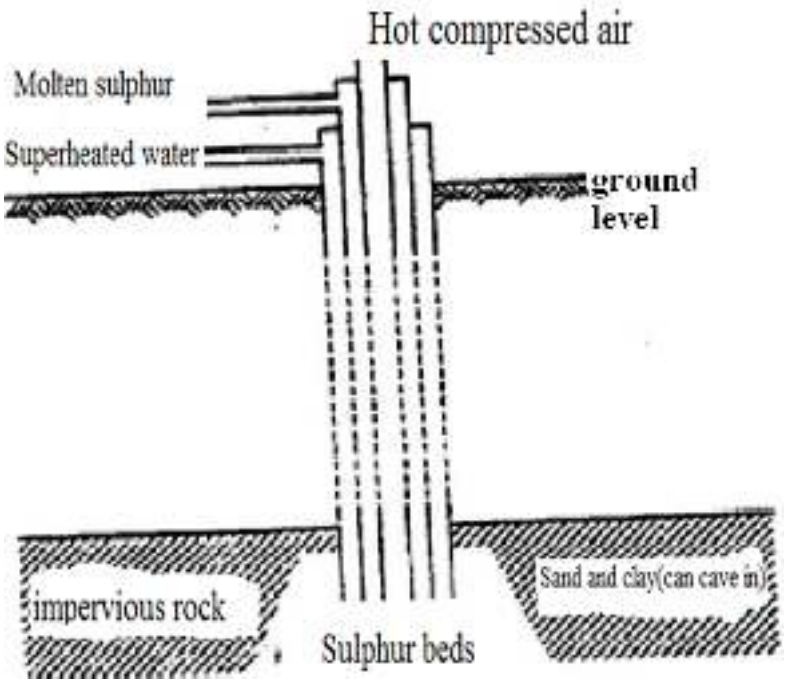
The superheated water melts the sulphur because the melting point of sulphur is lower at about at about 115ºC.

A compressed air at 15 atmospheres is forced /pumped through the innermost pipe.

The hot air forces the molten sulphur up the middle pipe where it is collected and solidifies in a large tank.

It is about 99% pure.

Diagram showing extraction of Sulphur from Fraschs Process



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C. Allotropes of Sulphur.

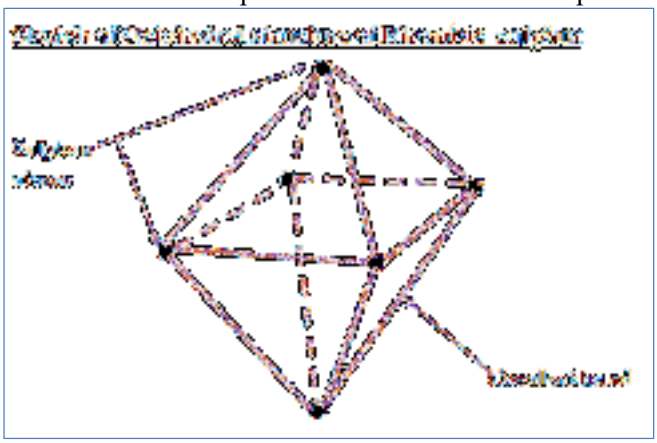
1. Sulphur exists as two crystalline allotropic forms:

(i)Rhombic sulphur

(ii)Monoclinic sulphur

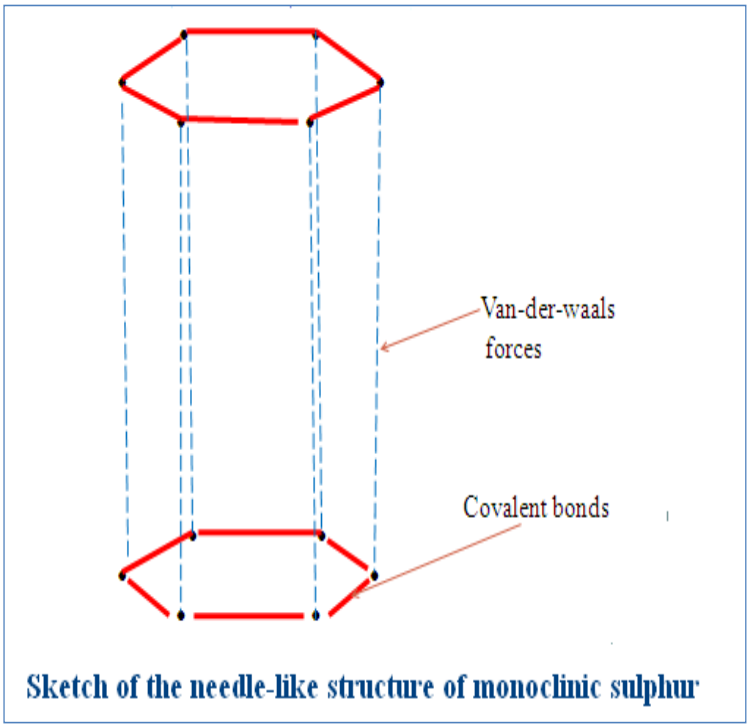
Rhombic sulphur	Monoclinic sulphur
Bright yellow crystalline solid	Pale yellow crystalline solid
Has a melting point of 113°C	Has a melting point of 119°C
Has a density of 2.06gcm^{-3}	Has a density of 1.96gcm^{-3}
Stable below 96°C	Stable above 96°C
Has octahedral structure	Has a needle-like structure

Rhombic sulphur and Monoclinic sulphur have a **transition** temperature of 96°C.This is the temperature at which one allotrope changes to the other.



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Sketch of the needle-like structure of monoclinic sulphur

2. Sulphur exists in non-crystalline forms as:

(i)Plastic sulphur-

Plastic sulphur is prepared from heating powdered sulphur to boil then pouring a thin continuous stream in a beaker with cold water. A long thin elastic yellow thread of plastic sulphur is formed .If left for long it turn to bright yellow crystalline rhombic sulphur.

(ii)Colloidal sulphur-

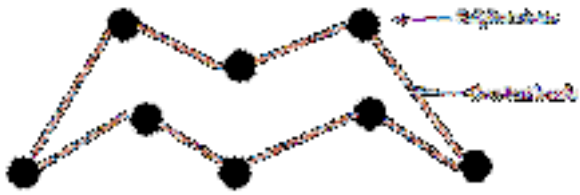
Colloidal sulphur is formed when sodium thiosulphate ($Na_2S_2O_3$) is added hydrochloric acid to form a yellow precipitate.

D. Heating Sulphur.

A molecule of sulphur exists as puckered ring of eight atoms joined by covalent bonds as S_8.

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On heating the yellow sulphur powder melts at 113°C to clear amber liquid with low viscosity and thus flows easily.
On further heating to 160°C the molten liquid darkens to a brown very viscous liquid that does not flow easily.
This is because the S_8 rings break into S_8 chain that join together to form very long chains made of over 100000 atoms of Sulphur.
The long chains **entangle** each other reducing their mobility /flow and hence increases their viscosity.
On continued further heating to above 160°C, the viscous liquid darkens but becomes more mobile/flows easily and thus less viscous.
This is because the long chains break to smaller/shorter chains.
At 444°C, the liquid boils and forms brown vapour of a mixture of S_8 ,S_6 ,S_2 molecules that solidifies to S_8 ring of "flowers of sulphur" on the cooler parts.

Summary of changes on heating sulphur

Observation on heating	Explanation/structure of Sulphur
Solid sulphur	Puckered S_8 ring
Heat to 113°C Amber yellow liquid	Puckered S_8 ring in liquid form (low viscosity/flow easily)
Heat to 160°C Liquid darkens	Puckered S_8 ring break/opens then join to form long chains that entangle (very high viscosity/very low rate of flow)
Heat to 444°C Liquid boils to brown vapour	Mixture of S_8 ,S_6 ,S_2 vapour
	Puckered S_8 ring

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Cool to room temperature Yellow sublimate (Flowers of Sulphur)	

E. Physical and Chemical properties of Sulphur.(**Questions**)

1. State three physical properties unique to Sulphur

Sulphur is a yellow solid, insoluble in water, soluble in carbon disulphide/tetrachloromethane/benzene, poor conductor of heat and electricity. It has a melting point of 115°C and a boiling point of 444°C.

2. Moist/damp/wet blue and red litmus papers were put in a gas jar containing air/oxygen. Burning sulphur was then lowered into the gas jar. State and explain the observation made.

Observations

-Sulphur melts then burns with a blue flame

Colourless gas produced that has a pungent smell

Red litmus paper remains red. Blue litmus paper turns red.

Explanation

Sulphur burns in air and faster in Oxygen to form Sulphur(IV)Oxide gas and traces/small amount of Sulphur(VI)Oxide gas. Both oxides react with water to form the corresponding acidic solution i.e

(i) Sulphur(IV)Oxide gas reacts with water to form sulphuric(IV)acid

(ii) Sulphur(VI)Oxide gas reacts with water to form sulphuric(VI)acid

Chemical equation

$S(s) + O_2(g) \rightarrow SO_2(g)$ (Sulphur(IV)Oxide gas)

$2S(s) + 3O_2(g) \rightarrow 2SO_3(g)$ (Sulphur(VI)Oxide gas traces)

$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$ (sulphuric(IV)acid)

$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$ (sulphuric(VI)acid).

3. Iron filings were put in a test tube containing powdered sulphur then heated on a Bunsen flame. Stop heating when reaction starts. State and explain the observations made. Test the effects of a magnet on the mixture before and after heating. Explain.

Observations

Before heating, the magnet attracts iron filings leaving sulphur

After heating, the magnet does not attract the mixture.

After heating, a red glow is observed that continues even when heating is stopped..

Black solid is formed.

Explanation Iron is attracted to a magnet because it is ferromagnetic.

When a mixture of iron and sulphur is heated, the reaction is exothermic giving out heat energy that makes the mixture to continue glowing even after stopping heating.

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Black Iron(II)sulphide is formed which is a compound and thus not ferromagnetic.
Chemical equation Fe(s) + S(s) -> FeS(s) (Exothermic reaction/ -ΔH)

Heated powdered heavy metals combine with sulphur to form **black** sulphides.

Cu(s) + S(s) -> CuS(s)
Zn(s) + S(s) -> ZnS(s)
Pb(s) + S(s) -> PbS(s)

4.The set up below show the reaction of sulphur on heated concentrated sulphuric(VI)acid.



(i)State and explain the observation made.

Observation
Yellow colour of sulphur fades
Orange colour of potassium dichromate(VI)paper turns to green.

Explanation

Hot concentrated sulphuric(VI)acid oxidizes sulphur to sulphur (IV)oxide gas. The oxide is also reduced to water. Traces of sulphur (VI)oxide is formed.
Chemical equation

$S(s) + 3H_2SO_4(l) \rightarrow 3SO_2(g) + 3H_2O(l) + SO_3(g)$

Sulphur (IV)oxide gas turns Orange potassium dichromate(VI)paper to green.

(ii)State and explain the observation made if concentrated sulphuric (VI) acid is replaced with concentrated Nitric (V) acid in the above set up.

Observation

Yellow colour of sulphur fades
Colurless solution formed
Brown fumes/gas produced.

Explanation

Hot concentrated Nitric(V)acid oxidizes sulphur to sulphuric (VI)acid. The Nitric (V) acid is reduced to brown nitrogen(IV)oxide gas.
Chemical equation

$S(s) + 6HNO_3(l) \rightarrow 6NO_2(g) + 2H_2O(l) + H_2SO_4(l)$

NB:
Hydrochloric acid is a weaker oxidizing agent and thus cannot oxidize sulphur like the other mineral acids.

5. State three main uses of sulphur .

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Sulphur is mainly used in:

(i)Contact process for the manufacture/industrial/large scale production of concentrated sulphuric(VI)acid.

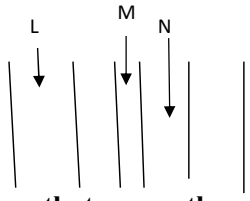
(ii)Vulcanization of rubber to make it harder, tougher, stronger, and more durable.

(iii)Making gun powder and match stick heads

(iv) As ointments to treat fungal infections

6. Revision Practice

The diagram below represents the extraction of sulphur by Fraschs process. Use it to answer the questions that follow.



(a)Name the substances that passes through:

M Superheated water at 170oC and 10 atmosphere pressure

L Hot compressed air

N Molten sulphur

(b)What is the purpose of the substance that passes through L and M?

M- Superheated water at 170°C and 10 atmosphere pressure is used to melt the sulphur

L- Hot compressed air is used to force up the molten sulphur.

(c) The properties of the two main allotropes of sulphur represented by letters A and B are given in the table below. Use it to answer the questions that follow.

	A	**B**
Appearance	**Bright yellow**	**Pale yellow**
Density(gcm^{-3})	**1.93**	**2.08**
Melting point(^{o}C)	**119**	**113**
Stability	**Above 96°C**	**Below 96°C**

I.What are allotropes?

Different forms of the same element existing at the same temperature and pressure without change of state.

II. Identify allotrope:

A. Monoclinic sulphur

B . Rhombic sulphur

III. State two main uses of sulphur.

-Manufacture of sulphuric(VI)acid

-as fungicide

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-in vulcanization of rubber to make it harder/tougher/ stronger
-manufacture of dyes /fibres

(d)Calculate the volume of sulphur (IV)oxide produced when 0.4 g of sulphur is completely burnt in excess air.(S = 32.0 ,I mole of a gas occupies 24 dm3 at room temperature)

Chemical equation

$S(s) + O_2(g) \rightarrow SO_2(g)$

Mole ratio S: SO_2 = 1:1

Method 1

32.0 g of sulphur -> 24 dm3 of $SO_2(g)$

0.4 g of sulphur -> $\frac{0.4 \text{ g} \times 24 \text{ dm3}}{32.0 \text{ g}}$ = **0.3 dm3**

Method 2

Moles of sulphur used = $\frac{\text{Mass of sulphur}}{\text{Molar mass of sulphur}}$ => $\frac{0.4}{32}$ = 0.0125 moles

Moles of sulphur used = Moles of sulphur(IV)oxide used=>0.0125 moles

Volume of sulphur(IV)oxide used = Moles of sulphur(IV)oxide x volume of one mole of gas =>0.0125 moles x 24 dm3 = **0.3 dm3**

B.COMPOUNDS OF SULPHUR

The following are the main compounds of sulphur:

(i) Sulphur(IV)oxide
(ii) Sulphur(VI)oxide
(iii) Sulphuric(VI)acid
(iv) Hydrogen Sulphide
(v) Sulphate(IV)/SO_3^{2-} and Sulphate(VI)/ SO_4^{2-} salts

(i) Sulphur(IV)oxide(SO_2)

(a) Occurrence

Sulphur (IV)oxide is found in volcanic areas as a gas or dissolved in water from geysersand hot springs in active volcanic areas of the world e.g. Olkaria and Hells gate near Naivasha in Kenya.

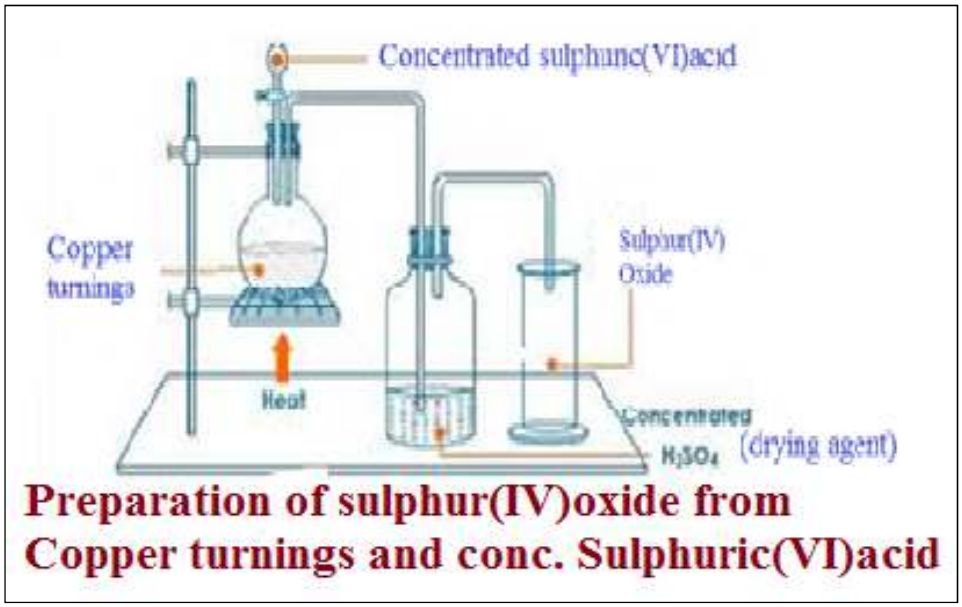
(b) School laboratory preparation

In a Chemistry school laboratory Sulphur (IV)oxide is prepared from the reaction of

Method 1:Using Copper and Sulphuric(VI)acid.

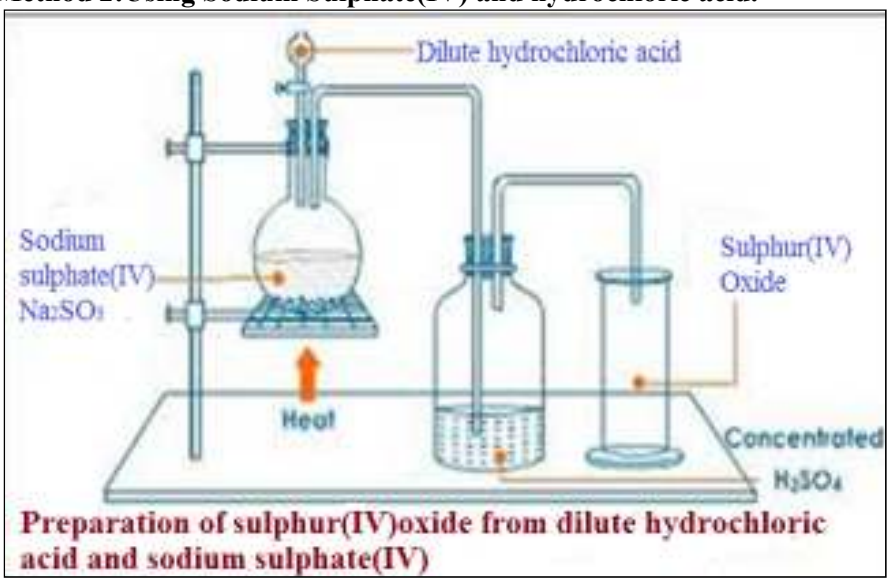
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Preparation of sulphur(IV)oxide from Copper turnings and conc. Sulphuric(VI)acid

Method 2:Using Sodium Sulphate(IV) and hydrochloric acid.



Preparation of sulphur(IV)oxide from dilute hydrochloric acid and sodium sulphate(IV)

(c)Properties of Sulphur(IV)oxide(Questions)

1. Write the equations for the reaction for the formation of sulphur (IV)oxide using:

(i)Method 1

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$Cu(s) + 2H_2SO_4(l) \rightarrow CuSO_4(aq) + SO_2(g) + 2H_2O(l)$
$Zn(s) + 2H_2SO_4(l) \rightarrow ZnSO_4(aq) + SO_2(g) + 2H_2O(l)$
$Mg(s) + 2H_2SO_4(l) \rightarrow MgSO_4(aq) + SO_2(g) + 2H_2O(l)$
$Fe(s) + 2H_2SO_4(l) \rightarrow FeSO_4(aq) + SO_2(g) + 2H_2O(l)$

Calcium ,Lead and Barium will form insoluble sulphate(VI)salts that will cover unreacted metals stopping further reaction thus producing very small amount/quantity of sulphur (IV)oxide gas.

(ii)Method 2

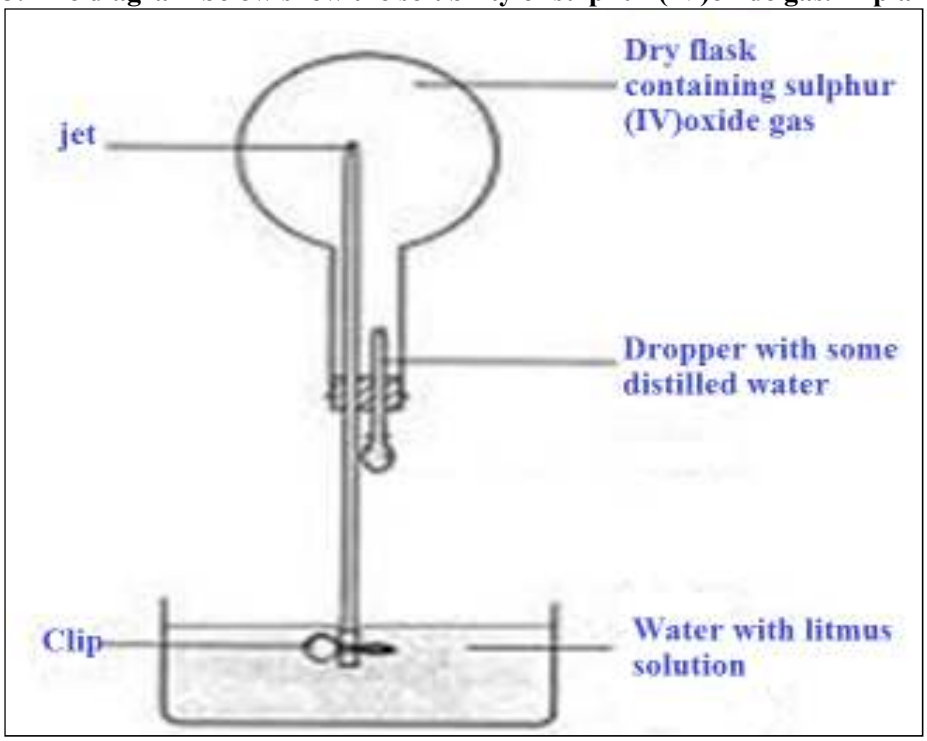
$Na_2SO_3(aq) + HCl(aq) \rightarrow NaCl(aq) + SO_2(g) + 2H_2O(l)$
$K_2SO_3(aq) + HCl(aq) \rightarrow KCl(aq) + SO_2(g) + 2H_2O(l)$
$BaSO_3(s) + 2HCl(aq) \rightarrow BaCl_2(aq) + SO_2(g) + H_2O(l)$
$CaSO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + SO_2(g) + H_2O(l)$
$PbSO_3(s) + 2HCl(aq) \rightarrow PbCl_2(\mathbf{s}) + SO_2(g) + H_2O(l)$

Lead(II)chloride is soluble on heating thus reactants should be heated to prevent it coating/covering unreacted $PbSO_3(s)$

2.State the physical properties unique to sulphur (IV)oxide gas.

Sulphur (IV)oxide gas is a colourless gas with a pungent irritating and choking smell which **liquidifies** easily. It is about two times denser than air.

3. The diagram below show the solubility of sulphur (IV)oxide gas. Explain.



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Sulphur(IV) oxide is very soluble in water.
One drop of water dissolves all the Sulphur (IV) oxide in the flask leaving a vacuum.
If the clip is removed, atmospheric pressure forces the water up through the narrow tube to form a fountain to occupy the vacuum.
An acidic solution of sulphuric (IV)acid is formed which turns litmus solution red.
Chemical equation
$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$ (sulphuric(IV)acid turn litmus red)

4.Dry litmus papers and wet/damp/moist litmus papers were put in a gas jar containing sulphur(IV) oxide gas. State and explain the observations made.

Observations
(i)Dry Blue litmus paper remains blue.
Dry red litmus paper remains red.
(ii) Wet/damp/moist blue litmus paper turns red.
Moist/damp/wet red litmus paper remains red.
Both litmus papers are then bleached /decolorized.
Explanation
Dry sulphur(IV) oxide gas is a molecular compound that does not dissociate/ionize to release H+(aq)ions and thus has no effect on dry blue/red litmus papers.
Wet/damp/moist litmus papers contain water that dissolves /react with dry sulphur(IV) oxide gas to form a solution of weak sulphuric(IV)acid ($H_2SO_3(aq)$).
Weak sulphuric(IV)acid($H_2SO_3(aq)$) dissociates /ionizes into free H+(aq)ions:

$$H_2SO_3(aq) \rightarrow 2H^+(aq) + SO_3^{2-}(aq)$$

The free $H^+(aq)$ions are responsible for turning blue litmus paper turns red showing the gas is acidic.
The $SO_3^{2-}(aq)$ ions in wet/damp/moist sulphur(IV) oxide gas is responsible for many reactions of the gas.
It is easily/readily oxidized to sulphate(VI) $SO_4^{2-}(aq)$ ions making sulphur(IV) oxide gas act as a reducing agent as in the following examples:

(a)Bleaching agent

Wet/damp/moist coloured flowers/litmus papers are bleached/decolorized when put in sulphur(IV) oxide gas.
This is because sulphur(IV) oxide removes atomic oxygen from the coloured dye/ material to form sulphuric(VI)acid.
Chemical equations
(i)Formation of sulphuric(IV)acid

$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$$

(ii)Decolorization/bleaching of the dye/removal of atomic oxygen.

Method I. $H_2SO_3(aq) + (dye + O) \rightarrow H_2SO_4(aq) + dye$
(coloured) (colourless)

Method II. $H_2SO_3(aq) + (dye) \rightarrow H_2SO_4(aq) + (dye - O)$

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(coloured) (colourless)

Sulphur(IV) oxide gas therefore bleaches by reduction /removing oxygen from a dye unlike chlorine that bleaches by oxidation /adding oxygen.

The bleaching by removing oxygen from Sulphur(IV) oxide gas is temporary. This is because the bleached dye regains the atomic oxygen from the atmosphere/air in presence of sunlight as catalyst thus regaining/restoring its original colour. e.g.

Old newspapers turn brown on exposure to air on regaining the atomic oxygen.

The bleaching through adding oxygen by chlorine gas is permanent.

(b)Turns Orange acidified potassium dichromate(VI) to green

Experiment:

(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containing acidified potassium dichromate(VI) solution. or;

(ii)Dip a filter paper soaked in acidified potassium dichromate(VI) into a gas jar containing Sulphur(IV) oxide gas.

Observation:

Orange acidified potassium dichromate(VI) turns to green.

Explanation:

Sulphur(IV) oxide gas reduces acidified potassium dichromate(VI) from orange $Cr_2O_7^{2-}$ ions to green Cr^{3+} ions without leaving a residue itself oxidized from SO_3^{2-} ions in sulphuric(IV) acid to SO_4^{2-} ions in sulphuric(VI) acid.

Chemical/ionic equation:

(i)Reaction of Sulphur(IV) oxide gas with water

$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

(ii)Dissociation /ionization of Sulphuric(IV)acid.

$H_2SO_3(aq) \rightarrow 2H^+(aq) + SO_3^{2-}(aq)$

(iii)Oxidation of $SO_3^{2-}(aq)$ and reduction of $Cr_2O_7^{2-}(aq)$

$3SO_3^{2-}(aq) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \rightarrow 3SO_4^{2-}(aq) + 2Cr^{3+}(aq) + 4H_2O(l)$

This is a **confirmatory** test for the presence of Sulphur(IV) oxide gas.

Hydrogen sulphide also reduces acidified potassium dichromate(VI) from orange $Cr_2O_7^{2-}$ ions to green Cr^{3+} ions leaving a **yellow** residue.

(c)Decolorizes acidified potassium manganate(VII)

Experiment:

(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containing acidified potassium manganate(VII) solution. or;

(ii)Dip a filter paper soaked in acidified potassium manganate(VII) into a gas jar containing Sulphur(IV) oxide gas.

Observation:

Purple acidified potassium manganate(VII) turns to colourless/ acidified potassium manganate(VII) is decolorized.

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Explanation:
Sulphur(IV) oxide gas reduces acidified potassium manganate(VII) from purple MnO_4^- ions to green Mn^{2+} ions without leaving a residue itself oxidized from SO_3^{2-} ions in sulphuric(IV) acid to SO_4^{2-} ions in sulphuric(VI) acid.
Chemical/ionic equation:
(i)Reaction of Sulphur(IV) oxide gas with water

$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

(ii)Dissociation /ionization of Sulphuric(IV)acid.

$H_2SO_3(aq) \rightarrow 2H^+(aq) + SO_3^{2-}(aq)$

(iii)Oxidation of SO_3^{2-} (aq)and reduction of MnO_4^- (aq)

$5SO_3^{2-}(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(l)$
(purple) (colourless)

This is another test for the presence of Sulphur(IV) oxide gas.
Hydrogen sulphide also decolorizes acidified potassium manganate(VII) from purple MnO_4^- ions to colourless Mn^{2+} ions leaving a **yellow** residue.

(d)Decolorizes bromine water

Experiment:
(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containing bromine water .
or;
(ii)Put three drops of bromine water into a gas jar containing Sulphur(IV) oxide gas. Swirl.
Observation:
Yellow bromine water turns to colourless/ bromine water is decolorized.
Explanation:
Sulphur(IV) oxide gas reduces yellow bromine water to colourless hydrobromic acid (HBr) without leaving a residue itself oxidized from SO_3^{2-} ions in sulphuric (IV) acid to SO_4^{2-} ions in sulphuric(VI) acid.
Chemical/ionic equation:
(i)Reaction of Sulphur(IV) oxide gas with water

$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

(ii)Dissociation /ionization of Sulphuric(IV)acid.

$H_2SO_3(aq) \rightarrow 2H^+(aq) + SO_3^{2-}(aq)$

(iii)Oxidation of SO_3^{2-} (aq)and reduction of MnO_4^- (aq)

$SO_3^{2-}(aq) + Br_2(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq) + 2HBr(aq)$
(yellow) (colourless)

This can also be used as another test for the presence of Sulphur(IV) oxide gas.
Hydrogen sulphide also decolorizes yellow bromine water to colourless leaving a **yellow** residue.

(e)Reduces Iron(III) Fe^{3+} **salts to Iron(II) salts** Fe^{2+}

Experiment:

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(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containing about 3 cm3 of Iron (III)chloride solution. or;

(ii)Place about 3cm3 of Iron (III)chloride solution into a gas jar containing Sulphur(IV) oxide gas.Swirl.
Observation:
Yellow/brown Iron (III)chloride solution turns to green
Explanation:
Sulphur(IV) oxide gas reduces Iron (III)chloride solution from yellow/brown Fe^{3+} ions to green Fe^{2+} ions without leaving a residue itself oxidized from SO_3^{2-} ions in sulphuric(IV) acid to SO_4^{2-} ions in sulphuric(VI) acid.
Chemical/ionic equation:
(i)Reaction of Sulphur(IV) oxide gas with water

$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

(ii)Dissociation /ionization of Sulphuric(IV)acid.

$H_2SO_3(aq) \rightarrow 2H^+(aq) + SO_3^{2-}(aq)$

(iii)Oxidation of SO_3^{2-} (aq)and reduction of Fe^{3+} (aq)

$SO_3^{2-}(aq) + 2Fe^{3+}(aq) + 3H_2O(l) \rightarrow SO_4^{2-}(aq) + 2Fe^{2+}(aq) + 2H^+(aq)$
(yellow) (green)

(f)Reduces Nitric(V)acid to Nitrogen(IV)oxide gas
Experiment:
(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containing about 3 cm3 of concentrated nitric(V)acid. or;
(ii)Place about 3cm3 of concentrated nitric(V)acid into a gas jar containing Sulphur(IV) oxide gas. Swirl.
Observation:
Brown fumes of a gas evolved/produced.
Explanation:
Sulphur(IV) oxide gas reduces concentrated nitric(V)acid to brown nitrogen(IV)oxide gas itself oxidized from SO_3^{2-} ions in sulphuric(IV) acid to SO_4^{2-} ions in sulphuric(VI) acid.
Chemical/ionic equation:

$SO_2(g) + 2HNO_3(l) \rightarrow H_2SO_4(l) + NO_2(g)$
(brown fumes/gas)

(g)Reduces Hydrogen peroxide to water
Experiment:
(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containing about 3 cm3 of 20 volume hydrogen peroxide. Add four drops of Barium nitrate(V)or Barium chloride followed by five drops of 2M hydrochloric acid/ 2M nitric(V) acid.
Observation:

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A white precipitate is formed that persist /remains on adding 2M hydrochloric acid/ 2M nitric(V) acid.

Explanation:
Sulphur(IV) oxide gas reduces 20 volume hydrogen peroxide and itself oxidized from SO_3^{2-} ions in sulphuric(IV) acid to SO_4^{2-} ions in sulphuric(VI) acid.
When Ba^{2+} ions in Barium Nitrate(V) or Barium chloride solution is added, a white precipitate of insoluble Barium salts is formed showing the presence of of either SO_3^{2-} ,SO_4^{2-} ,CO_3^{2-} ions. i.e.
Chemical/ionic equation:

$SO_3^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaSO_3(s)$
white precipitate

$SO_4^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaSO_4(s)$
white precipitate

$CO_3^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaCO_3(s)$
white precipitate

If nitric(V)/hydrochloric acid is added to the three suspected insoluble white precipitates above, the white precipitate:

(i) persist/remains if SO_4^{2-}(aq)ions ($BaSO_4(s)$) is present.
(ii)dissolves if SO_3^{2-}(aq)ions ($BaSO_3(s)$) and CO_3^{2-}(aq)ions ($BaCO_3(s)$)is present. This is because:

I. $BaSO_3(s)$ reacts with Nitric(V)/hydrochloric acid to produce acidic SO_2 gas that turns Orange moist filter paper dipped in acidified Potassium dichromate to green.
Chemical equation

$BaSO_3(s) + 2H^+(aq) \rightarrow Ba^{2+}(aq) + SO_2(g) + H_2O(l)$

I. $BaCO_3(s)$ reacts with Nitric(V)/hydrochloric acid to produce acidic CO_2 gas that forms a white precipitate when bubbled in lime water.
Chemical equation

$BaCO_3(s) + 2H^+(aq) \rightarrow Ba^{2+}(aq) + CO_2(g) + H_2O(l)$

5.Sulphur(IV)oxide also act as an oxidizing agent as in the following examples.
(a)Reduction by burning Magnesium
Experiment
Lower a burning Magnesium ribbon into agas jar containing Sulphur(IV)oxide gas
Observation
Magnesium ribbon continues to burn with difficulty.
White ash and yellow powder/speck
Explanation
Sulphur(IV)oxide does not support burning/combustion. Magnesium burns to produce enough heat energy to decompose Sulphur(IV)oxide to sulphur and oxygen.

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The metal continues to burn on Oxygen forming white Magnesium oxide solid/ash.
Yellow specks of sulphur residue form on the sides of reaction flask/gas jar.
During the reaction, Sulphur(IV)oxide is reduced(oxidizing agent)while the metal is oxidized (reducing agent)
Chemical equation

$SO_2(g) + 2Mg(s) \rightarrow 2MgO(s) + S(s)$
(white ash/solid) (**yellow** speck/powder)

(b)Reduction by Hydrogen sulphide gas
Experiment
Put two drops of water into a gas jar containing dry Sulphur(IV)oxide gas
Bubble hydrogen sulphide gas into the gas jar containing Sulphur(IV)oxide gas.
Or
Put two drops of water into a gas jar containing dry Sulphur(IV)oxide gas
Invert a gas jar full of hydrogen sulphide gas over the gas jar containing Sulphur(IV)oxide gas. Swirl
Observation
Yellow powder/speck
Explanation
Sulphur(IV)oxide oxidizes hydrogen sulphide to yellow specks of sulphur residue and itself reduced to also sulphur that form on the sides of reaction flask/gas jar.
A little moisture/water act as catalyst /speeds up the reaction.
Chemical equation

$SO_2(g) + 2H_2S(g) \rightarrow 2H_2O(l) + 3S(s)$
(yellow speck/powder)

6.Sulphur(IV)oxide has many industrial uses. State three.
(i)In the contact process for the manufacture of Sulphuric(VI)acid
(ii)As a bleaching agent of pulp and paper.
(iii)As a fungicide to kill microbes'
(iv)As a preservative of jam, juices to prevent fermentation

(ii) Sulphur(VI)oxide(SO_3)
(a) Occurrence
Sulphur (VI)oxide is does not occur free in nature/atmosphere
(b) Preparation
In a Chemistry school laboratory Sulphur (VI)oxide may prepared from:
Method 1;**Catalytic oxidation of sulphur(IV)oxide gas**.
Sulphur(IV)oxide gas and oxygen mixture are first dried by being passed through Concentrated Sulphuric(VI)acid .
The dry mixture is then passed through platinised asbestos to catalyse/speed up the combination to form Sulphur (VI)oxide gas.
Sulphur (VI)oxide gas readily solidify as silky white needles if passed through a freezing mixture /ice cold water.

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The solid fumes out on heating to a highly acidic poisonous gas.
Chemical equation

$$2SO_2(g) + O_2(g) \xrightarrow{\text{platinised asbestos}} 2SO_3(g)$$

Method 2; Heating Iron(II)sulphate(VI) heptahydrate
When green hydrated Iron(II)sulphate(VI) heptahydrate crystals are heated in a boiling tube ,it loses the water of crystallization and colour changes from green to white.
Chemical equation

$$FeSO_4.7H_2O(s) \rightarrow FeSO_4(s) + 7H_2O(l)$$

(green solid) (white solid)

On further heating ,the white anhydrous Iron(II)sulphate(VI) solid decomposes to a mixture of Sulphur (VI)oxide and Sulphur (IV)oxide gas.
Sulphur (VI) oxide readily / easily solidify as white silky needles when the mixture is passed through a freezing mixture/ice cold water.
Iron(III)oxide is left as a brown residue/solid.
Chemical equation

$$2FeSO_4(s) \rightarrow Fe_2O_3(s) + SO_2(g) + SO_3(g)$$

(green solid) (brown solid)

Caution
On exposure to air Sulphur (VI)oxide gas produces highly corrosive poisonous fumes of concentrated sulphuric(VI)acid and thus its preparation in a school laboratory is very risky.
(c) Uses of sulphur(VI)oxide
One of the main uses of sulphur(VI)oxide gas is as an intermediate product in the contact process for industrial/manufacture/large scale/production of sulphuric(VI)acid.
(iii) Sulphuric(VI)acid(H_2SO_4)
(a) Occurrence
Sulphuric (VI)acid(H_2SO_4) is one of the three mineral acids. There are three mineral acids;

Nitric(V)acid
Sulphuric(VI)acid
Hydrochloric acid.

Mineral acids do not occur naturally but are prepared in a school laboratory and manufactured at industrial level.
(b)The Contact process for industrial manufacture of H_2SO_4 .
I. Raw materials
The main raw materials for industrial preparation of Sulphuric(VI)acid include:

(i)**Sulphur** from Fraschs process or from heating metal sulphide ore like Galena(PbS),Zinc blende(ZnS)

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(ii)**Oxygen** from fractional distillation of air
(iii)**Water** from rivers/lakes

II. Chemical processes

The contact process involves four main chemical processes:

(i)Production of Sulphur (IV)oxide

As one of the raw materials, Sulphur (IV)oxide gas is got from the following sources;

I. Burning/roasting sulphur in air.

Sulphur from Fraschs process is roasted/burnt in air to form Sulphur (IV)oxide gas in the **burners**

Chemical equation

$S(s) + O_2(g) \rightarrow SO_2(g)$

II. Burning/roasting sulphide ores in air.

Sulphur (IV)oxide gas is produced as a by product in extraction of some metals like:

- Lead from Lead(II)sulphide/Galena,(PbS)
- Zinc from zinc(II)sulphide/Zinc blende, (ZnS)
- Copper from Copper iron sulphide/Copper pyrites, ($CuFeS_2$)

On roasting/burning, large amount /quantity of sulphur(IV)oxide is generated/produced.

Chemical equation

(i) $2PbS(s) + 3O_2(g) \rightarrow 2PbO(s) + 2SO_2(g)$

(ii) $2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$

(ii) $2CuFeS_2(s) + 4O_2(g) \rightarrow 2FeO(s) + 3SO_2(g) + Cu_2O(s)$

Sulphur(IV)oxide easily/readily liquefies and thus can be transported to a far distance safely.

(ii)Purification of Sulphur(IV)oxide

Sulphur(IV)oxide gas contain dust particles and Arsenic(IV)oxide as impurities. These impurities "poison"/impair the catalyst by adhering on/covering its surface. The impurities are removed by electrostatic precipitation method .

In the contact process Platinum or Vanadium(V)oxide may be used.

Vanadium(V)oxide is preferred because it is :

(i) cheaper/less expensive
(ii) less easily poisoned by impurities

(iii)Catalytic conversion of Sulphur(IV)oxide to Sulphur(VI)oxide

Pure and dry mixture of Sulphur (IV)oxide gas and Oxygen is heated to 450^oC in a heat exchanger.

The heated mixture is passed through long pipes coated with pellets of Vanadium (V)oxide catalyst.

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The close "contact" between the reacting gases and catalyst give the process its name.
Vanadium (V)oxide catalyse the conversion/oxidation of Sulphur(IV)oxide to Sulphur(VI)oxide gas.
Chemical equation

$$2SO_2\,(g) + O_2(g) \xrightarrow{-- V_2O_5 -->} 2SO_2\,(g)$$

This reaction is exothermic ($-\Delta H$) and the temperatures need to be maintained at around 450°C to ensure that:

(**i**)reaction rate/time taken for the formation of Sulphur(VI)oxide is not too **slow/long** at **lower** temperatures below 450°C

(**ii**) Sulphur(VI)oxide gas does not **decompose** back to Sulphur(IV)oxide gas and Oxygen gas at **higher** temperatures than 450°C.

(iv)Conversion of Sulphur(VI)oxide of Sulphuric(VI)acid

Sulphur(VI)oxide is the acid anhydride of concentrated Sulphuric(VI)acid.
Sulphur(VI)oxide reacts with water to form thick mist of fine droplets of very/highly corrosive concentrated Sulphuric(VI)acid because the reaction is highly exothermic.
To prevent this, Sulphur (VI)oxide is a passed up to meet downward flow of 98% Sulphuric(VI)acid in the absorption chamber/tower.
The reaction forms a very **viscous oily** liquid called **Oleum/fuming Sulphuric (VI) acid/ pyrosulphuric (VI) acid.**
Chemical equation

$$H_2SO_4\,(aq) + SO_3\,(g) \rightarrow H_2S_2O_7\,(l)$$

Oleum/fuming Sulphuric (VI) acid/ pyrosulphuric (VI) acid is diluted carefully with distilled water to give concentrated sulphuric (VI) acid .
Chemical equation

$$H_2S_2O_7\,(l) + H_2O\,(l) \rightarrow 2H_2SO_4\,(l)$$

The acid is stored ready for market/sale.

III. Environmental effects of contact process

Sulphur(VI)oxide and Sulphur(IV)oxide gases are atmospheric pollutants that form acid rain if they escape to the atmosphere.
In the Contact process, about 2% of these gases do not form sulphuric (VI) acid.
The following precautions prevent/minimize pollution from Contact process:

(i)recycling back any unreacted Sulphur(IV)oxide gas back to the heat exchangers.

(ii)dissolving Sulphur(VI)oxide gas in concentrated sulphuric (VI) acid instead of water.

This prevents the formation of fine droplets of the corrosive/ toxic/poisonous fumes of concentrated sulphuric (VI) acid.

(iii)**scrubbing**-This involves passing the exhaust gases through very tall chimneys lined with quicklime/calcium hydroxide solid.

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This reacts with Sulphur (VI)oxide gas forming harmless calcium(II)sulphate (IV) /$CaSO_3$

Chemical equation

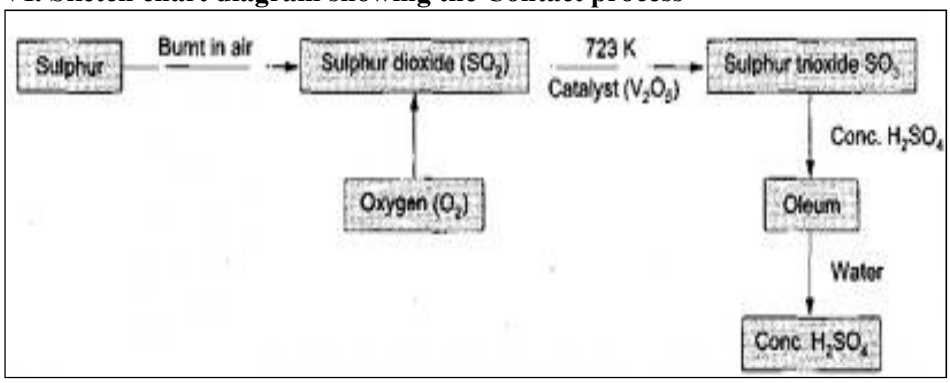
$Ca(OH)_2\ (aq) + SO_2(g) \quad \rightarrow \quad CaSO_3\ (aq) + H_2O\ (g)$

IV. Uses of Sulphuric(VI)acid

Sulphuric (VI) acid is used:

(i) in making dyes and paint
(ii)as acid in Lead-acid accumulator/battery
(iii) for making soapless detergents
(iv) for making sulphate agricultural fertilizers

VI. Sketch chart diagram showing the Contact process



(c) Properties of Concentrated sulphuric (VI) acid (i) Concentrated sulphuric (VI) acid is a colourless oily liquid with a density of $1.84gcm^{-3}$.It has a boiling point of 338ºC.

(ii) Concentrated sulphuric (VI) acid is very soluble in water.

The solubility /dissolution of the acid very highly exothermic.

The concentrated acid should thus be diluted slowly in excess water.

Water should never be added to the acid because the hot acid scatters highly corrosive fumes out of the container.

(iii) Concentrated sulphuric (VI)acid is a covalent compound. It has no free H^+ ions.

Free H^+ ions are responsible for turning the blue litmus paper red. Concentrated sulphuric (VI) acid thus do not change the blue litmus paper red.

(iv) Concentrated sulphuric (VI)acid is hygroscopic. It absorbs water from the atmosphere and do not form a solution.

This makes concentrated sulphuric (VI) acid very suitable as drying agent during preparation of gases.

(v)The following are some **chemical properties** of concentrated sulphuric (VI) acid:

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I. As a dehydrating agent

Experiment I;

Put about four spatula end full of brown sugar and glucose in separate 10cm3 beaker.

Carefully add about 10cm3 of concentrated sulphuric (VI) acid .Allow to stand for about 10 minutes.

Observation;

Colour(in brown sugar)change from brown to **black**.

Colour (in glucose) change from white to **black**.

10cm3 beaker becomes very hot.

Explanation

Concentrated sulphuric (VI) acid is strong dehydrating agent.

It removes chemically and physically combined elements of water(Hydrogen and Oxygen in ratio 2:1)from compounds.

When added to sugar /glucose a vigorous reaction that is highly exothermic take place.

The sugar/glucose is **charred** to **black mass** of carbon because the acid dehydrates the sugar/glucose leaving carbon.

Caution

This reaction is highly exothermic that start slowly but produce fine particles of carbon that if inhaled cause quick suffocation by blocking the lung villi.

Chemical equation

Glucose: $C_6H_{12}O_6(s)$ --conc.H_2SO_4--> 6C (**s**) + $6H_2O(l)$
(white) (**black**)

Sugar: $C_{12}H_{22}O_{11}(s)$ --conc.H_2SO_4--> 12C (**s**) +$11H_2O(l)$
(brown) (**black**)

Experiment II;

Put about two spatula end full of hydrated copper(II)sulphate(VI)crystals in a boiling tube .Carefully add about 10cm3 of concentrated sulphuric (VI) acid .Warm .

Observation;

Colour change from blue to white.

Explanation

Concentrated sulphuric (VI) acid is strong dehydrating agent.It removes physically combined elements of water(Hydrogen and Oxygen in ratio 2:1)from hydrated compounds.

The acid dehydrates blue copper(II)sulphate to white anhydrous copper(II)sulphate .

Chemical equation

$CuSO_4.5H_2O(s)$ --conc.H_2SO_4--> $CuSO_4$ (s) + $5H_2O(l)$
(blue) (white)

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Experiment III;
Put about 4cm3 of absolute ethanol in a boiling tube .Carefully add about 10cm3 of concentrated sulphuric (VI) acid.
Place moist/damp/wet filter paper dipped in acidified potassium dichromate(VI)solution on the mouth of the boiling tube. Heat strongly.
Caution:
Absolute ethanol is highly flammable.
Observation;
Colourless gas produced.
Orange acidified potassium dichromate (VI) paper turns to green.
Explanation
Concentrated sulphuric (VI) acid is strong dehydrating agent.
It removes chemically combined elements of water(Hydrogen and Oxygen in ratio 2:1)from compounds.
The acid dehydrates ethanol to ethene gas at about 170°C.
Ethene with =C=C= double bond turns orange acidified potassium dichromate (VI) paper turns to green.
Chemical equation

$$C_2H_5OH(l) \xrightarrow{\text{conc.}H_2SO_4/170^oC} C_2H_4(g) + H_2O(l)$$

NB: This reaction is used for the school laboratory preparation of ethene gas
Experiment IV;
Put about 4cm3 of methanoic acid in a boiling tube .Carefully add about 6 cm3 of concentrated sulphuric (VI) acid. Heat gently
Caution:
This should be done in a fume chamber/open
Observation;
Colourless gas produced.
Explanation
Concentrated sulphuric (VI) acid is strong dehydrating agent. It removes chemically combined elements of water (Hydrogen and Oxygen in ratio 2:1)from compounds.
The acid dehydrates methanoic acid to poisonous/toxic carbon(II)oxide gas.
Chemical equation

$$HCOOH(l) \xrightarrow{\text{conc.}H_2SO_4} CO(g) + H_2O(l)$$

NB: This reaction is used for the school laboratory preparation of small amount carbon (II)oxide gas
Experiment V;
Put about 4cm3 of ethan-1,2-dioic/oxalic acid in a boiling tube .Carefully add about 6 cm3 of concentrated sulphuric (VI) acid. Pass any gaseous product through lime water. Heat gently
Caution:

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This should be done in a fume chamber/open
Observation;
Colourless gas produced.
Gas produced forms a white precipitate with lime water.
Explanation
Concentrated sulphuric (VI) acid is strong dehydrating agent.
It removes chemically combined elements of water (Hydrogen and Oxygen in ratio 2:1)from compounds.
The acid dehydrates ethan-1,2-dioic/oxalic acid to a mixture of poisonous/toxic carbon(II)oxide and carbon(IV)oxide gases.
Chemical equation

$$HOOCCOOH(l) \xrightarrow{conc.H_2SO_4} CO(g) + CO_2(g) + H_2O(l)$$

NB: This reaction is also used for the school laboratory preparation of small amount carbon (II) oxide gas.
Carbon (IV) oxide gas is removed by passing the mixture through concentrated sodium/potassium hydroxide solution.

II. As an Oxidizing agent
Experiment I
Put about 2cm3 of Concentrated sulphuric (VI) acid into three separate boiling tubes. Place a thin moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution on the mouth of the boiling tube. Put about 0.5g of Copper turnings, Zinc granule and Iron filings to each boiling tube separately.
Observation;
Effervescence/fizzing/bubbles
Blue solution formed with copper,
Green solution formed with Iron
Colourless solution formed with Zinc
Colourless gas produced that has a pungent irritating choking smell.
Gas produced turn orange moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution to green.
Explanation
Concentrated sulphuric (VI) acid is strong oxidizing agent.
It oxidizes metals to metallic sulphate(VI) salts and itself reduced to sulphur(IV)oxide gas.
Sulphur (IV) oxide gas turn orange moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution to green.
$CuSO_4(aq)$ is a blue solution. $ZnSO_4(aq)$ is a colourless solution. $FeSO_4(aq)$ is a green solution.
Chemical equation

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$Cu(s) + 2H_2SO_4(aq) \rightarrow CuSO_4(aq) + SO_2(g) + 2H_2O(l)$

$Zn(s) + 2H_2SO_4(aq) \rightarrow ZnSO_4(aq) + SO_2(g) + 2H_2O(l)$

$Fe(s) + 2H_2SO_4(aq) \rightarrow FeSO_4(aq) + SO_2(g) + 2H_2O(l)$

Experiment II

Put about 2cm3 of Concentrated sulphuric (VI) acid into two separate boiling tubes.
Place a thin moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution on the mouth of the boiling tube.
Put about 0.5g of powdered charcoal and sulphur powder to each boiling tube separately.
Warm.
Observation;

Black solid charcoal dissolves/decrease
Yellow solid sulphur dissolves/decrease
Colourless gas produced that has a pungent irritating choking smell.
Gas produced turn orange moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution to green.

Explanation
Concentrated sulphuric (VI) acid is strong oxidizing agent. It oxidizes non-metals to non metallic oxides and itself reduced to sulphur(IV)oxide gas. Sulphur (IV) oxide gas turn orange moist/damp/wet filter paper dipped in acidified potassium dichromate (VI)solution to green.
Charcoal is oxidized to carbon(IV)oxide. Sulphur is oxidized to Sulphur(IV)oxide .
Chemical equation

$C(s) + 2H_2SO_4(aq) \rightarrow CO_2(aq) + 2SO_2(g) + 2H_2O(l)$

$S(s) + 2H_2SO_4(aq) \rightarrow 3SO_2(g) + 2H_2O(l)$

III. As the least volatile acid

Study the table below showing a comparison in boiling points of the three mineral acids

Mineral acid	Relative molecula mass	Boiling point(°C)
Hydrochloric acid(HCl)	36.5	35.0
Nitric(V)acid(HNO_3)	63.0	83.0
Sulphuric(VI)acid(H_2SO_4)	98.0	333

1.Which is the least volatile acid? Explain

Sulphuric(VI)acid(H_2SO_4) because it has the largest molecule and joined by Hydrogen bonds making it to have the highest boiling point/least volatile.

2. Using chemical equations, explain how sulphuric(VI)acid displaces the less volatile mineral acids.

(i)Chemical equation

$KNO_3(s) + H_2SO_4(aq) \rightarrow KHSO_4(l) + HNO_3(g)$

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$NaNO_3(s) + H_2SO_4(aq) \rightarrow NaHSO_4(l) + HNO_3(g)$

This reaction is used in the school laboratory preparation of Nitric(V) acid (HNO_3).

(ii)Chemical equation

$KCl(s) + H_2SO_4(aq) \rightarrow KHSO_4(s) + HCl(g)$

$NaCl(s) + H_2SO_4(aq) \rightarrow NaHSO_4(s) + HCl(g)$

This reaction is used in the school laboratory **preparation** of Hydrochloric acid (HCl).

(d) Properties of dilute sulphuric(VI)acid.

Dilute sulphuric(VI)acid is made when about 10cm3 of concentrated sulphuric (VI) acid is carefully added to about 90cm3 of distilled water.

Diluting concentrated sulphuric (VI) acid should be done carefully because the reaction is highly exothermic.

Diluting concentrated sulphuric (VI) acid decreases the number of moles present in a given volume of solution which makes the acid less corrosive.

On diluting concentrated sulphuric(VI) acid, water ionizes /dissociates the acid fully/wholly into two(**dibasic**)free $H^+(aq)$ and $SO_4^{2-}(aq)$ions:

$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$

The presence of free $H^+(aq)$ions is responsible for ;

(**i**)turn litmus red because of the presence of free $H^+(aq)$ions

(ii)have pH 1/2/3 because of the presence of many free $H^+(aq)$ions hence a strongly acidic solution.

(iii)Reaction with metals

Experiment:

Place 5cm3 of 0.2M dilute sulphuric(VI)acid into four separate clean test tubes. Add about 0.1g of Magnesium ribbon to one test tube. Cover the mixture with a finger as stopper. Introduce a burning splint on top of the finger and release the finger "stopper". Repeat by adding Zinc, Copper and Iron instead of the Magnesium ribbon.

Observation:

No effervescence/ bubbles/ fizzing with copper

Effervescence/ bubbles/ fizzing with Iron ,Zinc and Magnesium

Colourless gas produced that extinguishes burning splint with a "pop" sound.

Colourless solution formed with Zinc and Magnesium.

Green solution formed with Iron

Explanation:

When a metal higher than hydrogen in the reactivity/electrochemical series is put in a test tube containing dilute sulphuric(VI)acid, effervescence/ bubbling/ fizzing takes place with evolution of Hydrogen gas.

Impure hydrogen gas extinguishes burning splint with a "pop" sound.

A sulphate (VI) salts is formed. Iron, Zinc and Magnesium are higher than hydrogen in the reactivity/electrochemical series.

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They form Iron (II)sulphate(VI), Magnesium sulphate(VI) and Zinc sulphate(VI).

. When a metal lower than hydrogen in the reactivity/electrochemical series is put in a test tube containing dilute sulphuric(VI)acid, there is no effervescence/ bubbling/ fizzing that take place.

Copper thus do not react with dilute sulphuric(VI)acid.

Chemical/ionic equation

$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

$Fe(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2(g)$

$Fe(s) + H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$

NB:(i) Calcium,Lead and Barium forms insoluble sulphate(VI)salts that cover/coat the unreacted metals.

(ii)Sodium and Potassium react explosively with dilute sulphuric(VI)acid

(**iv**)Reaction with metal carbonates and hydrogen carbonates

Experiment:

Place 5cm3 of 0.2M dilute sulphuric(VI)acid into four separate clean boiling tubes. Add about 0.1g of sodium carbonate to one boiling tube. Introduce a burning splint on top of the boiling tube. Repeat by adding Zinc carbonate, Copper (II)carbonate and Iron(II)Carbonate in place of the sodium hydrogen carbonate.

Observation:

Effervescence/ bubbles/ fizzing.

Colourless gas produced that extinguishes burning splint.

Colourless solution formed with Zinc carbonate, sodium hydrogen carbonate and sodium carbonate.

Green solution formed with Iron(II)Carbonate

Blue solution formed with Copper(II)Carbonate

Explanation:

When a metal carbonate or a hydrogen carbonates is put in a test tube containing dilute sulphuric(VI)acid, effervescence/ bubbling/ fizzing takes place with evolution of carbon(IV)oxide gas. carbon(IV)oxide gas extinguishes a burning splint and forms a white precipitate when bubbled in lime water.

A sulphate (VI) salts is formed.

Chemical/ionic equation

$ZnCO_3(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2O(l) + CO_2(g)$

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$ZnCO_3(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2O(l) + CO_2(g)$
$CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l) + CO_2(g)$
$CuCO_3(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l) + CO_2(g)$
$FeCO_3(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2O(l) + CO_2(g)$
$FeCO_3(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2O(l) + CO_2(g)$
$2NaHCO_3(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l) + 2CO_2(g)$
$NaHCO_3(s) + H^+(aq) \rightarrow Na^+(aq) + H_2O(l) + CO_2(g)$
$Na_2CO_3(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)$
$NaHCO_3(s) + H^+(aq) \rightarrow Na^+(aq) + H_2O(l) + CO_2(g)$
$(NH_4)_2CO_3(s) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq) + H_2O(l) + CO_2(g)$
$(NH_4)_2CO_3(s) + H^+(aq) \rightarrow NH_4^+(aq) + H_2O(l) + CO_2(g)$
$2NH_4HCO_3(aq) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq) + H_2O(l) + CO_2(g)$
$NH_4HCO_3(aq) + H^+(aq) \rightarrow NH_4^+(aq) + H_2O(l) + CO_2(g)$

NB:

Calcium, Lead and Barium carbonates forms insoluble sulphate(VI)salts that cover/coat the unreacted metals.

(**v**)Neutralization-reaction of metal oxides and alkalis/bases

Experiment I:

Place 5cm3 of 0.2M dilute sulphuric(VI)acid into four separate clean boiling tubes. Add about 0.1g of copper(II)oxide to one boiling tube. Stir.

Repeat by adding Zinc oxide, calcium carbonate and Sodium (II)Oxide in place of the Copper(II)Oxide.

Observation:

Blue solution formed with Copper(II)Oxide

Colourless solution formed with other oxides

Explanation:

When a metal oxide is put in a test tube containing dilute sulphuric(VI)acid, the oxide dissolves forming a sulphate (VI) salt.

Chemical/ionic equation

$ZnO(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2O(l)$
$ZnO(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2O(l)$
$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$
$CuO(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l)$
$MgO(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2O(l)$
$MgO(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2O(l)$
$Na_2O(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$
$Na_2O(s) + 2H^+(aq) \rightarrow 2Na^+(aq) + H_2O(l)$
$K_2CO_3(s) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + H_2O(l)$
$K_2O(s) + H^+(aq) \rightarrow 2K^+(aq) + H_2O(l)$

NB:

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Calcium, Lead and Barium oxides forms insoluble sulphate(VI)salts that cover/coat the unreacted metals oxides.

Experiment II:

Fill a burette with 0.1M dilute sulphuric(VI)acid. Pipette 20.0cm3 of 0.1Msodium hydroxide solution into a 250cm3 conical flask. Add three drops of phenolphthalein indicator. Titrate the acid to get a permanent colour change. Repeat with0.1M potassium hydroxide solution inplace of 0.1Msodium hydroxide solution

Observation:

Colour of phenolphthalein changes from pink to colourless at the end point.

Explanation

Like other (mineral) acids dilute sulphuric(VI)acid neutralizes bases/alkalis to a sulphate salt and water only.

Colour of the indicator used changes when a slight excess of acid is added to the base at the end point

Chemical equation:

$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$

$OH^-(s) + H^+(aq) \rightarrow H_2O(l)$

$2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + H_2O(l)$

$OH^-(s) + H^+(aq) \rightarrow H_2O(l)$

$2NH_4OH(aq) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq) + H_2O(l)$

$OH^-(s) + H^+(aq) \rightarrow H_2O(l)$

(iv) Hydrogen sulphide(H_2S)

(a) Occurrence

Hydrogen sulphide is found in volcanic areas as a gas or dissolved in water from geysers and hot springs in active volcanic areas of the world e.g. Olkaria and Hells gate near Naivasha in Kenya.

It is present in rotten eggs and human excreta.

(b) Preparation

Hydrogen sulphide is prepared in a school laboratory by heating Iron (II) sulphide with dilute hydrochloric acid.

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(c) Properties of Hydrogen sulphide(Questions)

1. Write the equation for the reaction for the school laboratory preparation of Hydrogen sulphide.

Chemical equation: $FeS\ (s) + 2HCl\ (aq) \rightarrow H_2S\ (g)\ FeCl_2\ (aq)$

2. State three physical properties unique to Hydrogen sulphide.

Hydrogen sulphide is a colourless gas with characteristic pungent poisonous smell of rotten eggs. It is soluble in cold water but insoluble in warm water. It is denser than water and turns blue litmus paper red.

3. Hydrogen sulphide exist as a dibasic acid when dissolved in water. Using a chemical equation show how it ionizes in aqueous state.

$H_2S(aq) \rightarrow H^+(aq) + HS^-(aq)$

$H_2S(aq) \rightarrow 2H^+(aq) + S^{2-}(aq)$

Hydrogen sulphide therefore can form both normal and acid salts e.g

Sodium hydrogen sulphide and sodium sulphide both exist

4. State and explain one gaseous impurity likely to be present in the gas jar containing hydrogen sulphide above.

Hydrogen/ H_2

Iron(II)sulphide contains Iron as impurity .The iron will react with dilute hydrochloric acid to form iron(II)chloride and produce hydrogen gas that mixes with hydrogen sulphide gas.

5. State and explain the observations made when a filter paper dipped in Lead(II) ethanoate /Lead (II) nitrate(V) solution is put in a gas jar containing hydrogen sulphide gas.

Observations

Moist Lead(II) ethanoate /Lead (II) nitrate(V) paper turns black.

Explanation

When hydrogen sulphide is bubbled in a metallic salt solution, a metallic sulphide is formed.

All sulphides are insoluble black salts except sodium sulphide, potassium sulphide and ammonium sulphides.

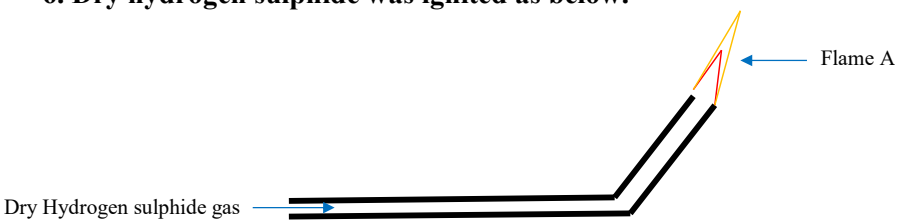
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Hydrogen sulphide gas blackens moist Lead (II) ethanoate /Lead (II) nitrate(V) paper .
The gas reacts with Pb^{2+} in the paper to form black Lead(II)sulphide.
This is the chemical test for the presence of H_2S other than the physical smell of rotten eggs.
Chemical equations

$Pb^{2+}(aq)$	$+ H_2S$	-> PbS (black)	$+ 2H^+(aq)$
$Fe^{2+}(aq)$	$+ H_2S$	-> FeS (black)	$+ 2H^+(aq)$
$Zn^{2+}(aq)$	$+ H_2S$	-> ZnS (black)	$+ 2H^+(aq)$
$Cu^{2+}(aq)$	$+ H_2S$	-> CuS (black)	$+ 2H^+(aq)$
$2Cu^{+}(aq)$	$+ H_2S$	-> Cu_2S (black)	$+ 2H^+(aq)$

6. Dry hydrogen sulphide was ignited as below.



(i) State the observations made in flame A
Hydrogen sulphide burns in excess air with a blue flame to form sulphur(IV)oxide gas and water.
Chemical equation: $2H_2S(g) + 3O_2(g) \rightarrow 2H_2O(l) + 2SO_2(g)$
Hydrogen sulphide burns in limited air with a blue flame to form sulphur solid and water.
Chemical equation: $2H_2S(g) + O_2(g) \rightarrow 2H_2O(l) + 2S(s)$

7. Hydrogen sulphide is a strong reducing agent that is oxidized to yellow solid sulphur as precipitate. The following experiments illustrate the reducing properties of Hydrogen sulphide.
(a)Turns Orange acidified potassium dichromate(VI) to green
Experiment:
(i)Pass a stream of Hydrogen sulphide gas in a test tube containing acidified potassium dichromate (VI) solution. or;
(ii)Dip a filter paper soaked in acidified potassium dichromate (VI) into a gas jar containing Hydrogen sulphide gas.

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Observation:
Orange acidified potassium dichromate (VI) turns to green.
Yellow solid residue.
Explanation:
Hydrogen sulphide gas reduces acidified potassium dichromate(VI) from orange $Cr_2O_7^{2-}$ ions to green Cr^{3+} ions leaving a yellow solid residue as itself is oxidized to sulphur.
Chemical/ionic equation:
$4H_2S(aq) + Cr_2O_7^{2-}(aq) + 6H^+(aq) \rightarrow 4S(aq) + 2Cr^{3+}(aq) + 7H_2O(l)$
This test is used for differentiating Hydrogen sulphide and sulphur (IV)oxide gas.
Sulphur(IV)oxide also reduces acidified potassium dichromate(VI) from orange $Cr_2O_7^{2-}$ ions to green Cr^{3+} ions without leaving a **yellow** residue.

(b)Decolorizes acidified potassium manganate(VII)

Experiment:
(i)Pass a stream of Sulphur(IV) oxide gas in a test tube containing acidified potassium manganate(VII) solution. or;
(ii)Dip a filter paper soaked in acidified potassium manganate(VII) into a gas jar containing Hydrogen Sulphide gas.
Observation:
Purple acidified potassium manganate(VII) turns to colourless/ acidified potassium manganate(VII) is decolorized.
Yellow solid residue.
Explanation:
Hydrogen sulphide gas reduces acidified potassium manganate(VII) from purple MnO_4^- ions to green Mn^{2+} ions leaving a residue as the gas itself is oxidized to sulphur.
Chemical/ionic equation:
$5H_2S(g) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 5S(s) + 2Mn^{2+}(aq) + 8H_2O(l)$
(purple) (colourless)
This is another test for differentiating Hydrogen sulphide and Sulphur(IV) oxide gas.
Sulphur(IV) oxide also decolorizes acidified potassium manganate(VII) from purple MnO_4^- ions to colourless Mn^{2+} ions leaving no **yellow** residue.

(c)Decolorizes bromine water

Experiment:
(i)Pass a stream of Hydrogen sulphide gas in a test tube containing bromine water . or;
(ii)Put three drops of bromine water into a gas jar containing Hydrogen sulphide gas. Swirl.
Observation:

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Yellow bromine water turns to colourless/ bromine water is decolorized.
Yellow solid residue
Explanation:
Hydrogen sulphide gas reduces yellow bromine water to colourless hydrobromic acid (HBr) leaving a yellow residue as the gas itself is oxidized to sulphur.
Chemical/ionic equation:

$H_2S(g) + Br_2(aq) \rightarrow S(s) + 2HBr(aq)$
(yellow solution) (yellow solid) (colourless)

This is another test for differentiating Hydrogen sulphide and Sulphur(IV) oxide gas.
Sulphur(IV) oxide also decolorizes acidified potassium manganate(VII) from purple MnO_4^- ions to colourless Mn^{2+} ions leaving no **yellow** residue.

(d)Reduces Iron(III) Fe^{3+} **salts to Iron(II) salts** Fe^{2+}

Experiment:
(i)Pass a stream of Hydrogen sulphide gas in a test tube containing about 3 cm3 of Iron (III)chloride solution. or;
(ii)Place about 3cm3 of Iron (III)chloride solution into a gas jar containing Hydrogen sulphide gas. Swirl.
Observation:
Yellow/brown Iron (III)chloride solution turns to green.
Yellow solid
Explanation:
Hydrogen sulphide gas reduces Iron (III)chloride solution from yellow/brown Fe^{3+} ions to green Fe^{2+} ions leaving a yellow residue.The gas is itself oxidized to sulphur.
Chemical/ionic equation:

$H_2S(aq) + 2Fe^{3+}(aq) \rightarrow S(s) + Fe^{2+}(aq) + 2H^+(aq)$
(yellow solution) (yellow residue) (green)

(e)Reduces Nitric(V)acid to Nitrogen(IV)oxide gas

Experiment:
(i)Pass a stream of Hydrogen sulphide gas in a test tube containing about 3 cm3 of concentrated nitric(V)acid. or;
(ii)Place about 3cm3 of concentrated nitric(V)acid into a gas jar containing Hydrogen sulphide gas. Swirl.
Observation:
Brown fumes of a gas evolved/produced.
Yellow solid residue

Explanation:
Hydrogen sulphide gas reduces concentrated nitric(V)acid to brown nitrogen(IV)oxide gas itself oxidized to yellow sulphur.

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Chemical/ionic equation:

$H_2S(g) + 2HNO_3(l) \rightarrow 2H_2O(l) + S(s) + 2NO_2(g)$

(yellow residue) (brown fumes)

(f)Reduces sulphuric(VI)acid to Sulphur

Experiment:

(i)Pass a stream of Hydrogen sulphide gas in a test tube containing about 3 cm3 of concentrated sulphuric(VI)acid. or;

(ii)Place about 3cm3 of concentrated sulphuric (VI) acid into a gas jar containing Hydrogen sulphide gas. Swirl.

Observation:

Yellow solid residue

Explanation:

Hydrogen sulphide gas reduces concentrated sulphuric(VI)acid to yellow sulphur.

Chemical/ionic equation:

$3H_2S(g) + H_2SO_4(l) \rightarrow 4H_2O(l) + 4S(s)$

(yellow residue)

(g)Reduces Hydrogen peroxide to water

Experiment:

(i)Pass a stream of Hydrogen sulphide gas in a test tube containing about 3 cm3 of 20 volume hydrogen peroxide.

Observation:

Yellow solid residue

Explanation:

Hydrogen sulphide gas reduces 20 volume hydrogen peroxide to water and itself oxidized to yellow sulphur

Chemical/ionic equation:

$H_2S(g) + H_2O_2(l) \rightarrow 2H_2O(l) + S(s)$

(yellow residue)

8.Name the salt formed when:

(i)equal volumes of equimolar hydrogen sulphide neutralizes sodium hydroxide solution:

Sodium hydrogen sulphide

Chemical/ionic equation:

$H_2S(g) + NaOH(l) \rightarrow H_2O(l) + NaHS(aq)$

(ii) hydrogen sulphide neutralizes excess concentrated sodium hydroxide solution:

Sodium sulphide

Chemical/ionic equation:

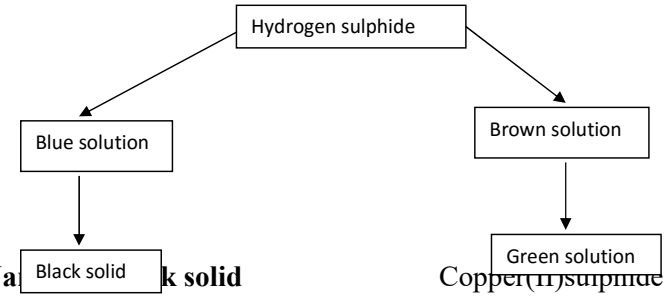
$H_2S(g) + 2NaOH(l) \rightarrow 2H_2O(l) + Na_2S(aq)$

Practice

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Hydrogen sulphide gas was bubbled into a solution of metallic nitrate(V)salts as in the flow chart below



(a)Na k solid Copper(II)sulphide

(b)Identify the cation responsible for the formation of:

I. Blue solution $Cu^{2+}(aq)$

II. Green solution $Fe^{2+}(aq)$

III. Brown solution $Fe^{3+}(aq)$

(c)Using acidified potassium dichromate(VI) describe how you would differentiate between sulphur(IV)Oxide and hydrogen sulphide

-Bubble the gases in separate test tubes containing acidified Potassium dichromate(VI) solution.

-Both changes the Orange colour of acidified Potassium dichromate(VI) solution to green.

-Yellow solid residue/deposit is formed with Hydrogen sulphide

Chemical/ionic equation:

$4H_2S(aq) + Cr_2O_7^{2-}(aq) + 6H^+(aq) \rightarrow 4S(aq) + 2Cr^{3+}(aq) + 7H_2O(l)$

$3SO_3^{2-}(aq) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \rightarrow 3SO_4^{2-}(aq) + 2Cr^{3+}(aq) + 4H_2O(l)$

(d)State and explain the observations made if a burning splint is introduced at the mouth of a hydrogen sulphide generator.

Observation

Gas continues burning with a blue flame

Explanation: Hydrogen sulphide burns in excess air with a blue flame to form sulphur(IV)oxide gas and water.

Chemical equation: $2H_2S(g) + 3O_2(g) \rightarrow 2H_2O(l) + 2SO_2(g)$

(v)Sulphate (VI) (SO_4^{2-})and Sulphate(IV) (SO_3^{2-}) salts

1. Sulphate (VI) (SO_4^{2-}) salts are normal and acid salts derived from Sulphuric (VI)acid H_2SO_4.
2. Sulphate(IV) (SO_3^{2-}) salts are normal and acid salts derived from Sulphuric (IV)acid H_2SO_3.
3. Sulphuric (VI)acid H_2SO_4 is formed when sulphur(VI)oxide gas is bubbled in water.

The acid exist as a dibasic acid with two ionisable hydrogen. It forms therefore the Sulphate (VI) (SO_4^{2-}) and hydrogen sulphate (VI) (HSO_4^-) salts.

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i.e.

H_2SO_4 (aq) -> $2H^+$(aq) + SO_4^{2-}(aq)

H_2SO_4 (aq) -> H^+(aq) + HSO_4^-(aq)

All Sulphate (VI) (SO_4^{2-}) salts **dissolve** in water/are soluble except Calcium (II) sulphate (VI) ($CaSO_4$), Barium (II) sulphate (VI) ($BaSO_4$) and Lead (II) sulphate (VI) ($PbSO_4$)

All Hydrogen sulphate (VI) (HSO_3^-) salts **exist** in solution/dissolved in water.

Sodium (I) hydrogen sulphate (VI) ($NaHSO_4$), Potassium (I) hydrogen sulphate (VI) ($KHSO_4$) and Ammonium hydrogen sulphate (VI) (NH_4HSO_4) exist also as solids.

Other Hydrogen sulphate (VI) (HSO_4^-) salts do not **exist** except those of Calcium (II) hydrogen sulphate (VI) ($Ca(HSO_4)_2$) and Magnesium (II) hydrogen sulphate (VI) ($Mg(HSO_4)_2$).

4. Sulphuric (IV)acid H_2SO_3 is formed when sulphur(IV)oxide gas is bubbled in water.

The acid exist as a dibasic acid with two ionisable hydrogen. It forms therefore the Sulphate (IV) (SO_3^{2-}) and hydrogen sulphate (VI) (HSO_4^-) salts.

i.e.

H_2SO_3 (aq) -> $2H^+$(aq) + SO_3^{2-}(aq)

H_2SO_3 (aq) -> H^+(aq) + HSO_3^-(aq)

All Sulphate (IV) (SO_3^{2-}) salts **dissolve** in water/are soluble except Calcium (II) sulphate (IV) ($CaSO_3$), Barium (II) sulphate (IV) ($BaSO_3$) and Lead (II) sulphate (IV) ($PbSO_3$)

All Hydrogen sulphate (IV) (HSO_3^-) salts **exist** in solution/dissolved in water.

Sodium (I) hydrogen sulphate (IV) ($NaHSO_3$), Potassium (I) hydrogen sulphate (IV) ($KHSO_3$) and Ammonium hydrogen sulphate (IV) (NH_4HSO_3) exist also as solids.

Other Hydrogen sulphate (IV) (HSO_3^-) salts do not **exist** except those of Calcium (II) hydrogen sulphate (IV) ($Ca(HSO_3)_2$) and Magnesium (II) hydrogen sulphate (IV) ($Mg(HSO_3)_2$).

5.The following experiments show the effect of heat on sulphate(VI) (SO_4^{2-})and sulphate(IV) (SO_3^{2-}) salts:

Experiment:

In a clean dry test tube place separately about 1.0g of :

Zinc(II)sulphate (VI), Iron(II)sulphate(VI), Copper(II)sulphate(VI),Sodium (I) sulphate (VI), Sodium (I) sulphate (IV).Heat gently then strongly. Test any gases produced using litmus papers.

Observations:

-Colourless droplets of liquid forms on the cooler parts of the test tube in all cases.

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-White solid residue is left in case of Zinc (II)sulphate(VI),Sodium (I) sulphate (VI) and Sodium (I) sulphate (IV).
-Colour changes from green to brown /yellow in case of Iron (II)sulphate(VI)
-Colour changes from blue to white then black in case of Copper (II) sulphate (VI)
-Blue litmus paper remain and blue and red litmus paper remain red in case of Zinc(II)sulphate(VI), Sodium (I) sulphate (VI) and Sodium (I) sulphate (IV)
-Blue litmus paper turns red and red litmus paper remain red in case of Iron (II)sulphate(VI) and Copper (II) sulphate (VI).
Explanation
(i)All Sulphate (VI) (SO_4^{2-}) salts exist as **hydrated** salts with water of crystallization that condenses and collects on cooler parts of test tube as a colourless liquid on gentle heating. e.g.

$K_2SO_4.10H_2O(s) \rightarrow K_2SO_4(s) + 10H_2O(l)$
$Na_2SO_4.10H_2O(s) \rightarrow Na_2SO_4(s) + 10H_2O(l)$
$MgSO_4.7H_2O(s) \rightarrow MgSO_4(s) + 7H_2O(l)$
$CaSO_4.7H_2O(s) \rightarrow CaSO_4(s) + 7H_2O(l)$
$ZnSO_4.7H_2O(s) \rightarrow ZnSO_4(s) + 7H_2O(l)$
$FeSO_4.7H_2O(s) \rightarrow FeSO_4(s) + 7H_2O(l)$
$Al_2(SO_4)_3.6H_2O(s) \rightarrow Al_2(SO_4)_3(s) + 6H_2O(l)$
$CuSO_4.5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(l)$

All Sulphate (VI) (SO_4^{2-}) salts do not decompose on heating **except** Iron (II) sulphate (VI) and Copper (II) sulphate (VI).
(i)Iron (II) sulphate (VI) decomposes on strong heating to produce acidic sulphur (IV)oxide and sulphur(VI)oxide gases. Iron(III)oxide is formed as a brown /yellow residue.
Chemical equation

$2FeSO_4(s) \rightarrow Fe_2O_3(s) + SO_2(g) + SO_3(g)$

This reaction is used for the school laboratory preparation of small amount of sulphur(VI)oxide gas.
Sulphur (VI) oxide readily /easily solidifies as white silky needles when the mixture is passed through freezing mixture/ice cold water.
Sulphur (IV) oxide does not.
(ii) Copper(II)sulphate(VI) decomposes on strong heating to black copper (II) oxide and Sulphur (VI) oxide gas.
Chemical equation

$2CuSO_4(s) \rightarrow CuO(s) + SO_3(g)$

This reaction is used for the school laboratory preparation of small amount of sulphur(VI)oxide gas.

6. The following experiments show the test for the presence of sulphate (VI) (SO_4^{2-})and sulphate(IV) (SO_3^{2-}) ions in a sample of a salt/compound:
Experiments/Observations:

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(a)Using Lead(II)nitrate(V)

I. To about 5cm3 of a salt solution in a test tube add four drops of Lead(II)nitrate(V)solution. Preserve.

Observation	Inference
White precipitate/ppt	SO_4^{2-} , SO_3^{2-} , CO_3^{2-} , Cl^- ions

II. To the preserved sample in (I) above, add six drops of 2M nitric(V) acid . Preserve.

Observation 1

Observation	Inference
White precipitate/ppt persists	SO_4^{2-} , Cl^- ions

Observation 2

Observation	Inference
White precipitate/ppt dissolves	SO_3^{2-} , CO_3^{2-} , ions

III.(a)To the preserved sample observation 1 in (II) above, Heat to boil.

Observation 1

Observation	Inference
White precipitate/ppt persists on boiling	SO_4^{2-} ions

Observation 2

Observation	Inference
White precipitate/ppt dissolves on boiling	Cl^- ions

.(b)To the preserved sample observation 2 in (II) above, add 4 drops of acidified potassium manganate(VII) /dichromate(VI).

Observation 1

Observation	Inference
(i)acidified potassium manganate(VII)decolorized (ii)Orange colour of acidified potassium dichromate(VI) turns to green	SO_3^{2-} ions

Observation 2

Observation	Inference
(i)acidified potassium manganate(VII) not decolorized (ii)Orange colour of acidified potassium dichromate(VI) does not turns to green	CO_3^{2-} ions

Experiments/Observations:

(b)Using Barium(II)nitrate(V)/ Barium(II)chloride

I. To about 5cm3 of a salt solution in a test tube add four drops of Barium(II) nitrate (V) / Barium(II)chloride. Preserve.

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Observation	Inference
White precipitate/ppt	SO_4^{2-} , SO_3^{2-} , CO_3^{2-} ions

II. To the preserved sample in (I) above, add six drops of 2M nitric(V) acid . Preserve.

Observation 1

Observation	Inference
White precipitate/ppt persists	SO_4^{2-} , ions

Observation 2

Observation	Inference
White precipitate/ppt dissolves	SO_3^{2-} , CO_3^{2-} , ions

III.To the preserved sample observation 2 in (II) above, add 4 drops of acidified potassium manganate(VII) /dichromate(VI).

Observation 1

Observation	Inference
(i)acidified potassium manganate(VII)decolorized (ii)Orange colour of acidified potassium dichromate(VI) turns to green	SO_3^{2-} ions

Observation 2

Observation	Inference
(i)acidified potassium manganate(VII) not decolorized (ii)Orange colour of acidified potassium dichromate(VI) does not turns to green	CO_3^{2-} ions

Explanations

Using Lead(II)nitrate(V)

(i)Lead(II)nitrate(V) solution reacts with chlorides(Cl^-), Sulphate (VI) salts (SO_4^{2-}), Sulphate (IV)salts (SO_3^{2-}) and carbonates(CO_3^{2-}) to form the insoluble white precipitate of Lead(II)chloride, Lead(II)sulphate(VI), Lead(II) sulphate (IV) and Lead(II)carbonate(IV).

Chemical/ionic equation:

$Pb^{2+}(aq) + Cl^-(aq) \rightarrow PbCl_2(s)$

$Pb^{2+}(aq) + SO_4^{2+}(aq) \rightarrow PbSO_4(s)$

$Pb^{2+}(aq) + SO_3^{2+}(aq) \rightarrow PbSO_3(s)$

$Pb^{2+}(aq) + CO_3^{2+}(aq) \rightarrow PbCO_3(s)$

(ii)When the insoluble precipitates are acidified with nitric(V) acid,

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- Lead(II)chloride and Lead(II)sulphate(VI) do not react with the acid and thus their white precipitates remain/ persists.
- Lead(II) sulphate (IV) and Lead(II)carbonate(IV) reacts with the acid to form **soluble** Lead(II) nitrate (V) and produce/effervesces/fizzes/bubbles out **sulphur(IV)oxide** and **carbon(IV)oxide** gases respectively.

. Chemical/ionic equation:

$PbSO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + Pb^{2+}(aq) + SO_2(g)$

$PbCO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + Pb^{2+}(aq) + CO_2(g)$

(iii)When Lead(II)chloride and Lead(II)sulphate(VI) are heated/warmed;
- Lead(II)chloride dissolves in hot water/on boiling(recrystallizes on cooling)
- Lead(II)sulphate(VI) do not dissolve in hot water thus its white precipitate persists/remains on heating/boiling.

(iv)When sulphur(IV)oxide and carbon(IV)oxide gases are produced;
- **sulphur(IV)oxide** will decolorize acidified potassium manganate(VII) and / or Orange colour of acidified potassium dichromate(VI) will turns to green.

Carbon(IV)oxide will not.

Chemical equation:

$5SO_3^{2-}(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(l)$
(purple) (colourless)

$3SO_3^{2-}(aq) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \rightarrow 3SO_4^{2-}(aq) + 2Cr^{3+}(aq) + 4H_2O(l)$
(Orange) (green)

- **Carbon(IV)oxide** forms an insoluble white precipitate of calcium carbonate if three drops of lime water are added into the reaction test tube when effervescence is taking place. **Sulphur(IV)oxide will not.**

Chemical equation:

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

These tests should be done immediately after acidifying to ensure the gases produced react with the oxidizing agents/lime water.

Using Barium(II)nitrate(V)/ Barium(II)Chloride

(i)Barium(II)nitrate(V) and/ or Barium(II)chlo**ride** solution reacts with Sulphate (VI) salts (SO_4^{2-}), Sulphate (IV)salts (SO_3^{2-}) and carbonates(CO_3^{2-}) to form the insoluble white precipitate of Barium(II)sulphate(VI), Barium(II) sulphate (IV) and Barium(II)carbonate(IV).

Chemical/ionic equation:

$Ba^{2+}(aq) + SO_4^{2+}(aq) \rightarrow BaSO_4(s)$

$Ba^{2+}(aq) + SO_3^{2+}(aq) \rightarrow BaSO_3(s)$

$Ba^{2+}(aq) + CO_3^{2+}(aq) \rightarrow BaCO_3(s)$

(ii)When the insoluble precipitates are acidified with nitric(V) acid,

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- Barium (II)sulphate(VI) do not react with the acid and thus its white precipitates remain/ persists.
- Barium(II) sulphate (IV) and Barium(II)carbonate(IV) reacts with the acid to form **soluble** Barium(II) nitrate (V) and produce /effervesces /fizzes/ bubbles out **sulphur(IV)oxide** and **carbon(IV)oxide** gases respectively.

. Chemical/ionic equation:

$BaSO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + Ba^{2+}(aq) + SO_2(g)$

$BaCO_3(s) + 2H^+(aq) \rightarrow H_2O(l) + Ba^{2+}(aq) + CO_2(g)$

(iii) When sulphur(IV)oxide and carbon(IV)oxide gases are produced;
- **sulphur(IV)oxide** will decolorize acidified potassium manganate(VII) and / or Orange colour of acidified potassium dichromate(VI) will turns to green.
Carbon(IV)oxide will not.
Chemical equation:

$5SO_3^{2-}(aq) + 2MnO_4^-(aq) + 6H^+(aq) \rightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 3H_2O(l)$
(purple) (colourless)

$3SO_3^{2-}(aq) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \rightarrow 3SO_4^{2-}(aq) + 2Cr^{3+}(aq) + 4H_2O(l)$
(Orange) (green)

- **Carbon(IV)oxide** forms an insoluble white precipitate of calcium carbonate if three drops of lime water are added into the reaction test tube when effervescence is taking place. **Sulphur(IV)oxide will not.**
Chemical equation:

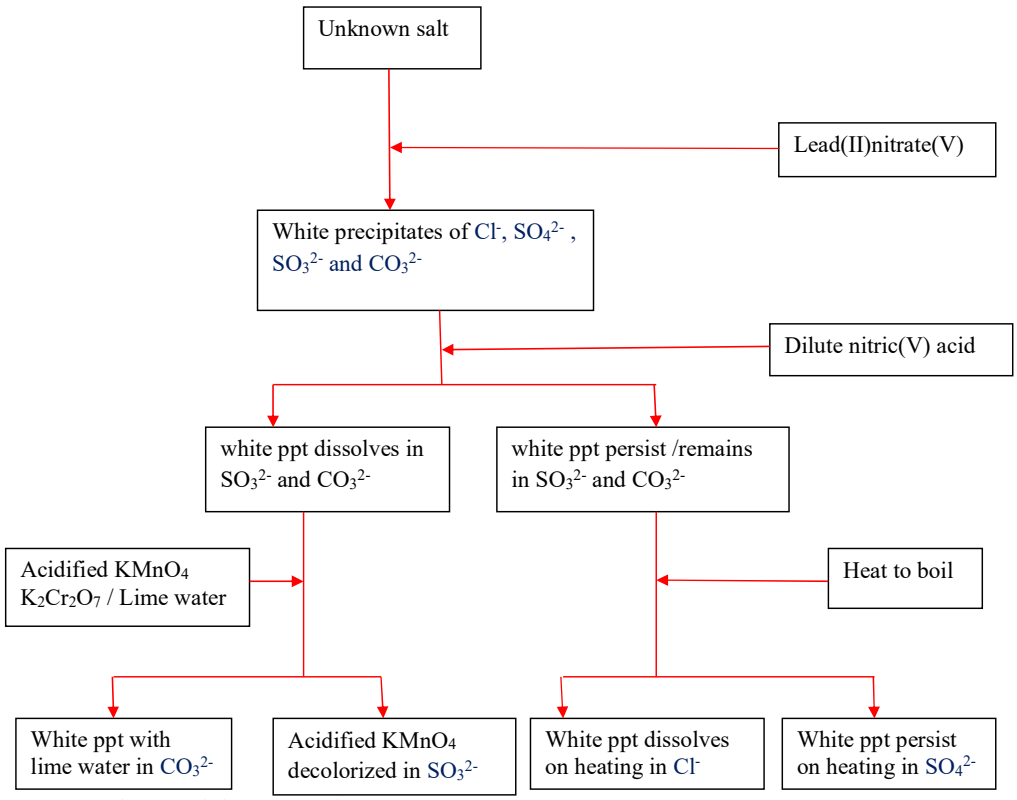
$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

These tests should be done immediately after acidifying to ensure the gases produced react with the oxidizing agents/lime water.

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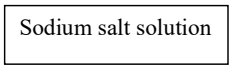
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Summary test for Sulphate (VI) (SO_4^{2-})and Sulphate(IV) (SO_3^{2-}) salts



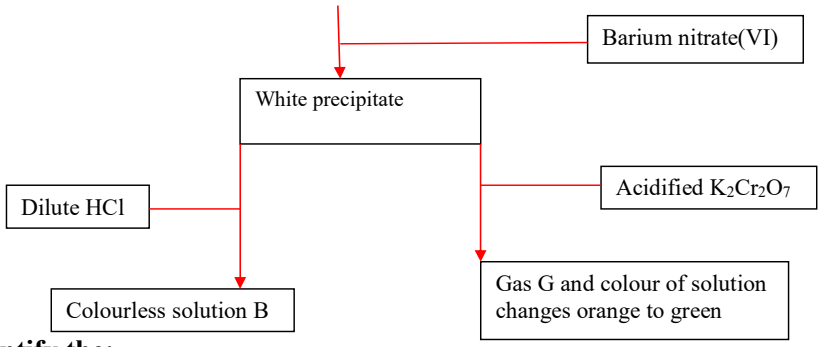
Practice revision question

1. Study the flow chart below and use it to answer the questions that follow



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(a)**Identify the:**

I: Sodium salt solution

Sodium sulphate(IV)/Na_2SO_3

II: White precipitate

Barium sulphate(IV)/$BaSO_3$

III: Gas G

Sulphur (IV)Oxide /SO_2

IV: Colourless solution H

Barium chloride /$BaCl_2$

(b)**Write an ionic equation for the formation of:**

I.White precipitate

Ionic equation $Ba^{2+}(aq) + SO_3^{2-}(aq) \rightarrow BaSO_3(s)$

II.Gas G

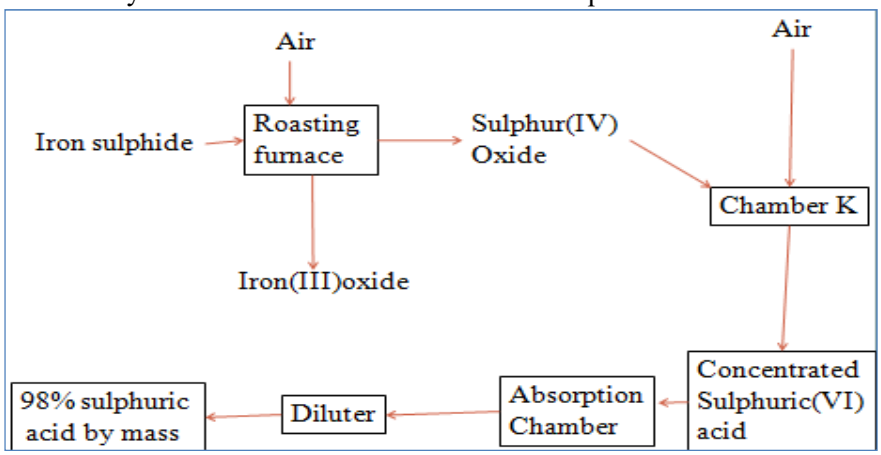
Ionic equation $BaSO_3(s) + 2H^+(aq) \rightarrow SO_2(g) + H_2O(l) + Ba^{2+}(aq)$

III. Green solution from the orange solution

$3SO_3^{2-}(aq) + Cr_2O_7^{2-}(aq) + 8H^+(aq) \rightarrow 3SO_4^{2-}(aq) + 2Cr^{3+}(aq) + 4H_2O(l)$

(Orange) (green)

2. Study the flow chart below and answer the questions that follow.



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(i)Write equation for the reaction taking place at:

I.The roasting furnace (1mk)

$2FeS_2$ (s) + $5O_2$ (g) -> $2FeO$(s) + $4SO_2$ (g)

II.The absorption tower (1mk)

H_2SO_4 (l) + SO_3 (g) -> $H_2S_2O_7$(l)

III.The diluter (1mk)

$H_2S_2O_7$(l) + H_2O(l) -> $2H_2SO_4$ (l)

(ii)The reaction taking place in chamber K is

$SO_{2\,(g)} + {}^1/_2O_{2\,(g)} \rightleftharpoons SO_{3\,(g)}$

I. Explain why it is necessary to use excess air in chamber K

To ensure all the SO_2 reacts

II.Name another substance used in chamber K

Vanadium(V)oxide

3.(a)Describe a chemical test that can be used to differentiate between sodium sulphate (IV) and sodium sulphate (VI).

Add acidified Barium nitrate(V)/chloride.

White precipitate formed with sodium sulphate (VI)

No white precipitate formed with sodium sulphate (IV)

(b)Calculate the volume of sulphur (IV) oxide formed when 120 kg of copper is reacted with excess concentrated sulphuric(VI)acid.(Cu = 63.5 ,1 mole of a gas at s.t.p =22.4dm3)

Chemical equation

$Cu(s) + 2H_2SO_4(l) \rightarrow CuSO_4(aq) + H_2O(l) + SO_2(g)$

Mole ratio Cu(s: SO_2 (g) = 1:1

Method 1

1 Mole Cu =63.5 g -> 1 mole SO_2 = 22.4dm3

(120 x 1000) g -> (120 x 1000) g x 22.4.dm3)

63.5 g

= **42330.7087**

Method 2

Moles of Cu = $\frac{(120 \times 1000)\ g}{63.5}$ = **1889.7639** moles

Moles SO_2 = Moles of Cu = **1889.7639** moles

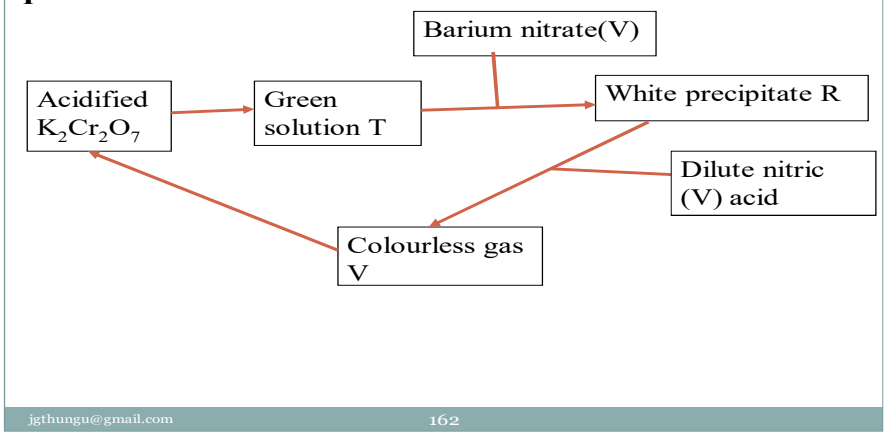
Volume of SO_2 = Mole x molar gas volume = (1889.7639 moles x 22.4)

= **42330.7114**

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4.Use the reaction scheme below to answer the questions that follow.



(a)Identify the:

(i)cation responsible for the green solution T

Cr^{3+}

(ii)possible anions present in white precipitate R

CO_3^{2-}, SO_3^{2-}, SO_4^{2-}

(b)Name gas V

Sulphur (IV)oxide

(c)Write a possible ionic equation for the formation of white precipitate R.

$Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$

$Ba^{2+}(aq) + SO_3^{2-}(aq) \rightarrow BaSO_3(s)$

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

ORGANIC CHEMISTRY 1 & 2

Introduction to Organic chemistry

Organic chemistry is the branch of chemistry that studies carbon compounds present in living things, once living things or synthetic/man-made.

Compounds that makes up living things whether alive or dead mainly contain carbon. Carbon is tetravalent.

It is able to form stable covalent bonds with itself and many non-metals like hydrogen, nitrogen ,oxygen and halogens to form a variety of compounds. This is because:

(i) carbon uses all the four valence electrons to form four strong covalent bond.

(ii)carbon can covalently bond to form a single, double or triple covalent bond with itself.

(iii)carbon atoms can covalently bond to form a very long chain or ring.

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When carbon covalently bond with Hydrogen, it forms a group of organic compounds called **Hydrocarbons**

A.HYDROCARBONS (HCs)

Hydrocarbons are a group of organic compounds containing /made up of hydrogen and carbon atoms only.

Depending on the type of bond that exist between the individual carbon atoms, hydrocarbon are classified as:

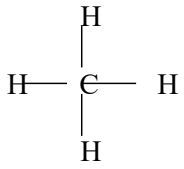
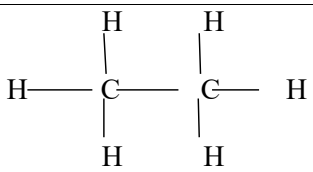
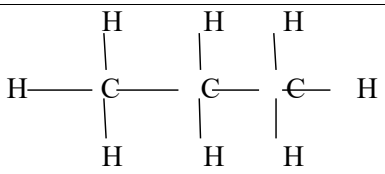
(i) Alk**a**nes
(ii) Alk**e**nes
(iii) Alk**y**nes

(i) Alkanes

(a)Nomenclature/Naming

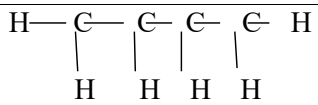
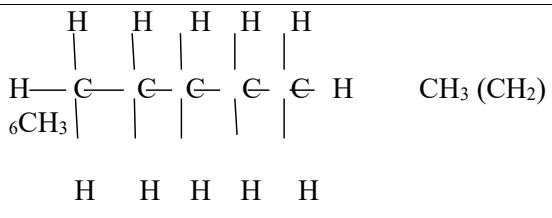
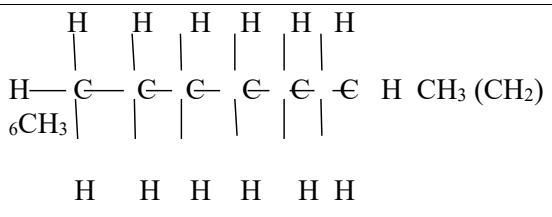
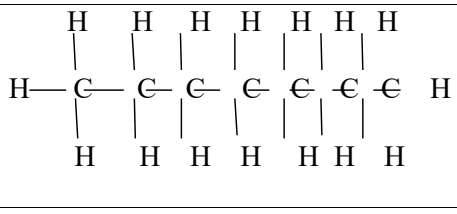
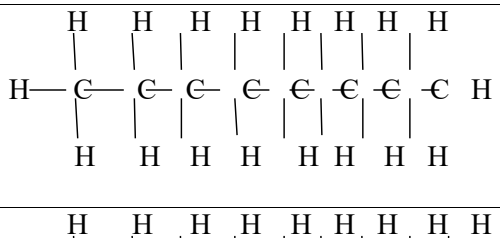
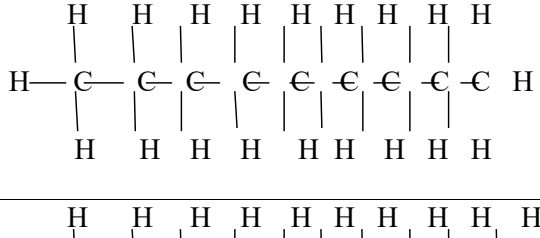
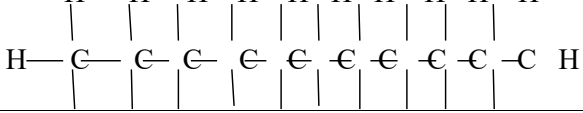
These are hydrocarbons with a general formula $\mathbf{C_nH_{2n+2}}$ where **n** is the number of Carbon atoms in a molecule.

The carbon atoms are linked by single bond to each other and to hydrogen atoms. They include:

n	General/ Molecular formula	Structural formula	Name
1	CH_4		Meth**ane**
2	C_2H_6		Eth**ane**
3	C_3H_8		Prop**ane**
4	C_4H_{10}		But**ane**

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5	C_5H_{12}		Pent**ane**
6	C_6H_{14}		Hex**ane**
7	C_7H_{16}		Hept**ane**
8	C_8H_{18}		Oct**ane**
9	C_9H_{20}		Non**ane**
10	$C_{10}H_{22}$		dec**ane**

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		H H H H H H H H H H

Note

1.The **general formula/molecular formular** of a compound shows the number of each atoms of elements making the compound e.g.

Decane has a general/molecular formula $\mathbf{C_{10}H_{22}}$;this means there are 10 carbon atoms and 22 hydrogen atoms in a molecule of decane.

2.The **structural formula** shows the arrangement/bonding of atoms of each element making the compound e.g

Decane has the structural formula as in the table above ;this means the 1st carbon from left to right is bonded to three hydrogen atoms and one carbon atom.

The 2nd carbon atom is joined/bonded to two other carbon atoms and two Hydrogen atoms.

3.Since carbon is **tetravalent** ,each atom of carbon in the alkane **MUST** always be bonded using **four** covalent bond /four shared pairs of electrons.

4.Since Hydrogen is **monovalent** ,each atom of hydrogen in the alkane **MUST** always be bonded using **one** covalent bond/one shared pair of electrons.

5.One member of the alkane differ from the next/previous by a CH_2 group.

e.g

Propane differ from ethane by one carbon and two Hydrogen atoms form ethane.

Ethane differ from methane also by one carbon and two Hydrogen atoms

6.A group of compounds that differ by a CH_2 group from the next /previous **consecutively** is called a **homologous series**.

7.A homologous series:

(i) differ by a CH_2 group from the next /previous consecutively

(ii)have similar chemical properties

(iii)have similar chemical formula that can be represented by a general formula e.g alkanes have the general formula C_nH_{2n+2}.

(iv)the physical properties (e.g.melting/boiling points)show steady gradual change)

8.The 1st four alkanes have the prefix **meth_,eth_,prop_** and **but_** to represent 1,2,3 and 4 carbons in the compound. All other use the numeral prefix **pent_,Hex_,hept_** , etc to show also the number of carbon atoms.

9.If one hydrogen atom in an alkane is removed, an alkyl group is formed.e.g

Alkane name	molecular structure C_nH_{2n+2}	**Alkyl name**	**Molecula structure C_nH_{2n+1}**
methane	CH_4	**methyl**	**CH_3**
ethane	CH_3CH_3	**ethyl**	**CH_3 CH_2**
propane	CH_3 CH_2 CH_3	**propyl**	**CH_3 CH_2 CH_2**

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butane	$CH_3\ CH_2\ CH_2\ CH_3$	**butyl**	**$CH_3\ CH_2\ CH_2\ CH_2$**

(b)Isomers of alkanes

Isomers are compounds with the same molecular **general formula** but different molecular **structural formula**.

Isomerism is the existence of a compounds having the same general/molecular formula but different structural formula.

The 1st three alkanes do not form isomers.Isomers are named by using the IUPAC(**I**nternational **U**nion of **P**ure and **A**pplied **C**hemistry) system of **nomenclature/naming.**

The IUPAC system of nomenclature uses the following basic rules/guidelines:

1.Identify the longest continuous carbon chain to get/determine the parent alkane.

2.Number the longest chain form the end of the chain that is near the branches so as the branch get the lowest number possible

3. Determine the position, number and type of branches. Name them as methyl, ethyl, propyl e.tc. according to the number of carbon chains attached to the parent alkane. Name them fluoro-,chloro-,bromo-,iodo- if they are halogens

4.Use prefix di-,tri-,tetra-,penta-,hexa- to show the number of branches attached to the parent alkane.

Practice on IUPAC nomenclature of alkanes

(a)Draw the structure of:

(i)2-methylpentane

Procedure

1. Identify the longest continuous carbon chain to get/determine the parent alkane.

Butane is the parent name **$CH_3\ CH_2\ CH_2\ CH_3$**

2. Number the longest chain form the end of the chain that is near the branches so as the branch get the lowest number possible

The methyl group is attached to Carbon **"2"**

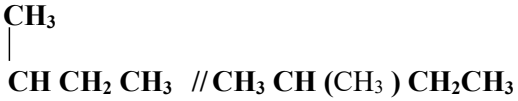
3. Determine the position, number and type of branches. Name them as methyl, ethyl, propyl e.tc. according to the number of carbon chains attached to the parent alkane i.e

Position of the branch at carbon "2"

Number of branches at carbon "1"

Type of the branch "methyl" hence

Molecular formula

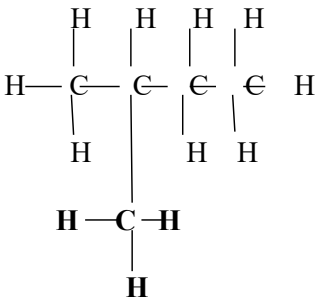


$CH_3\ CH\ CH_2\ CH_3$ // **$CH_3\ CH\ (CH_3)\ CH_2CH_3$**

Structural formula

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(ii)2,2-dimethylpentane

Procedure

1. Identify the longest continuous carbon chain to get/determine the parent alkane.

Butane is the parent name **$CH_3\ CH_2\ CH_2\ CH_3$**

2. Number the longest chain form the end of the chain that is near the branches so as the branch get the lowest number possible

The methyl group is attached to Carbon **"2"**

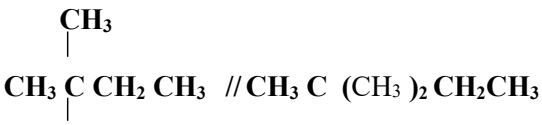
3. Determine the position, number and type of branches. Name them as methyl, ethyl, propyl e.tc. according to the number of carbon chains attached to the parent alkane i.e

Position of the branch at carbon "2"

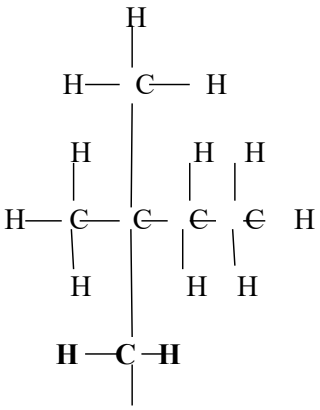
Number of branches at carbon "2"

Type of the branch two"methyl" hence

Molecular formular



Structural formula



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(iii) 2,2,3-trimethylbutane

Procedure

1. Identify the longest continuous carbon chain to get/determine the parent alkane.
   Butane is the parent name **$CH_3\ CH_2\ CH_2\ CH_3$**
2. Number the longest chain form the end of the chain that is near the branches so as the branch get the lowest number possible

The methyl group is attached to Carbon **"2 and 3"**

3. Determine the position, number and type of branches. Name them as methyl, ethyl, propyl e.tc. according to the number of carbon chains attached to the parent alkane i.e

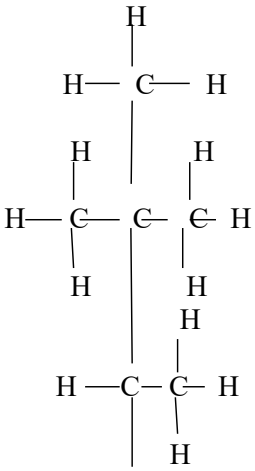
   Position of the branch at carbon "2 and 3"
   Number of branches at carbon "3"
   Type of the branch three "methyl" hence
   Molecular formular

CH_3

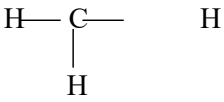
$CH_3\ C\ CH\ CH_3$ // $CH_3\ C\ (CH_3)_3\ CH_2CH_3$

$CH_3\ CH_3$

Structural formula



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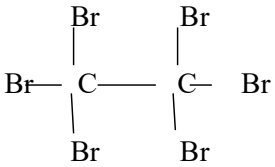


(iv) 1,1,1,2,2,2-hexabromoethane

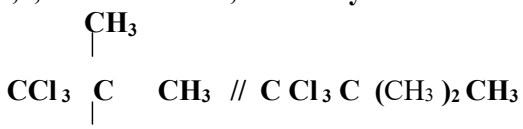
Molecular formula

CBr_3 CBr_3

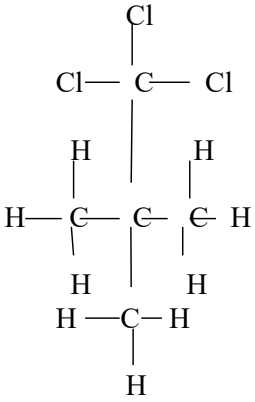
Structural formula



(v) 1,1,1-tetrachloro-2,2-dimethylbutane



Structural formula



(c)Occurrence and extraction

Crude oil ,**natural gas** and **biogas** are the main sources of alkanes:

(i)Natural gas is found on top of crude oil deposits and consists mainly of methane.

(ii)Biogas is formed from the decay of waste organic products like animal dung and cellulose. When the decay takes place in absence of oxygen , 60-75% by volume of the gaseous mixture of methane gas is produced.

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(iii)Crude oil is a mixture of many flammable hydrocarbons/substances. Using fractional distillation, each hydrocarbon fraction can be separated from the other. The hydrocarbon with lower /smaller number of carbon atoms in the chain have lower boiling point and thus collected first.

As the carbon **chain increase**, the **boiling** point, **viscosity** (ease of flow) and colour **intensity** increase as **flammability** decrease. Hydrocarbons in crude oil are not pure. They thus have no sharp fixed boiling point.

Uses of different crude oil fractions

Carbon atoms in a molecule	Common name of fraction	Uses of fraction
1-4	Gas	L.P.G gas for domestic use
5-12	Petrol	Fuel for petrol engines
9-16	Kerosene/Paraffin	Jet fuel and domestic lighting/cooking
15-18	Light diesel	Heavy diesel engine fuel
18-25	Diesel oil	Light diesel engine fuel
20-70	Lubricating oil	Lubricating oil to reduce friction.
Over 70	Bitumen/Asphalt	Tarmacking roads

(d)School laboratory preparation of alkanes

In a school laboratory, alkanes may be prepared from the reaction of a sodium alkanoate with solid sodium hydroxide/soda lime.

Chemical equation:

Sodium alkanoate + soda lime -> alkane + Sodium carbonate

$C_nH_{2n+1}COONa(s) + NaOH(s) \rightarrow C_nH_{2n+2} + Na_2CO_3(s)$

The "H" in NaO**H** is transferred/moves to the C_nH_{2n+1} in $C_nH_{2n+1}COONa(s)$ to form C_nH_{2n+2}.

Examples

1. **Meth**ane is prepared from the heating of a mixture of sodium **ethan**oate and soda lime and collecting over water

Sodium **ethan**oate + soda lime -> **meth**ane + Sodium carbonate

$CH_3COONa(s) + NaOH(s) \rightarrow CH_4 + Na_2CO_3(s)$

The "H" in NaO**H** is transferred/moves to the CH_3 in $CH_3COONa(s)$ to form CH_4.

2. **Eth**ane is prepared from the heating of a mixture of sodium **prop**anoate and soda lime and collecting over water

Sodium **prop**anoate + soda lime -> **eth**ane + Sodium carbonate

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$CH_3CH_2COONa(s)$ + NaO**H**(s) -> CH_3CH_3 + $Na_2CO_3(s)$
The "H" in NaO**H** is transferred/moves to the CH_3CH_2 in CH_3CH_2COONa (s) to form CH_3CH_3

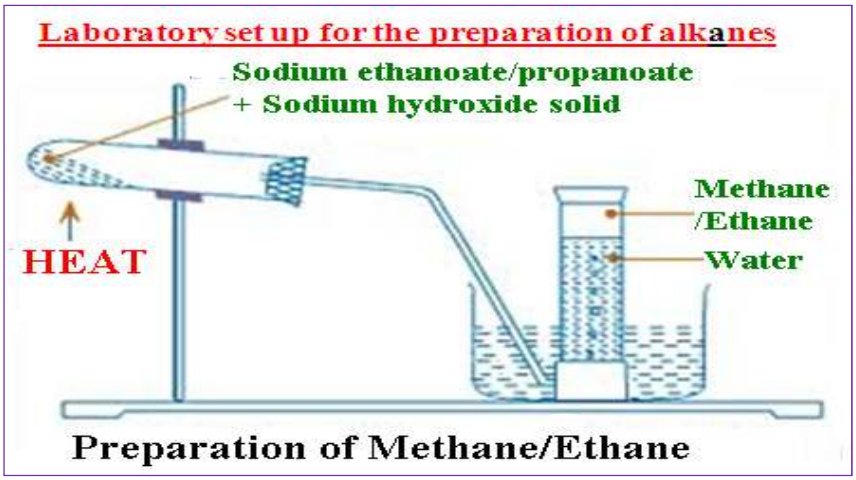
3. **Prop**ane is prepared from the heating of a mixture of sodium **but**anoate and soda lime and collecting over water

Sodium **but**anoate + soda lime -> **prop**ane + Sodium carbonate
$CH_3CH_2CH_2COONa(s)$ + NaO**H**(s) -> $CH_3CH_2CH_3$ + $Na_2CO_3(s)$
The "H" in NaO**H** is transferred/moves to the $CH_3CH_2CH_2$ in $CH_3CH_2CH_2COONa$ (s) to form $CH_3CH_2CH_3$

4. **But**ane is prepared from the heating of a mixture of sodium **pent**anoate and soda lime and collecting over water

Sodium **pent**anoate + soda lime -> **but**ane + Sodium carbonate
$CH_3CH_2CH_2CH_2COONa(s)$+NaO**H**(s) -> $CH_3CH_2CH_2CH_3$ + $Na_2CO_3(s)$
The "H" in NaO**H** is transferred/moves to the $CH_3CH_2CH_2CH_2$ in $CH_3CH_2CH_2CH_2COONa$ (s) to form $CH_3CH_2CH_2CH_3$

Laboratory set up for the preparation of alkanes



(d)Properties of alkanes

I. Physical properties

Alkanes are colourless gases, solids and liquids that are not poisonous.
They are slightly soluble in water.
The solubility decrease as the carbon chain and thus the molar mass increase
The melting and boiling point increase as the carbon chain increase.
This is because of the increase in van-der-waals /intermolecular forces as the carbon chain increase.

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The 1st four straight chain alkanes (methane,ethane,propane and butane)are therefore gases ,the nect six(pentane ,hexane, heptane,octane,nonane, and decane) are liquids while the rest from unidecane(11 carbon atoms) are solids .
The density of straight chain alkanes increase with increasing carbon chain as the intermolecular forces increases.
This reduces the volume occupied by a given mass of the compound.

Summary of physical properties of alkanes

Alkane	General formula	Melting point(K)	Boiling point(K)	Density gcm^{-3}	State at room(298K) temperature and pressure atmosphere (101300Pa)
Methane	CH4	90	112	0.424	gas
Ethane	CH_3CH_3	91	184	0.546	gas
Propane	$CH_3CH_2CH_3$	105	231	0.501	gas
Butane	$CH_3(CH_2)_2CH_3$	138	275	0.579	gas
Pentane	$CH_3(CH_2)_3CH_3$	143	309	0.626	liquid
Hexane	$CH_3(CH_2)_4CH_3$	178	342	0.657	liquid
Heptane	$CH_3(CH_2)_5CH_3$	182	372	0.684	liquid
Octane	$CH_3(CH_2)_6CH_3$	216	399	0.703	liquid
Nonane	$CH_3(CH_2)_7CH_3$	219	424	0.708	liquid
Octane	$CH_3(CH_2)_8CH_3$	243	447	0.730	liquid

II.Chemical properties

(i)Burning.

Alkanes burn with a **blue**/non-luminous **non-sooty**/non-smoky flame in **excess** air to form carbon(IV) oxide and water.

Alkane + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

Alkanes burn with a **blue**/non-luminous **no-sooty**/non-smoky flame in **limited** air to form carbon(II) oxide and water.

Alkane + Air -> carbon(**II**) oxide + water (limited air)

Examples

1.(a) Methane when ignited burns with a **blue non sooty** flame in **excess** air to form carbon(IV) oxide and water.

Methane + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l/g)$

(b) Methane when ignited burns with a **blue non sooty** flame in **limited** air to form carbon(II) oxide and water.

Methane + Air -> carbon(**II**) oxide + water (excess air/oxygen)

$2CH_4(g) + 3O_2(g) \rightarrow 2CO(g) + 4H_2O(l/g)$

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2.(a) Ethane when ignited burns with a **blue non sooty** flame in **excess** air to form carbon(IV) oxide and water.

Ethane + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l/g)$

(b) Ethane when ignited burns with a **blue non sooty** flame in **limited** air to form carbon(II) oxide and water.

Ethane + Air -> carbon(**II**) oxide + water (excess air/oxygen)

$2C_2H_6(g) + 5O_2(g) \rightarrow 4CO(g) + 6H_2O(l/g)$

3.(a) Propane when ignited burns with a **blue non sooty** flame in **excess** air to form carbon(IV) oxide and water.

Propane + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l/g)$

(b) Ethane when ignited burns with a **blue non sooty** flame in **limited** air to form carbon(II) oxide and water.

Ethane + Air -> carbon(**II**) oxide + water (excess air/oxygen)

$2C_3H_8(g) + 7O_2(g) \rightarrow 6CO(g) + 8H_2O(l/g)$

ii)Substitution

Substitution reaction is one in which a hydrogen atom is replaced by a halogen in presence of ultraviolet light.

Alkanes react with halogens in presence of ultraviolet light to form halogenoalkanes.

During substitution:

(i)the halogen molecule is split into free atom/radicals.

(ii)one free halogen radical/atoms knock /remove one hydrogen from the alkane leaving an alkyl radical.

(iii) the alkyl radical combine with the other free halogen atom/radical to form halogenoalkane.

(iv)the chlorine atoms substitute repeatedly in the alkane. Each substitution removes a hydrogen atom from the alkane and form hydrogen halide.

(v)substitution stops when all the hydrogen in alkanes are replaced with halogens.

Substitution reaction is a highly **explosive** reaction in presence of **sunlight / ultraviolet** light that act as **catalyst**.

Examples of substitution reactions

Methane has no effect on bromine or chlorine in diffused light/dark. In sunlight , a mixture of chlorine and methane explode to form colourless mixture of chloromethane and hydrogen chloride gas. The pale green colour of chlorine gas fades.

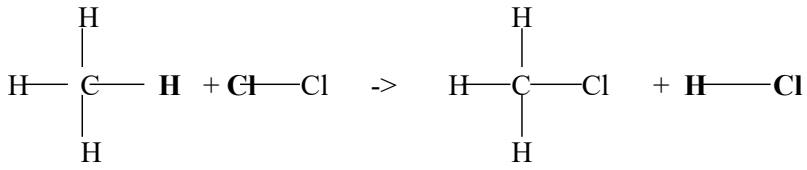
Chemical equation

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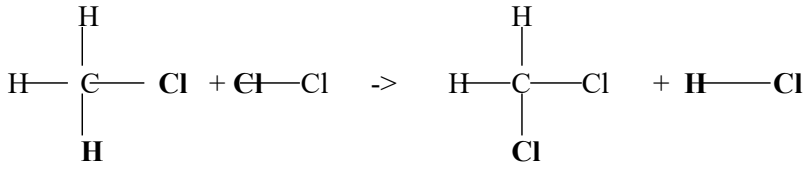
1.(a)Methane + chlorine -> Chloromethane + Hydrogen chloride

$CH_4(g) + Cl_2(g) \rightarrow CH_3Cl\ (g) + HCl\ (g)$



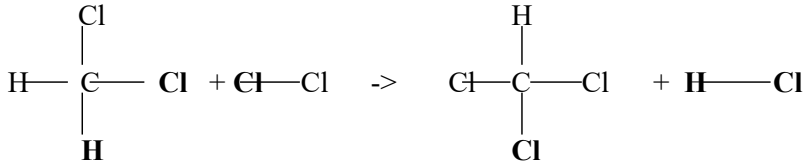
(b) Chloromethane + chlorine -> dichloromethane + Hydrogen chloride

$CH_3Cl\ (g) + Cl_2(g) \rightarrow CH_2Cl_2\ (g) + HCl\ (g)$



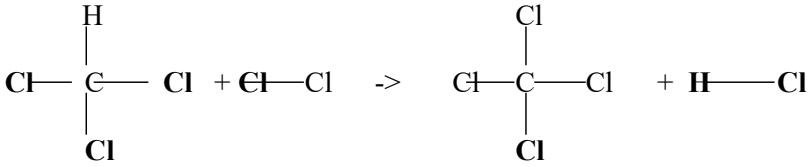
(c) dichloromethane + chlorine -> trichloromethane + Hydrogen chloride

$CH_2Cl_2\ (g) + Cl_2(g) \rightarrow CHCl_3\ (g) + HCl\ (g)$



(c) trichloromethane + chlorine -> tetrachloromethane + Hydrogen chloride

$CHCl_3\ (g) + Cl_2(g) \rightarrow CCl_4\ (g) + HCl\ (g)$



Ethane has no effect on bromine or chlorine in diffused light/dark. In sunlight , a mixture of bromine and ethane explode to form colourless mixture of bromoethane and hydrogen chloride gas. The red/brown colour of bromine gas fades.

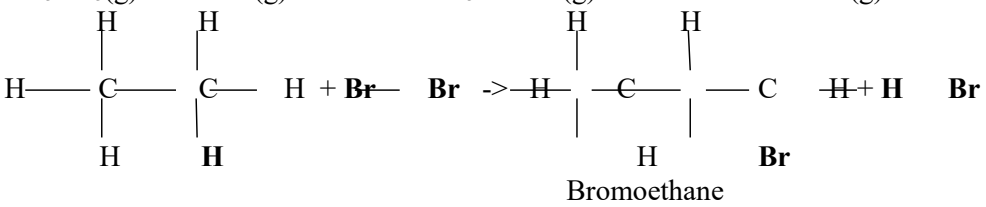
Chemical equation

(a)Ethane + chlorine -> Chloroethane + Hydrogen chloride

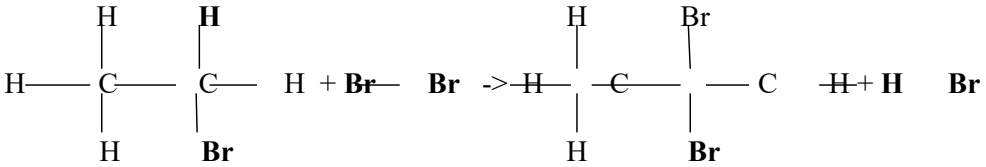
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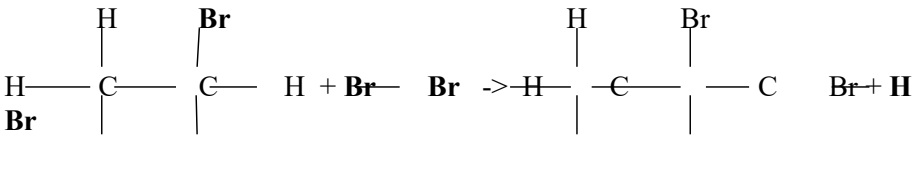
$CH_3CH_3(g) + Br_2(g) \rightarrow CH_3CH_2Br(g) + HBr(g)$



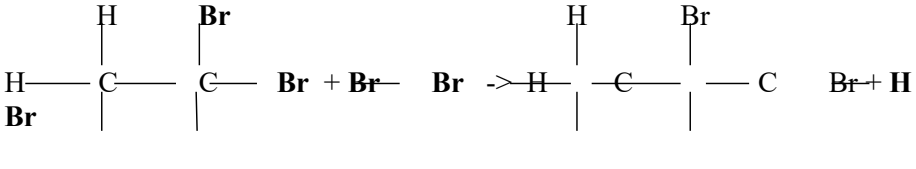
Bromoethane



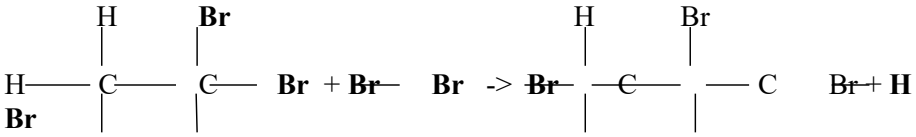
1,1-dibromoethane



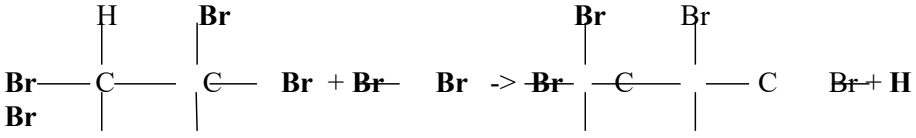
1,1,1-tribromoethane



1,1,1,2-tetrabromoethane



1,1,1,2,2-pentabromoethane



1,1,1,2,2,2-hexabromoethane

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Uses of alkanes

1.Most alkanes are used as fuel e.g. Methane is used as biogas in homes. Butane is used as the Laboratory gas.

2.On cracking ,alkanes are a major source of Hydrogen for the manufacture of ammonia/Haber process.

3.In manufacture of Carbon black which is a component in printers ink.

4.In manufacture of useful industrial chemicals like methanol, methanol, and chloromethane.

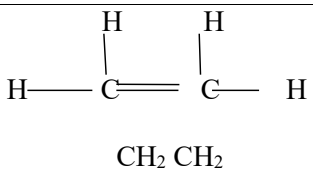
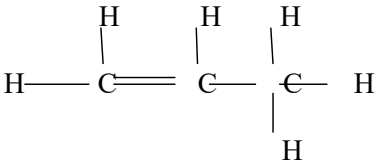
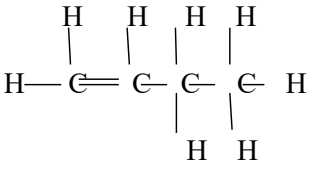
(ii) Alkenes

(a)Nomenclature/Naming

These are hydrocarbons with a general formula $\mathbf{C_nH_{2n}}$ and $=C-C$ double bond as the functional group . **n** is the number of Carbon atoms in the molecule.

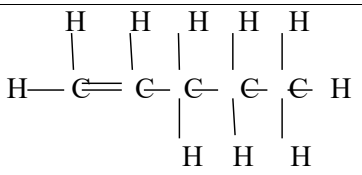
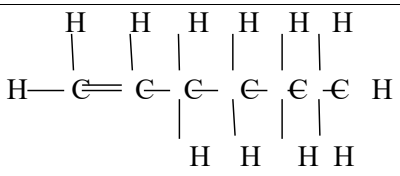
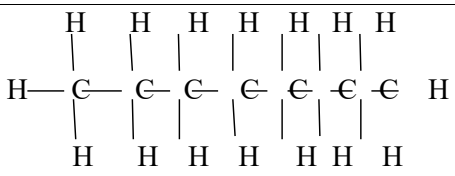
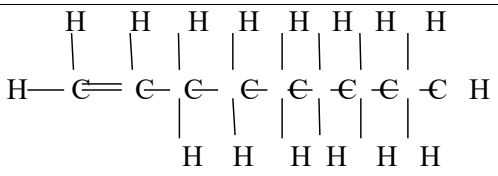
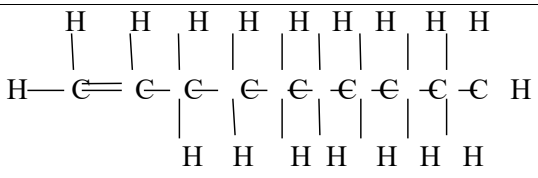
The carbon atoms are linked by at least one **double** bond to each other and single bonds to hydrogen atoms.

They include:

n	General/ Molecular formula	Structural formula	Name
1		Does not exist	
2	C_2H_6	 $CH_2\ CH_2$	Eth**ene**
3	C_3H_8	 $CH_2\ CH\ CH_3$	Prop**en e**
4	C_4H_{10}	 $CH_2\ CH\ CH_2CH_3$	But**ene**

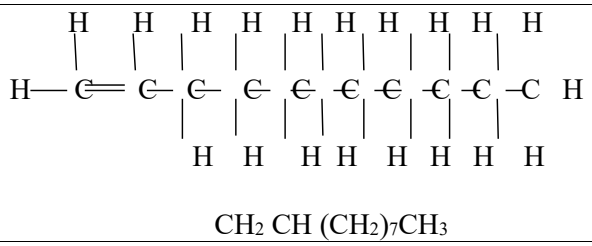
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5	C_5H_{12}		Pent**ene**
6	C_6H_{14}		Hex**ene**
7	C_7H_{16}		Hept**en e**
8	C_8H_{18}		Oct**ene**
9	C_9H_{20}		Non**ene**

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10	$C_{10}H_{22}$		dec**ene**

$CH_2\ CH\ (CH_2)_7CH_3$

Note

1.Since carbon is **tetravalent** ,each atom of carbon in the alkene **MUST** always be bonded using **four** covalent bond /four shared pairs of electrons including at the double bond.

2.Since Hydrogen is **monovalent** ,each atom of hydrogen in the alkene **MUST** always be bonded using **one** covalent bond/one shared pair of electrons.

3.One member of the alkene ,like alkanes,differ from the next/previous by a CH_2 group.They also form a homologous series.

e.g

Propene differ from ethene by one carbon and two Hydrogen atoms from ethene.

4.A homologous series of alkenes like that of alkanes:

(i) differ by a CH_2 group from the next /previous consecutively

(ii)have similar chemical properties

(iii)have similar chemical formula represented by the general formula C_nH_{2n}

(iv)the physical properties also show steady gradual change

5.The = C= C = double bond in alkene is the functional group. A functional group is the **reacting site** of a molecule/compound.

6. The = C= C = double bond in alkene can easily be broken to accommodate more two more monovalent atoms. The = C= C = double bond in alkenes make it thus **unsaturated.**

7. An unsaturated hydrocarbon is one with a double =C=C= or triple≡C C – carbon bonds in their molecular structure. Unsaturated hydrocarbon easily reacts to be **saturated**.

8.A saturated hydrocarbon is one without a double =C=C= or triple≡C C – carbon bonds in their molecular structure.

Most of the reactions of alkenes take place at the = C = C =bond.

(b)Isomers of alkenes

Isomers are alkenes lie alkanes have the same molecular **general formula** but different molecular **structural formula**.

Ethene and propene do not form isomers. Isomers of alkenes are also named by using the IUPAC(**I**nternational **U**nion of **P**ure and **A**pplied **C**hemistry) system of **nomenclature/naming.**

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The IUPAC system of nomenclature of naming alkenes uses the following basic rules/guidelines:

1.Identify the longest continuous/straight carbon chain which contains the **=C = C= double** bond get/determine the **parent** alkene.

2.Number the longest chain form the end of the chain which contains the **=C = C= double** bond so he **=C = C= double** bond lowest number possible.

3 Indicate the positions by splitting "**alk**-positions-**ene**" e.g. but-2-ene, pent-1,3-diene.

4.The position **indicated** must be for the carbon atom at the **lower** position in the **=C = C= double bond.**i.e

But-2-ene means the double =C = C= is between Carbon "2"and "3"

Pent-1,3-diene means there are two double bond one between carbon "1" and "2"and another between carbon "3" and "4"

5. Determine the position, number and type of branches. Name them as methyl, ethyl, propyl e.tc. according to the number of alkyl carbon chains attached to the alkene. Name them fluoro-,chloro-,bromo-,iodo- if they are halogens

6.Use prefix di-,tri-,tetra-,penta-,hexa- to show the number of **double** C = C bonds and **branches** attached to the alkene.

7.Position isomers can be formed when the=C = C= double bond is shifted between carbon atoms e.g.

But-2-ene means the double =C = C= is between Carbon "2"and "3"

But-1-ene means the double =C = C= is between Carbon "1"and "2"

Both But-1-ene and But-2-ene are position isomers of Butene

8.Position isomers are molecules/compounds having the same general formular but different position of the functional group.i.e.

Butene has the molecular/general formular C_4H_8 position but can form both But-1-ene and But-2-ene as position isomers.

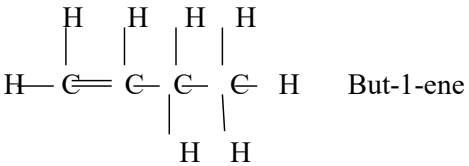
9. Like alkanes ,an alkyl group can be attached to the alkene. Chain/branch isomers are thus formed.

10.Chain/branch isomers are molecules/compounds having the same general formula but different structural formula e.g

Butene and 2-methyl propene both have the same general formualr but different branching chain.

Practice on IUPAC nomenclature of alkenes

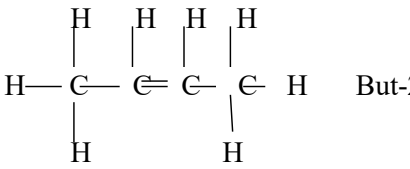
Name the following isomers of alkene



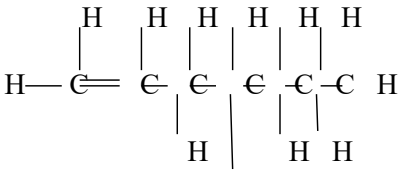
But-1-ene

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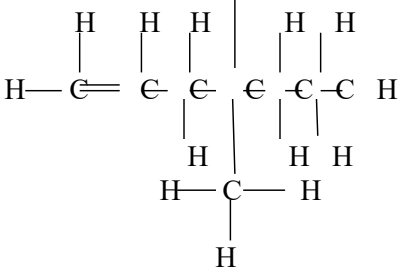
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But-2-ene

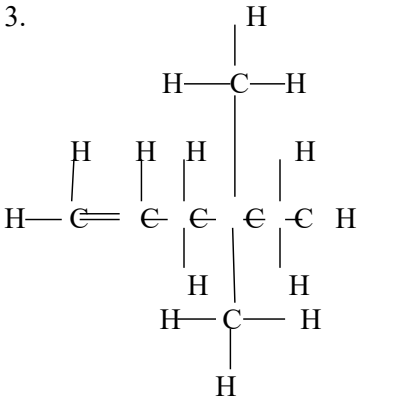


4-methylhex-1-ene



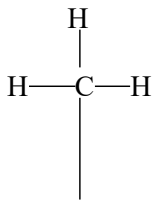
4,4-dimethylhex-1-ene

3.



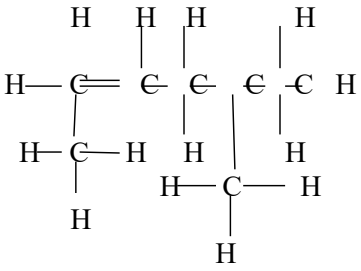
4,4-dimethylpent -1- ene

4.



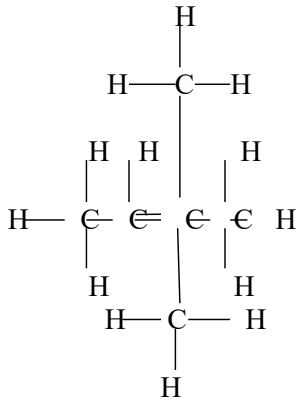
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5,5-dimethylhex-2- ene

5.



2,2-dimethylbut -2- ene

8.H_2C $CHCH_2$ CH_2 CH_3 pent -1- ene
9.H_2C $C(CH_3)CH_2$ CH_2 CH_3 2-methylpent -1- ene
10.H_2C $C(CH_3)C(CH_3)_2$ CH_2 CH_3 2,3,3-trimethylpent -1- ene
11.H_2C $C(CH_3)C(CH_3)_2$ $C(CH_3)_2$ CH_3 2,3,3,4,4-pentamethylpent -1- ene
12.H_3C $C(CH_3)C(CH_3)$ $C(CH_3)_2$ CH_3 2,3,4,4-tetramethylpent -2- ene
13. H_2C $C(CH_3)C(CH_3)$ $C(CH_3)$ CH_3 2,3,4-trimethylpent -1,3- diene
14. H_2C CBrCBr CBr CH_3 2,3,4-tribromopent -1,3- diene
15. H_2C $CHCH$ CH_2 But -1,3- diene
16. Br_2C CBrCBr CBr_2 1,1,2,3,4,4-hexabromobut -1,3- diene
17. I_2C CICI CI_2 1,1,2,3,4,4-hexaiodobut -1,3- diene
18. H_2C $C(CH_3)C(CH_3)$ CH_2 2,3-dimethylbut -1,3- diene

(c)Occurrence and extraction

At indusrial level,alkenes are obtained from the cracking of alk**a**nes.Cracking is the process of breaking long chain alkanes to smaller/shorter alk**a**nes, an alk**e**ne and hydrogen gas at high temperatures.

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Cracking is a major source of useful hydrogen gas for manufacture of ammonia/nitric(V)acid/HCl i.e.

Long chain alkane -> smaller/shorter alk**a**ne + Alk**e**ne + Hydrogen gas

Examples

1.When irradiated with high energy radiation,Propane undergo cracking to form methane gas, ethene and hydrogen gas.

Chemical equation

$CH_3CH_2CH_3$ (g) -> CH_4(g) + $CH_2=CH_2$(g) + H_2(g)

2.Octane undergo cracking to form hydrogen gas, butene and butane gases

Chemical equation

$CH_3(CH_2)_6CH_3$ (g) -> $CH_3CH_2CH_2CH_3$(g) + $CH_3CH_2CH=CH_2$(g) + H_2(g)

(d)School laboratory preparation of alke**nes**

In a school laboratory, alk**e**nes may be prepared from dehydration of alkanols using:

(i) concentrated sulphuric(VI)acid(H_2SO_4).

(a) aluminium(III)oxide(Al_2O_3) i.e

Alkanol --Conc. H_2SO_4 --> Alkene + Water

Alkanol --Al_2O_3 --> Alkene + Water e.g.

1.(a)At about 180ºC,concentrated sulphuric(VI)acid dehydrates/removes water from ethanol to form eth**e**ne.

The gas produced contain traces of carbon(IV)oxide and sulphur(IV)oxide gas as impurities.

It is thus passed through concentrated sodium/potassium hydroxide solution to remove the impurities.

Chemical equation

CH_3CH_2**OH** (l) --conc H_2SO_4/180ºC--> $CH_2=CH_2$(g) + **H_2O**(l)

(b)On heating strongly aluminium(III)oxide(Al_2O_3),it dehydrates/removes water from ethanol to form eth**e**ne.

Ethanol vapour passes through the hot aluminium (III) oxide which catalyses the dehydration.

Activated aluminium(III)oxide has a very high affinity for water molecules/elements of water and thus dehydrates/ removes water from ethanol to form eth**e**ne.

Chemical equation

CH_3CH_2**OH** (l) --(Al_2O_3/strong heat--> $CH_2=CH_2$(g) + **H_2O**(l)

2(a) Propan-1-ol and Propan-2-ol(position isomers of propanol) are dehydrated by conc H_2SO_4 at about 180ºC to prop**e**ne(propene has no position isomers).

Chemical equation

$CH_3CH_2CH_2$**OH** (l) -- conc H_2SO_4/180ºC --> $CH_3CH_2=CH_2$(g) + **H_2O**(l)

Propan-**1**-ol Prop-**1**-ene

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CH_3CH**OH** CH_3 (l) -- conc $H_2SO_4/180^oC$ --> $CH_3CH_2=CH_2(g)$ + **H_2O**(l)
Propan-**2**-ol Prop-**1**-ene

(b) Propan-1-ol and Propan-2-ol(position isomers of propanol) are dehydrated by heating strongly aluminium(III)oxide(Al_2O_3) form prop**e**ne

Chemical equation

CH_3CH_2 CH_2**OH** (l) -- Heat/Al_2O_3 --> $CH_3CH_2=CH_2(g)$ + **H_2O**(l)
Propan-**1**-ol Prop-**1**-ene

CH_3CH**OH** CH_3 (l) -- Heat/Al_2O_3 --> $CH_3CH_2=CH_2(g)$ + **H_2O**(l)
Propan-**2**-ol Prop-**1**-ene

3(a) Butan-1-ol and Butan-2-ol(position isomers of butanol) are dehydrated by conc H_2SO_4 at about 180°C to But-1-**e**ne and But-2-**e**ne respectively

Chemical equation

CH_3CH_2 CH_2 CH_2**OH** (l) -- conc $H_2SO_4/180^oC$ -->CH_3 $CH_2CH_2=CH_2(g)$ + **H_2O**(l)
Butan-**1**-ol But-**1**-ene

CH_3CH**OH** CH_2CH_3 (l)-- conc $H_2SO_4/180^oC$ -->$CH_3CH=CH$ $CH_2(g)$ + **H_2O**(l)
Butan-**2**-ol But-**2**-ene

(b) Butan-1-ol and Butan-2-ol are dehydrated by heating strongly aluminium (III) oxide (Al_2O_3) form But-1-**e**ne and But-2-**e**ne respectively.

Chemical equation

CH_3CH_2 CH_2 CH_2**OH** (l) -- Heat/Al_2O_3 --> CH_3 $CH_2CH_2=CH_2(g)$ + **H_2O**(l)
Butan-**1**-ol But-**1**-ene

CH_3CH**OH** CH_2CH_3 (l) -- Heat/Al_2O_3 --> $CH_3CH=CH$ $CH_2(g)$ + **H_2O**(l)
Butan-**2**-ol But-**2**-ene

Laboratory set up for the preparation of alk**e**nes/**ethene**

Caution

(i) Ethanol is highly inflammable

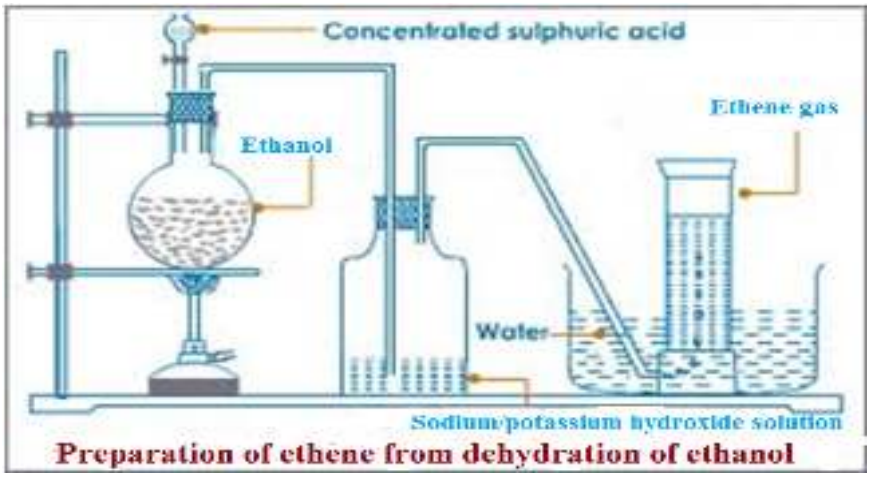
(ii) Conc H_2SO_4 is highly corrosive on skin contact.

(iii) Common school thermometer has maximum calibration of 110°C and thus cannot be used. It breaks/cracks.

(i)Using conentrated sulphuric(VI)acid

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Some broken porcelain or sand should be put in the flask when heating to:

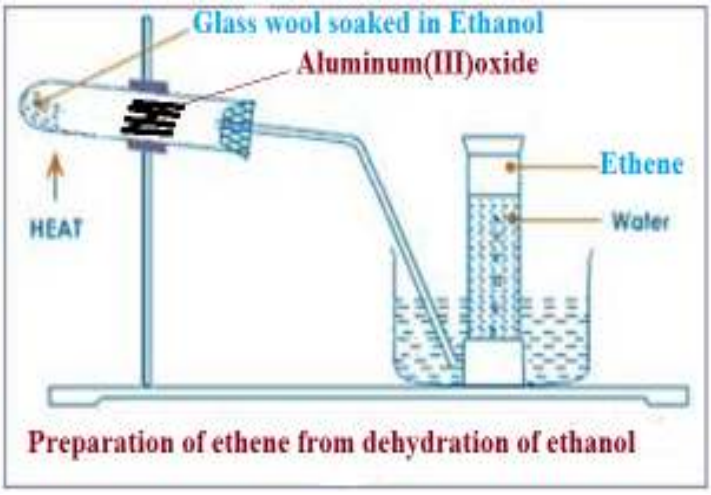
(i)prevent bumping which may break the flask.

(ii)ensure uniform and smooth boiling of the mixture

The temperatures should be maintained at above160ºC.

At lower temperatures another compound -**ether** is predominantly formed instead of ethene gas**.**

(ii)Using aluminium(III)oxide



(e)Properties of alkenes

I. Physical properties

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Like alkanes, alkenes are colourles gases, solids and liquids that are not poisonous. They are slightly soluble in water.
The solubility in water decrease as the carbon chain and as the molar mass increase but very soluble in organic solvents like tetrachloromethane and methylbenzene.
The melting and boiling point increase as the carbon chain increase.
This is because of the increase in van-der-waals /intermolecular forces as the carbon chain increase.
The 1st four straight chain alkenes (ethene,propane,but-1-ene and pent-1-ene)are gases at room temperature and pressure.
The density of straight chain alkenes,like alkanes, increase with increasing carbon chain as the intermolecular forces increases reducing the volume occupied by a given mass of the alkene.

Summary of physical properties of the 1st five alkenes

Alkene	General formula	Melting point(°C)	Boiling point(K)	State at room(298K) temperature and pressure atmosphere (101300Pa)
Ethene	CH_2CH_2	-169	-104	gas
Propene	CH_3 $CHCH_2$	-145	-47	gas
Butene	CH_3CH_2 $CHCH_2$	-141	-26	gas
Pent-1-ene	$CH_3(CH_2$ $CHCH_2$	-138	30	liquid
Hex-1-ene	$CH_3(CH_2)$ $CHCH_2$	-98	64	liquid

II. Chemical properties

(a)Burning/combustion

Alkenes burn with a **yellow**/ luminous **sooty**/ smoky flame in **excess** air to form carbon(IV) oxide and water.

Alkene + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

Alkenes burn with a **yellow**/ luminous **sooty**/ smoky flame in **limited** air to form carbon(II) oxide and water.

Alkene + Air -> carbon(**II**) oxide + water (limited air)

Burning of alkenes with a **yellow**/ luminous **sooty**/ smoky flame is a confirmatory test for the **presence** of the **=C=C=** double bond because they have **higher C:H ratio.**

A homologous series with $-\overset{|}{C}=\overset{|}{C}-$ double or $\equiv C-C$ triple bond is said to be **unsaturated.**

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A homologous series with —C— C— single bond is said to be **saturated**.Most of the reactions of the unsaturated compound involve trying to be saturated to form a

—C— C— single bond .

Examples of burning alkenes

1.(a) Ethene when ignited burns with a **yellow sooty** flame in **excess** air to form carbon(IV) oxide and water.

Ethene + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l/g)$

(b) Ethene when ignited burns with a **yellow sooty** flame in **limited** air to form carbon(II) oxide and water.

Ethene + Air -> carbon(**II**) oxide + water (limited air)

$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l/g)$

2.(a) Propene when ignited burns with a **yellow sooty** flame in **excess** air to form carbon(IV) oxide and water.

Propene + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

$2C_3H_6(g) + 9O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l/g)$

(a) Propene when ignited burns with a **yellow sooty** flame in **limited** air to form carbon(II) oxide and water.

Propene + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

$C_3H_6(g) + 3O_2(g) \rightarrow 3CO(g) + 3H_2O(l/g)$

(b)Addition reactions

An addition reaction is one which an unsaturated compound reacts to form a saturated compound.Addition reactions of alkenes are named from the reagent used to cause the addtion/convert the double =C=C= to single C-C bond.

(i)Hydrogenation

Hydrogenation is an addition reaction in which **hydrogen** in presence of **Palladium/Nickel** catalyst at high temperatures react with alkenes to form alkanes.

Examples

1.When Hydrogen gas is passed through liquid vegetable and animal **oil** at about 180ºC in presence of Nickel catalyst,solid **fat** is formed.

Hydrogenation is thus used to **harden** oils to solid fat especially margarine.

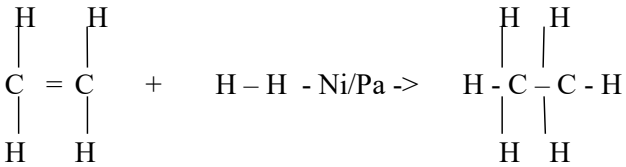
During hydrogenation, one hydrogen atom in the hydrogen molecule attach itself to one carbon and the other hydrogen to the second carbon breaking the double bond to single bond.

Chemical equation

$H_2C{=}CH_2 + H_2 \xrightarrow{Ni/Pa} H_3C - CH_3$

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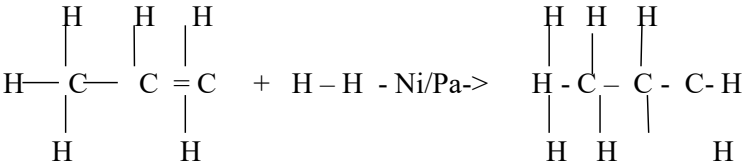
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2.Propene undergo hydrogenation to form Propane

Chemical equation

$H_3C\ CH{=}CH_2 + H_2 \xrightarrow{-Ni/Pa} H_3C\ CH - CH_3$

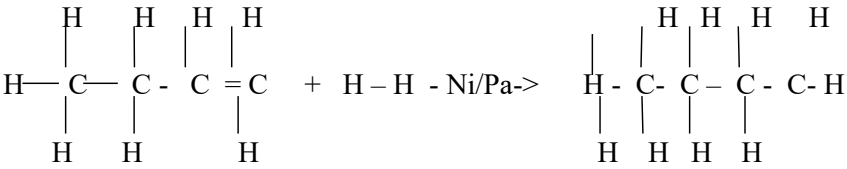


3.Both But-1-ene and But-2-ene undergo hydrogenation to form Butane

Chemical equation

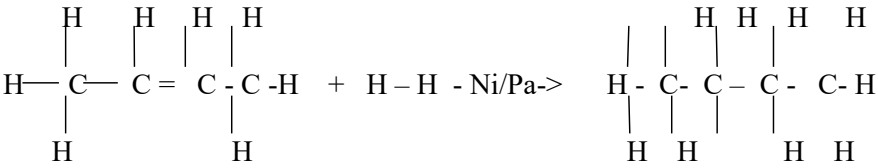
But-1-ene + Hydrogen –Ni/Pa-> Butane

$H_3C\ CH_2\ CH{=}CH_2 + H_2 \xrightarrow{-Ni/Pa} H_3C\ CH_2CH - CH_3$



But-2-ene + Hydrogen –Ni/Pa-> Butane

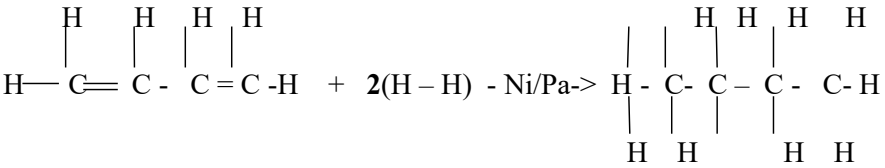
$H_3C\ CH_2 {=} CH\ CH_2 + H_2 \xrightarrow{-Ni/Pa} H_3C\ CH_2CH - CH_3$



4. But-1,3-diene should undergo hydrogenation to form Butane. The reaction uses **two** moles of hydrogen molecules/**four** hydrogen atoms to break the two double bonds.

But-1,3-diene + Hydrogen –Ni/Pa-> Butane

$H_2C\ CH\ CH{=}CH_2 + \mathbf{2}H_2 \xrightarrow{-Ni/Pa} H_3C\ CH_2CH - CH_3$



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(ii) Halogenation.

Halogenation is an addition reaction in which a halogen (Fluorine, chlorine, bromine, iodine) reacts with an alk**e**ne to form an alkane.

The double bond in the alk**e**ne break and form a single bond.

The colour of the halogen **fades** as the number of moles of the halogens remaining unreacted decreases/reduces.

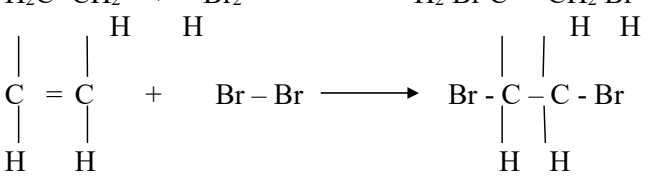
One bromine atom bond at the 1st carbon in the double bond while the other goes to the 2nd carbon.

Examples

1Ethene reacts with bromine to form 1,2-dibromoethane.

Chemical equation

$H_2C=CH_2 + Br_2 \longrightarrow H_2\,Br\,C - CH_2\,Br$



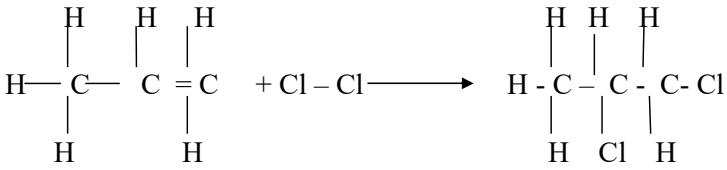
Ethene + Bromine ⟶ 1,2-dibromoethane

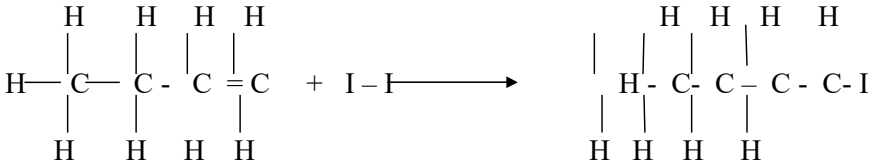
2.Propene reacts with chlorine to form 1,2-dichloropropane.

Chemical equation

$H_3C\,CH=CH_2 + Cl_2 \longrightarrow H_3C\,CHCl - CH_2Cl$

Propene + Chlorine ⟶ 1,2-dichloropropane





3.Both But-1-**e**ne and But-2-**e**ne undergo halogenation with iodine to form 1,2-diiodobutane and 2,3-diiodobutane

Chemical equation

But-1-**e**ne + iodine ⟶ 1,2 diiodobutane

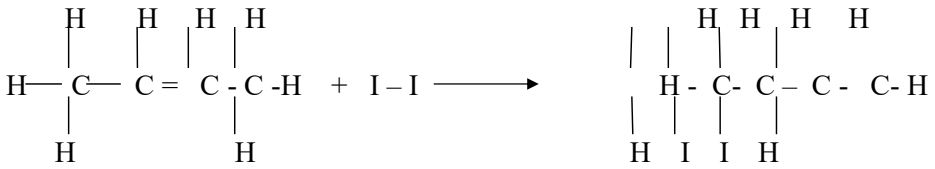
$H_3C\,CH_2\,CH=CH_2 + I_2 \longrightarrow H_3C\,CH_2CH\,I - CH_2I$

But-2-**e**ne + Iodine ⟶ 2,3-diiodobutane

$H_3C\,CH=CH-CH_2 + F_2 \longrightarrow H_3C\,CHICHI - CH_3$

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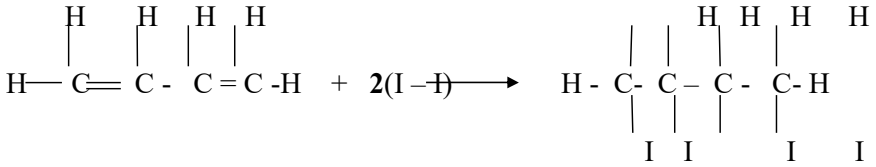
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4. But-1,3-diene should undergo halogenation to form Butane. The reaction uses **two** moles of iodine molecules/**four** iodine atoms to break the two double bonds.

But-1,3-diene + iodine ⟶ 1,2,3,4-tetraiodobutane

H_2C= CH CH=CH_2 + **2**I_2 ⟶ H_2CI CHICHI - CHI



(iii) Reaction with hydrogen halides.

Hydrogen halides reacts with alkene to form a halogenoalkane. The double bond in the alkene break and form a single bond.

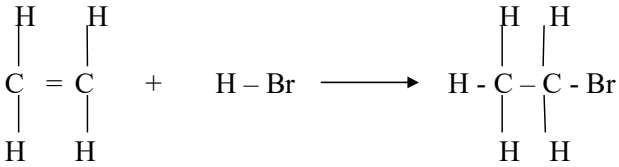
The main compound is one which the **hydrogen** atom bond at the carbon with **more hydrogen** .

Examples

1. Ethene reacts with hydrogen bromide to form bromoethane.

Chemical equation

H_2C=CH_2 + HBr ⟶ H_3 C - CH_2 Br



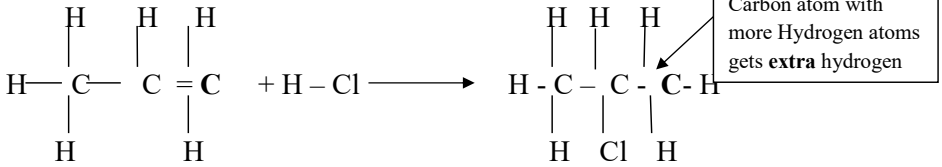
Ethene + Bromine ⟶ bromoethane

2. Propene reacts with hydrogen iodide to form 2-iodopropane.

Chemical equation

H_3C CH=CH_2 + HI ⟶ H_3C CHI - CH_3

Propene + Chlorine ⟶ 2-chloropropane



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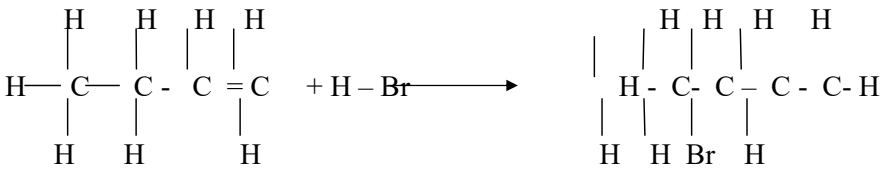
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3. Both But-1-**e**ne and But-2-**e**ne reacts with hydrogen bromide to form 2-bromobutane

Chemical equation

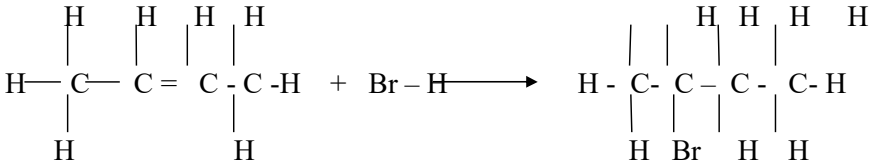
But-1-**e**ne + hydrogen bromide ⟶ 2-bromobutane

$H_3C\ CH_2\ CH{=}CH_2 + HBr \longrightarrow H_3C\ CH_2CHBr\ -CH_3$



But-2-**e**ne + Hydrogen bromide ⟶ 2-bromobutane

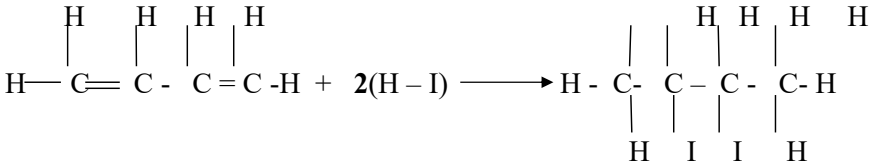
$H_3C\ CH{=}\ CH\text{-}CH_2 + HBr \longrightarrow H_3C\ CHBrCH_2\ -\ CH_3$



4. But-1,3-di**e**ne react with hydrogen iodide to form 2,3- diiodobutane. The reaction uses **two** moles of hydrogen iodide molecules/**two** iodine atoms and two hydrogen atoms to break the two double bonds.

But-1,3-di**e**ne + iodine ⟶ 2,3-diiodobutane

$H_2C{=}\ CH\ CH{=}CH_2 + \mathbf{2}HI_2 \longrightarrow H_3CCHICHI\ -\ CH_3$



(iv) Reaction with bromine/chlorine water.

Chlorine and bromine water is formed when the halogen is dissolved in distilled water.Chlorine water has the formular HOCl(hypochlorous/chloric(I)acid) .Bromine water has the formular HOBr(hydrobromic(I)acid).

During the addition reaction .the halogen move to one carbon and the OH to the other carbon in the alkene at the =C=C= double bond to form a **halogenoalkanol.**

Bromine water + Alkene -> bromoalkan**o**l

Chlorine water + Alkene -> bromoalkan**o**l

Examples

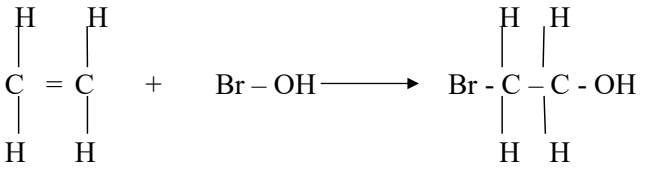
1Ethene reacts with bromine water to form bromoethan**ol**.

Chemical equation

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$H_2C{=}CH_2 + HOBr \longrightarrow H_2\,Br\,C - CH_2\,OH$



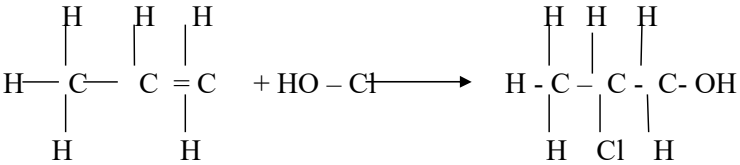
Ethene + Bromine water ⟶ bromoethan**ol**

2.Propene reacts with chlorine water to form chloropropan-2-ol / 2-chloropropan-1-ol.

Chemical equation

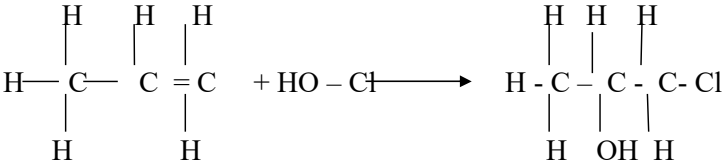
I. $H_3C\,CH{=}CH_2 + HOCl \longrightarrow H_3C\,CHCl - CH_2OH$

Propene + Chlorine water ⟶ 2-chloropropane



II. $H_3C\,CH{=}CH_2 + HOCl \longrightarrow H_3C\,CHOH - CH_2Cl$

Propene + Chlorine ⟶ chloropropan-2-ol

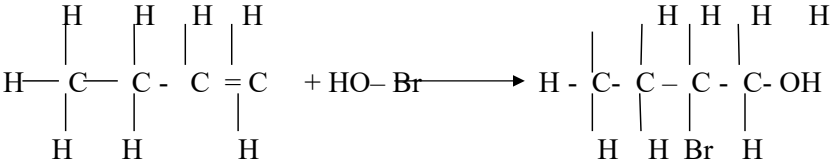


3.Both But-1-**e**ne and But-2-**e**ne react with bromine water to form 2-bromobutan-1-**ol** /3-bromobutan-2-**ol** respectively

Chemical equation

I. But-1-**e**ne + bromine water ⟶ 2-bromobutan-1-ol

$H_3C\,CH_2\,CH{=}CH_2 + HOBr \longrightarrow H_3C\,CH_2CH\,Br - CH_2OH$

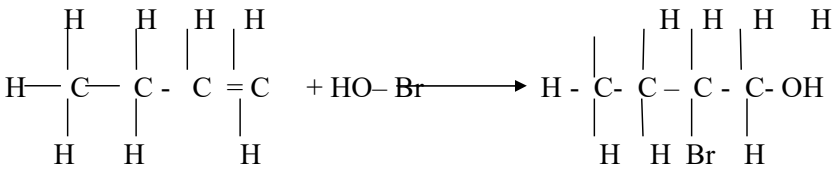


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II. But-2-ene + bromine water ⟶ 3-bromobutan-2-ol

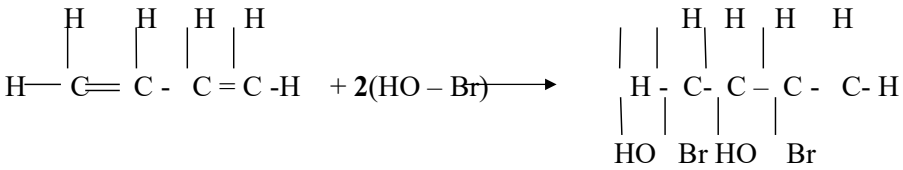
$H_3C\ CH{=}\ CHCH_3 + HOBr \longrightarrow H_3C\ CH_2OHCH\ Br\ CH_3$



4. But-1,3-diene reacts with bromine water to form Butan-1,3-diol.
The reaction uses **two** moles of bromine water molecules to break the two double bonds.

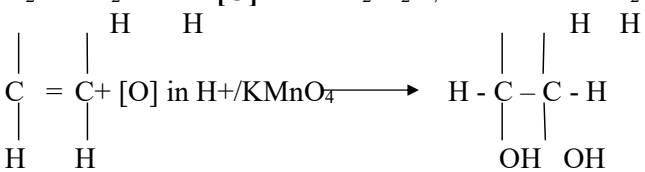
But-1,3-diene + bromine water ⟶ 2,4-dibromobutan-1,3-diol

$H_2C{=}\ CH\ CH{=}CH_2 + \mathbf{2}HOBr \longrightarrow H_2COH\ CHBrCHOH\ CHBr$



(v) Oxidation.
Alkenes are oxidized to alkanols with **duo/double** functional groups by oxidizing agents.
When an alkene is bubbled into orange acidified potassium/sodium dichromate (VI) solution,the colour of the oxidizing agent changes to green.
When an alkene is bubbled into purple acidified potassium/sodium manganate(VII) solution, the oxidizing agent is decolorized.
Examples
1Ethene is oxidized to ethan-1,2-di**ol** by acidified potassium/sodium manganate(VII) solution/ acidified potassium/sodium dichromate(VI) solution.
The purple acidified potassium/sodium manganate(VII) solution is decolorized.
The orange acidified potassium/sodium dichromate(VI) solution turns to green.
Chemical equation

$H_2C{=}CH_2 \xrightarrow{[O]\ \text{in}\ H^+/K_2Cr_2O_7} HO\ CH_2\ -\ CH_2\ OH$



Ethene + [O] in $H^+/KMnO_4$ ⟶ ethan-1,2-di**ol**

2. Propene is oxidized to propan-1,2-di**ol** by acidified potassium/sodium manganate(VII) solution/ acidified potassium/sodium dichromate(VI) solution.
The purple acidified potassium/sodium manganate(VII) solution is decolorized.

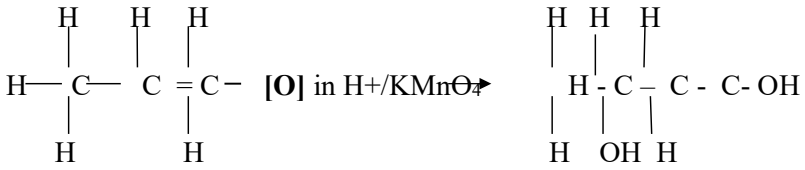
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The orange acidified potassium/sodium dichromate(VI) solution turns to green.
Chemical equation

$H_3C\ CH{=}CH_2 \xrightarrow{[O]\text{ in } H^+/KMnO_4} H_3C\ CHOH - CH_2OH$

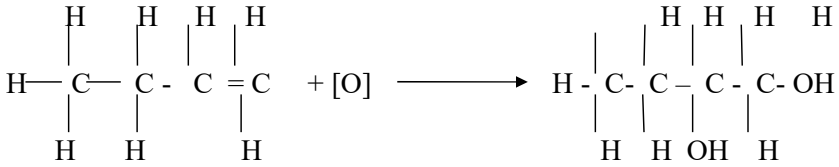
Propene —**[O]** in H+/$KMnO_4$→ propan-1,2-diol



3.Both But-1-**e**ne and But-2-**e**ne react with bromine water to form butan-1,2-diol and butan-2,3-diol
Chemical equation

I.But-1-**e**ne + **[O]** in $H^+/KMnO_4$ ⟶ butan-1,2-diol

$H_3C\ CH_2\ CH{=}CH_2 + [O] \longrightarrow H_3C\ CH_2CHOH - CH_2OH$



(v) Hydrolysis.
Hydrolysis is the reaction of a compound with water/addition of H-OH to a compound.
Alkenes undergo hydrolysis to form alkanols .
This takes place in two steps:
(i)Alkenes react with **concentrated sulphuric(VI)acid** at room temperature and pressure to form **alkylhydrogen sulphate(VI)**.

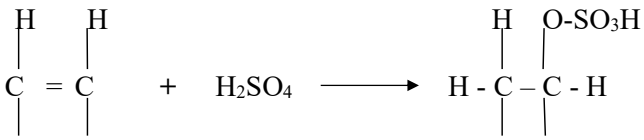
Alkenes + concentrated sulphuric(VI)acid -> alkylhydrogen sulphate(VI)

(ii)On adding **water** to alkylhydrogen sulphate(VI) then warming, an alkanol is formed.

alkylhydrogen sulphate(VI) + water -warm-> Alkanol.

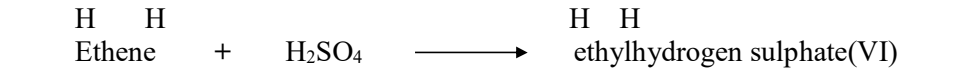
Examples
(i)Ethene reacts with cold concentrated sulphuric(VI)acid to form ethyl hydrogen sulphate(VII)
Chemical equation

$H_2C{=}CH_2 + H_2SO_4 \longrightarrow CH_3 - CH_2OSO_3H$



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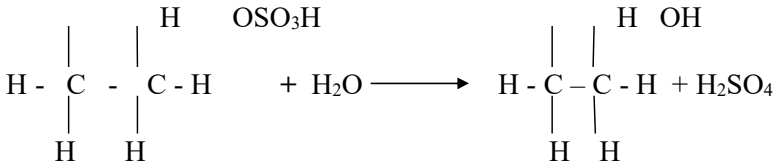
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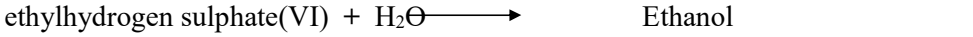


(ii) Ethylhydrogen sulphate(VI) is hydrolysed by water to ethanol

Chemical equation

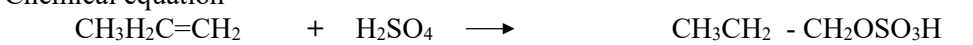


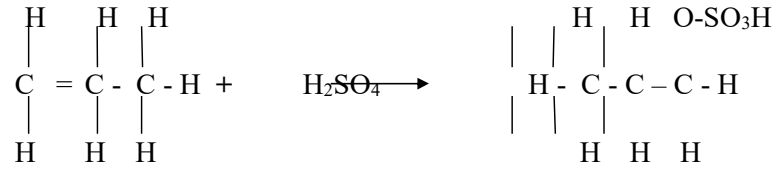


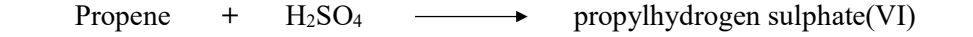


2. Propene reacts with cold concentrated sulphuric(VI)acid to form propyl hydrogen sulphate(VII)

Chemical equation

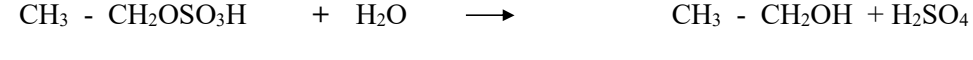


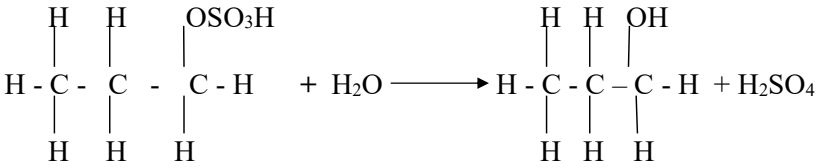


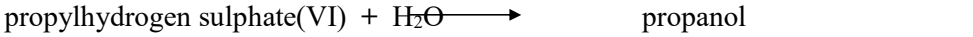


(ii) Propylhydrogen sulphate(VI) is hydrolysed by water to propanol

Chemical equation







(vi) Polymerization/self addition

Addition polymerization is the process where a small unsaturated monomer (alkene) molecule join together to form a large saturated molecule.

Only alkenes undergo addition polymerization.

Addition polymers are named from the alkene/monomer making the polymer and adding the prefix "**poly**" before the name of monomer to form a **polyalkene**

During addition polymerization

(i)the double bond in alkenes break

(ii)free radicals are formed

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(iii)the free radicals collide with each other and join to form a larger molecule. The more collisions the larger the molecule.

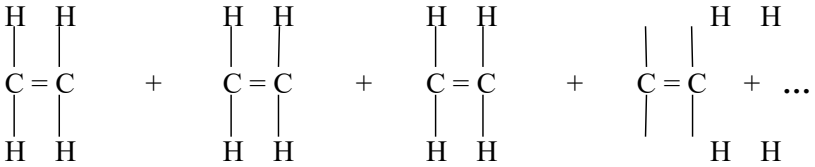
Examples of addition polymerization

1.Formation of Polyethene

Polyethene is an addition polymer formed when ethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

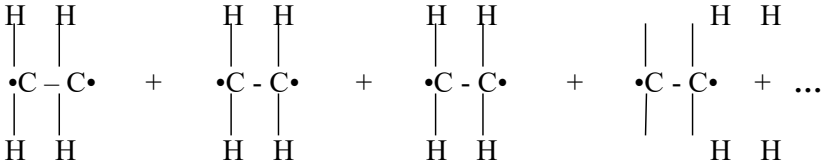
During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)



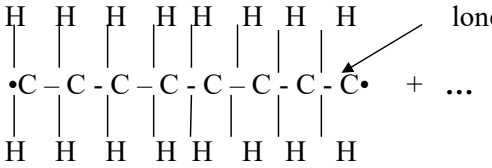
Ethene + Ethene + Ethene + Ethene + …

(ii)the double bond joining the ethane molecule break to free readicals



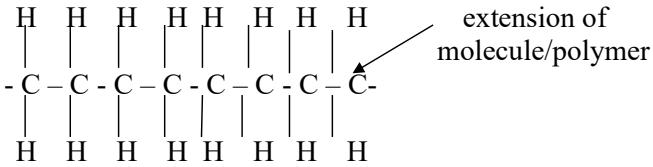
Ethene radical + Ethene radical + Ethene radical + Ethene radical + …

(iii)the free radicals collide with each other and join to form a larger molecule



Lone pair of electrons can be used to join more monomers to form longer polyethene.

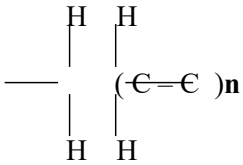
Polyethene molecule can be represented as:



Since the molecule is a **repetition** of one monomer, then the polymer is:

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Where **n** is the number of monomers in the polymer. The number of monomers in the polymer can be determined from the molar mass of the polymer and monomer from the relationship:

Number of monomers/repeating units in monomer = Molar mass polymer
Molar mass monomer

Examples

Polythene has a molar mass of 4760.Calculate the number of ethene molecules in the polymer(C=12.0, H=1.0)

Number of monomers/repeating units in polyomer = Molar mass polymer
Molar mass monomer

=> Molar mass ethene (C_2H_4)= 28 Molar mass polyethene = 4760

Substituting 4760 = 170 ethene molecules
28

The **commercial** name of polyethene is **polythene**.

It is an elastic, tough, transparent and durable plastic.

Polythene is used:

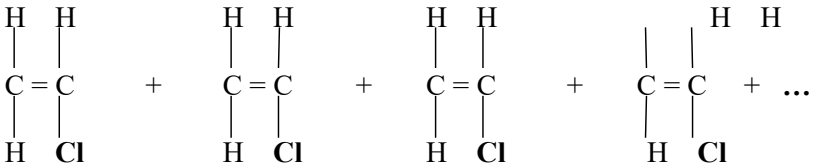
(i)in making plastic bag
(ii)bowls and plastic bags
(iii)packaging materials

2.Formation of Polychlorethene

Polychloroethene is an addition polymer formed when chloroethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)



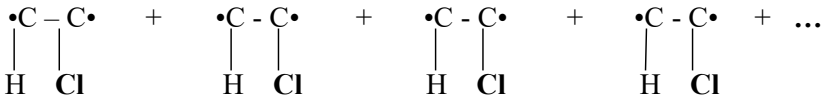
chloroethene + chloroethene + chloroethene + chloroethene + **…**

(ii)the double bond joining the chloroethene molecule break to free radicals

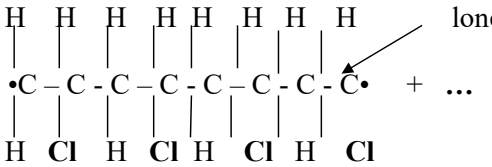


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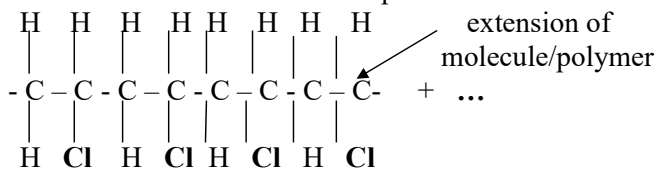


(iii)the free radicals collide with each other and join to form a larger molecule

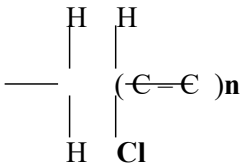


Lone pair of electrons can be used to join more monomers to form longer polychloroethene.

Polychloroethene molecule can be represented as:



Since the molecule is a repetition of one monomer, then the polymer is:



Examples

Polychlorothene has a molar mass of 4760.Calculate the number of chlorethene molecules in the polymer(C=12.0, H=1.0,Cl=35.5)

Number of monomers/repeating units in monomer = Molar mass polymer
Molar mass monomer

=> Molar mass ethene (C_2H_3Cl)= 62.5 Molar mass polyethene = 4760

Substituting 4760 = 77.16 => 77 polychloroethene molecules**(whole number)**
62.5

The **commercial** name of polychloroethene is **polyvinylchloride(PVC)**. It is a tough, non-transparent and durable plastic. PVC is used:

(i)in making plastic rope
(ii)water pipes
(iii)crates and boxes

3.Formation of Polyphenylethene

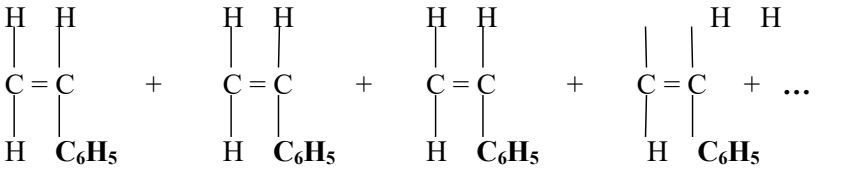
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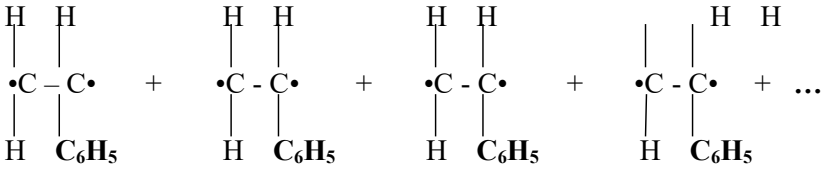
Polyphenylethene is an addition polymer formed when phenylethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

During polymerization:

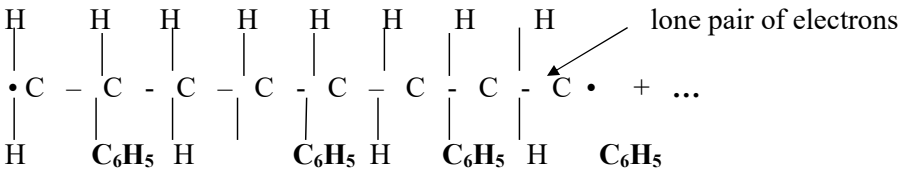
(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)



(ii)the double bond joining the phenylethene molecule break to free radicals

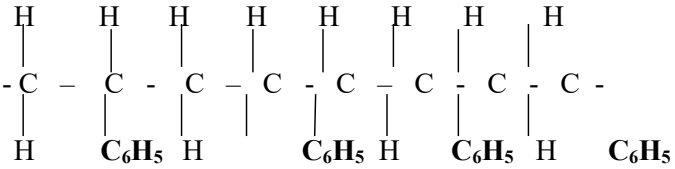


(iii)the free radicals collide with each other and join to form a larger molecule

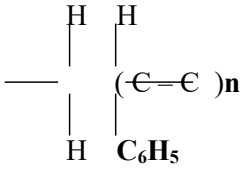


Lone pair of electrons can be used to join more monomers to form longer polyphenylethene.

Polyphenylethene molecule can be represented as:



Since the molecule is a repetition of one monomer, then the polymer is:



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Examples

Polyphenylthene has a molar mass of 4760.Calculate the number of phenylethene molecules in the polymer(C=12.0, H=1.0,)

Number of monomers/repeating units in monomer = Molar mass polymer
Molar mass monomer

=> Molar mass ethene (C_8H_8)= 104 Molar mass polyethene = 4760

Substituting 4760 = 45.7692 =>45 polyphenylethene molecules**(whole number)**
104

The **commercial** name of polyphenylethene is **polystyrene**. It is a very light durable plastic. Polystyrene is used:

(i)in making packaging material for carrying delicate items like computers, radion,calculators.

(ii)ceiling tiles

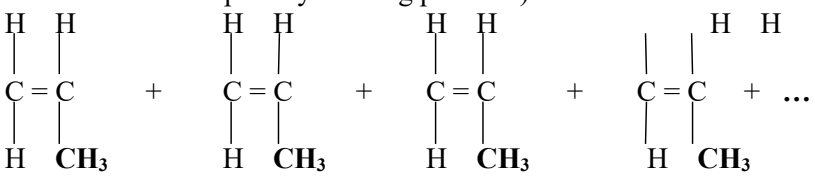
(iii)clothe linings

4.Formation of Polypropene

Polypropene is an addition polymer formed when propene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

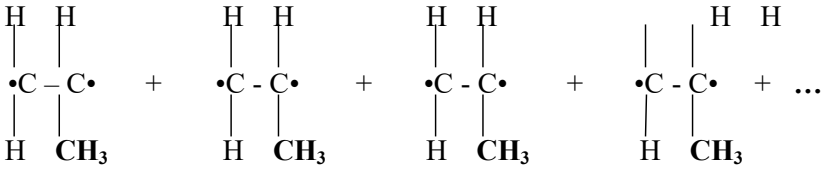
During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)

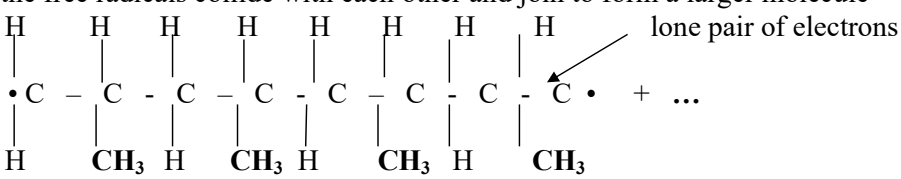


propene + propene + propene+ propene + …

(ii)the double bond joining the phenylethene molecule break to free radicals



(iii)the free radicals collide with each other and join to form a larger molecule

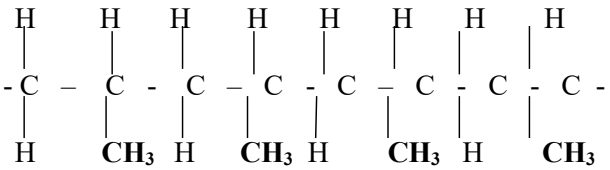


Lone pair of electrons can be used to join more monomers to form longer propene.

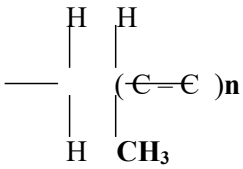
propene molecule can be represented as:

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Since the molecule is a repetition of one monomer, then the polymer is:



Examples

Polypropene has a molar mass of 4760.Calculate the number of propene molecules in the polymer(C=12.0, H=1.0,)

Number of monomers/repeating units in monomer = Molar mass polymer
Molar mass monomer

=> Molar mass propene (C_3H_8)= 44 Molar mass polyethene = 4760

Substituting 4760 = 108.1818 =>108 propene molecules**(whole number)**
44

The **commercial** name of polyphenylethene is **polystyrene**. It is a very light durable plastic. Polystyrene is used:

(i)in making packaging material for carrying delicate items like computers, radion,calculators.

(ii)ceiling tiles

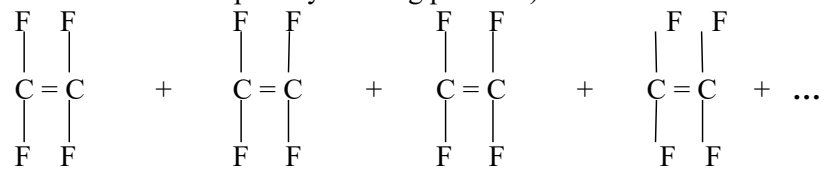
(iii)clothe linings

5.Formation of Polytetrafluorothene

Polytetrafluorothene is an addition polymer formed when tetrafluoroethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)



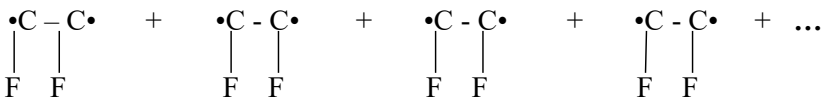
tetrafluoroethene + tetrafluoroethene+ tetrafluoroethene+ tetrafluoroethene + …

(ii)the double bond joining the tetrafluoroethene molecule break to free radicals

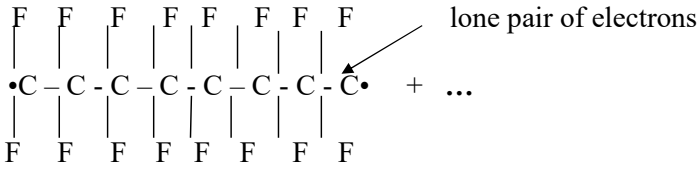


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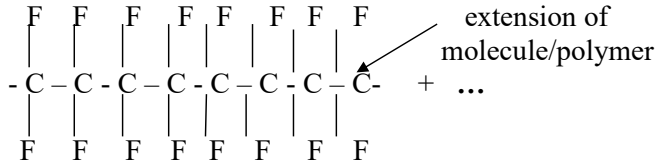


(iii)the free radicals collide with each other and join to form a larger molecule

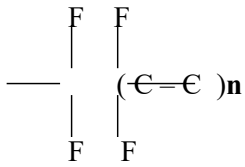


Lone pair of electrons can be used to join more monomers to form longer polytetrafluoroethene.

polytetrafluoroethene molecule can be represented as:



Since the molecule is a repetition of one monomer, then the polymer is:



Examples

Polytetrafluorothene has a molar mass of 4760.Calculate the number of tetrafluoroethene molecules in the polymer(C=12.0, ,F=19)

Number of monomers/repeating units in monomer = Molar mass polymer

Molar mass monomer

=> Molar mass ethene (C_2F_4)= 62.5 Molar mass polyethene = 4760

Substituting 4760 = 77.16 => 77 polychloroethene molecules**(whole number)**

62.5

The **commercial** name of polytetrafluorethene(**P.T.F.E**) is **Teflon(P.T.F.E)**. It is a tough, non-transparent and durable plastic. PVC is used:

(i)in making plastic rope

(ii)water pipes

(iii)crates and boxes

6.Formation of rubber from Latex

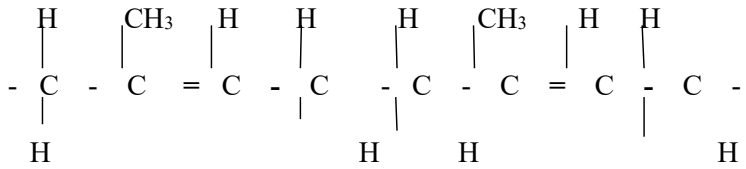
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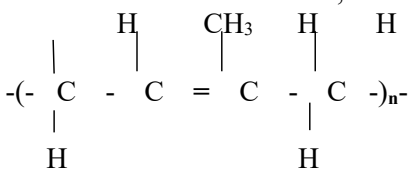
Natural rubber is obtained from rubber trees.
During harvesting an incision is made on the rubber tree to produce a milky white substance called **latex.**
Latex is a mixture of rubber and lots of water.
The latex is then added an acid to coagulate the rubber.
Natural rubber is a polymer of 2-methylbut-1,3-diene ;



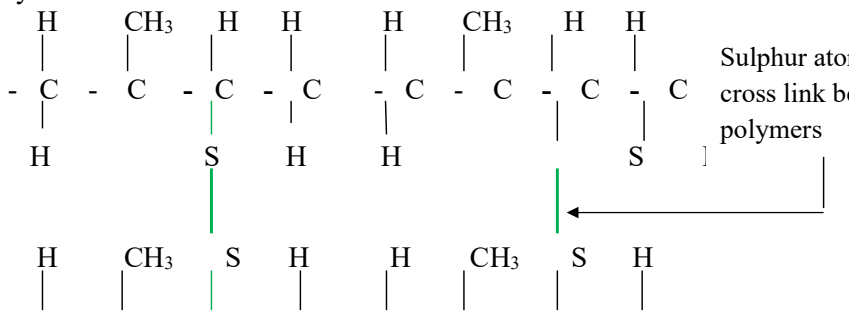
During natural polymerization to rubber, one double C=C bond break to self add to another molecule. The double bond remaining move to carbon "2" thus;



Generally the structure of rubber is thus;

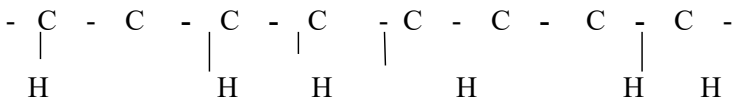


Pure rubber is soft and sticky.It is used to make erasers, car tyres. Most of it is vulcanized.Vulcanization is the process of heating rubber with sulphur to make it harder/tougher.
During vulcanization the sulphur atoms form a cross link between chains of rubber molecules/polymers. This decreases the number of C=C double bonds in the polymer.



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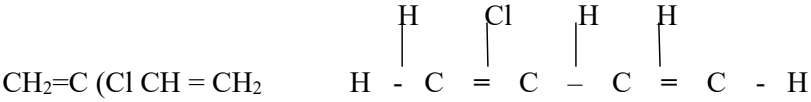


Vulcanized rubber is used to make **tyres**, **shoes** and **valves**.

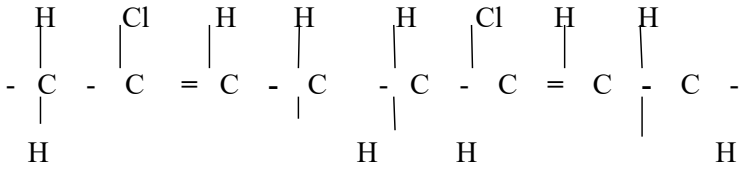
7.Formation of synthetic rubber

Synthetic rubber is able to resist action of oil,abrasion and organic solvents which rubber cannot.

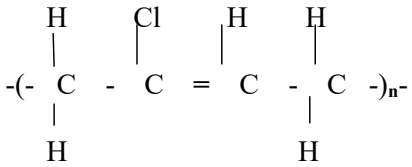
Common synthetic rubber is a polymer of 2-chlorobut-1,3-diene ;



During polymerization to synthetic rubber, one double C=C bond is broken to self add to another molecule. The double bond remaining move to carbon "2" thus;



Generally the structure of rubber is thus;



Rubber is thus strengthened through vulcanization and manufacture of synthetic rubber.

(c)Test for the presence of $\ulcorner C \urcorner C$ – double bond.

(i)Burning/combustion

All unsaturated hydrocarbons with a $\underline{\underline{\ }}C\underline{\underline{\ }}C$ – or $-\overline{C} = C$ – bond burn with a yellow sooty flame.

Experiment

Scoop a sample of the substance provided in a clean metallic spatula. Introduce it on a Bunsen burner.

Observation	Inference
Solid melt then burns with a yellow sooty flame	$-\underline{C}=\underline{C}-$, $-C\overline{=}C-$ bond

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(ii)Oxidation by acidified $KMnO_4/K_2Cr_2O_7$

Bromine water ,Chlorine water and Oxidizing agents acidified $KMnO_4/K_2Cr_2O_7$ change to **unique** colour in presence of $-C=C-$ or $-C\equiv C-$ bond.

Experiment

Scoop a sample of the substance provided into a clean test tube. Add 10cm3 of distilled water. Shake. Take a portion of the solution mixture. Add three drops of acidified $KMnO_4/K_2Cr_2O_7$.

Observation	Inference
Acidified $KMnO_4$ decolorized	$-C=C-$
Orange colour of acidified $K_2Cr_2O_7$ turns green	$-C\equiv C-$ bond
Bromine water is decolorized	
Chlorine water is decolorized	

(d)Some uses of Alkenes

1. In the manufacture of plastic
2. Hydrolysis of ethene is used in industrial manufacture of ethanol.
3. In ripening of fruits.
4. In the manufacture of detergents.

(iii) Alkynes

(a)Nomenclature/Naming

These are hydrocarbons with a general formula C_nH_{2n-2} and $\equiv C-C$ double bond as the functional group . **n** is the number of Carbon atoms in the molecule.

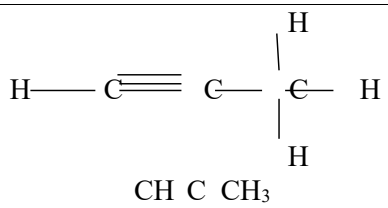
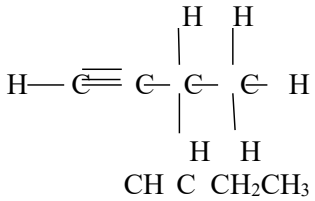
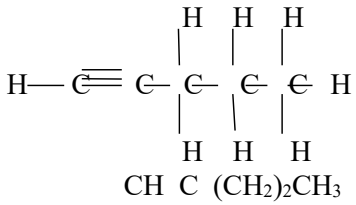
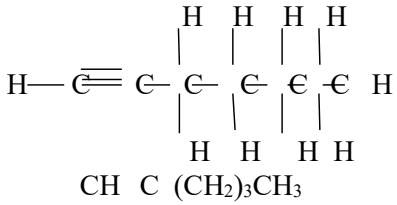
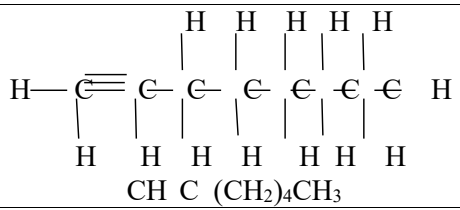
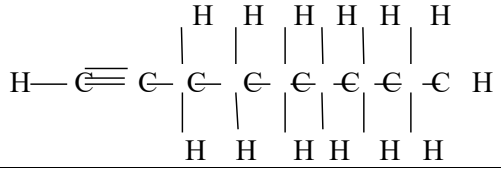
The carbon atoms are linked by at least one **triple** bond to each other and single bonds to hydrogen atoms.

They include:

n	General/ Molecular formula	Structural formula	Name
1		Does not exist	-
2	C_2H_2	H—— C≡ C—— H CH CH	Ethyne

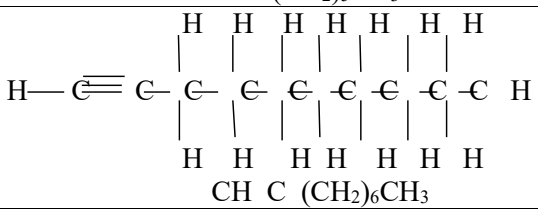
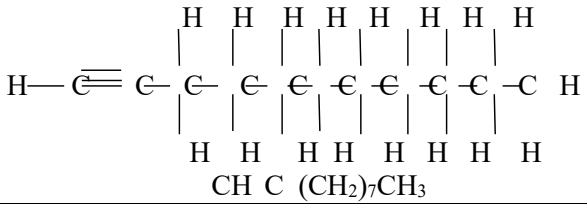
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3	C_3H_4		Propyne
4	C_4H_6		Butyne
5	C_5H_8		Pentyne
6	C_6H_{10}		Hexyne
7	C_7H_{12}		Heptyne
8	C_8H_{14}		Octyne

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		$CH\equiv C-(CH_2)_5CH_3$	
9	C_9H_{16}	 $CH\equiv C-(CH_2)_6CH_3$	Nonyne
10	$C_{10}H_{18}$	 $CH\equiv C-(CH_2)_7CH_3$	Decyne

Note

1. Since carbon is **tetravalent** ,each atom of carbon in the alkyne **MUST** always be bonded using **four** covalent bond /four shared pairs of electrons including at the triple bond.

2. Since Hydrogen is **monovalent** ,each atom of hydrogen in the alkyne **MUST** always be bonded using **one** covalent bond/one shared pair of electrons.

3. One member of the alkyne ,like alkenes and alkanes, differ from the next/previous by a CH_2 group(molar mass of 14 **a**tomic **m**ass **u**nits).They thus form a homologous series.

e.g

Propyne differ from ethyne by (14 a.m.u) one carbon and two Hydrogen atoms from ethyne.

4.A homologous series of alkenes like that of alkanes:

(i) differ by a CH_2 group from the next /previous consecutively

(ii) have similar chemical properties

(iii)have similar chemical formula with general formula C_nH_{2n-2}

(iv)the physical properties also show steady gradual change

5.The $-C\equiv C-$ triple bond in alkyne is the functional group. The functional group is the **reacting site** of the alkynes.

6. The $-C\equiv C-$ triple bond in alkyne can easily be broken to accommodate more /four more monovalent atoms. The $-C\equiv C-$ triple bond in alkynes make it thus **unsaturated** like alkenes**.**

7. Most of the reactions of alkynes like alkenes take place at the $-C\equiv C-$ triple bond.

(b)Isomers of alkynes

Isomers of alkynes have the same molecular **general formula** but different molecular **structural formula**.

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Isomers of alkynes are also named by using the IUPAC(**I**nternational **U**nion of **P**ure and **A**pplied **C**hemistry) system of **nomenclature/naming.**

The IUPAC system of nomenclature of naming alkynes uses the following basic rules/guidelines:

1.Identify the longest continuous/straight carbon chain which contains the **- C = C- triple** bond to get/determine the **parent** alkene.

2. Number the longest chain form the end of the chain which contains the **-C = C- triple** bond so as **- C = C- triple** bond get lowest number possible.

3 Indicate the positions by splitting "**alk**-positions-**yne**" e.g. but-2-yne, pent-1,3-diyne.

4.The position **indicated** must be for the carbon atom at the **lower** position in the **-C ≡ C- triple bond.** i.e

But-2-yne means the triple -C = C- is between Carbon "2"and "3"

Pent-1,3-diyne means there are two triple bonds; one between carbon "1" and "2"and another between carbon "3" and "4"

5. Determine the position, number and type of branches. Name them as methyl, ethyl, propyl e.tc. according to the number of alkyl carbon chains attached to the alkyne. Name them fluoro-,chloro-,bromo-,iodo- if they are halogens

6.Use prefix di-,tri-,tetra-,penta-,hexa- to show the number of **triple - C = C-** bonds and **branches** attached to the alkyne.

7.Position isomers can be formed when the -C = C- triple bond is shifted between carbon atoms e.g.

But-2-yne means the double - C = C- is between Carbon "2"and "3"

But-1-yne means the double - C = C- is between Carbon "1"and "2"

Both But-1-yne and But-2-yne are position isomers of Butyne.

9. Like alkanes and alkynes , an alkyl group can be attached to the alkyne. Chain/branch isomers are thus formed.

Butyne and 2-methyl propyne both have the same general formular but different branching chain.

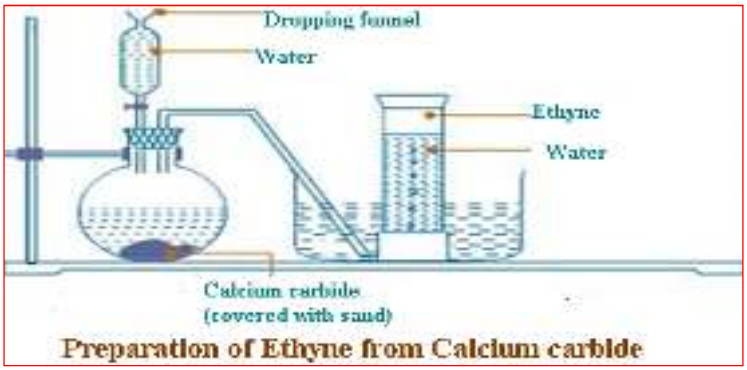
(c)Preparation of Alkynes.

Ethyne is prepared from the reaction of water on calcium carbide. The reaction is highly exothermic and thus a layer of sand should be put above the calcium carbide to absorb excess heat to prevent the reaction flask from breaking.

Copper(II)sulphate(VI) is used to catalyze the reaction

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Preparation of Ethyne from Calcium carbide

Chemical equation

$CaC_2(s) + 2\ H_2O(l) \rightarrow Ca(OH)_2(aq) + C_2H_2(g)$

(d)Properties of alkynes

I. Physical properties

Like alkanes and alkenes, alkynes are colourles gases, solids and liquids that are not poisonous.

They are slightly soluble in water. The solubility in water decrease as the carbon chain and as the molar mass increase but very soluble in organic solvents like tetrachloromethane and methylbenzene. Ethyne has a pleasant taste when pure.

The melting and boiling point increase as the carbon chain increase.

This is because of the increase in van-der-waals /intermolecular forces as the carbon chain increase. The 1st three straight chain alkynes (ethyne,propyne and but-1-yne)are gases at room temperature and pressure.

The density of straight chain alkynes increase with increasing carbon chain as the intermolecular forces increases reducing the volume occupied by a given mass of the alkyne.

Summary of physical properties of the 1st five alkenes

Alkyne	General formula	Melting point(°C)	Boiling point(°C)	State at room(298K) temperature and pressure atmosphere (101300Pa)
Ethyne	CH CH	-82	-84	gas
Propyne	CH_3 C CH	-103	-23	gas
Butyne	CH_3CH_2 CCH	-122	8	gas
Pent-1-yne	$CH_3(CH_2)_2$ CCH	-119	39	liquid
Hex-1-yne	$CH_3(CH_2)_3$C CH	-132	71	liquid

II. Chemical properties

(a)Burning/combustion

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Alkynes burn with a **yellow**/ luminous very **sooty**/ smoky flame in **excess** air to form carbon(IV) oxide and water.

Alkyne + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

Alkenes burn with a **yellow**/ luminous very**sooty**/ smoky flame in **limited** air to form carbon(II) oxide/carbon and water.

Alkyne + Air -> carbon(**II**) oxide /carbon + water (limited air)

Burning of alkynes with a **yellow**/ luminous **sooty**/ smoky flame is a confirmatory test for the **presence** of the **- C = C –** triple bond because they have very **high C:H ratio.**

Examples of burning alkynes

1.(a) Ethyne when ignited burns with a **yellow** very **sooty** flame in **excess** air to form carbon(IV) oxide and water.

Ethyne + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l/g)$

(b) Ethyne when ignited burns with a **yellow sooty** flame in **limited** air to form a mixture of unburnt carbon and carbon(II) oxide and water.

Ethyne + Air -> carbon(**II**) oxide + water (limited air)

$C_2H_2(g) + O_2(g) \rightarrow 2CO_2(g) + C + 2H_2O(l/g)$

2.(a) Propyne when ignited burns with a **yellow sooty** flame in **excess** air to form carbon(IV) oxide and water.

Propyne + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

$C_3H_4(g) + 4O_2(g) \rightarrow 3CO_2(g) + 2H_2O(l/g)$

(a) Propyne when ignited burns with a **yellow sooty** flame in **limited** air to form carbon(II) oxide and water.

Propene + Air -> carbon(**IV**) oxide + water (excess air/oxygen)

$2C_3H_4(g) + 5O_2(g) \rightarrow 6CO(g) + 4H_2O(l/g)$

(b)Addition reactions

An addition reaction is one which an unsaturated compound reacts to form a saturated compound. Addition reactions of alkynes are also named from the reagent used to cause the addition/convert the triple - C = C- to single C- C bond.

(i)Hydrogenation

Hydrogenation is an addition reaction in which **hydrogen** in presence of **Palladium/Nickel** catalyst at 150°C temperatures react with alkynes to form alkenes then alkanes.

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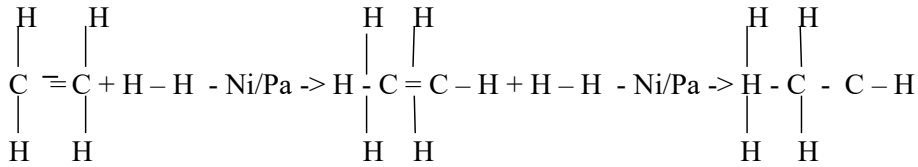
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Examples

1.During hydrogenation, **two** hydrogen atom in the hydrogen molecule attach itself to one carbon and the other **two** hydrogen to the second carbon breaking the **triple** bond to **double** the **single**.

Chemical equation

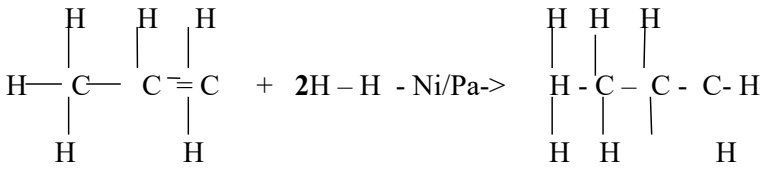
$$HC \equiv CH + H_2 \xrightarrow{-Ni/Pa} H_2C = CH_2 + H_2 \xrightarrow{-Ni/Pa} H_2C - CH_2$$



2.Propyne undergo hydrogenation to form Propane

Chemical equation

$$H_3C\,CH \equiv CH_2 + 2H_2 \xrightarrow{-Ni/Pa} H_3C\,CH - CH_3$$

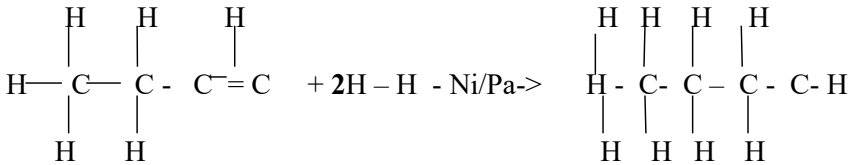


3(a) But-1-yne undergo hydrogenation to form Butane

Chemical equation

But-1-yne + Hydrogen –Ni/Pa-> Butane

$$H_3C\,CH_2\,C \equiv CH + 2H_2 \xrightarrow{-Ni/Pa} H_3C\,CH_2CH - CH_3$$

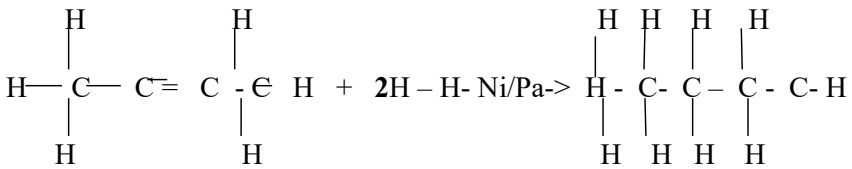


(b) But-2-yne undergo hydrogenation to form Butane

Chemical equation

But-2-yne + Hydrogen –Ni/Pa-> Butane

$$H_3C\,C \equiv C\,CH_2 + 2H_2 \xrightarrow{-Ni/Pa} H_3C\,CH_2CH - CH_3$$



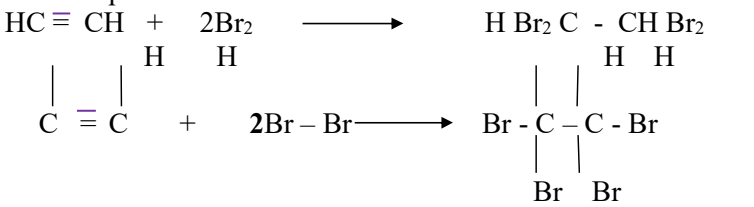
(ii) Halogenation.

Halogenation is an addition reaction in which a halogen (Fluorine, chlorine, bromine, iodine) reacts with an alkyne to form an alkene then alkane.

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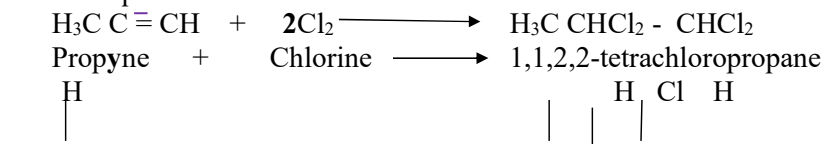
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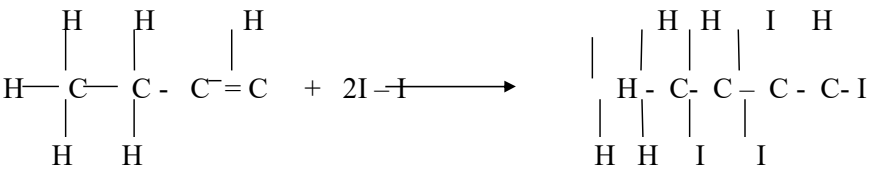
The reaction of alkynes with halogens with alkynes is **faster** than with alkenes. The triple bond in the alkyne break and form a double then single bond.
The colour of the halogen **fades** as the number of moles of the halogens remaining unreacted decreases.
Two bromine atoms bond at the 1st carbon in the triple bond while the other two goes to the 2nd carbon.
Examples
1Ethyne reacts with brown bromine vapour to form 1,1,2,2-tetrabromoethane.
Chemical equation



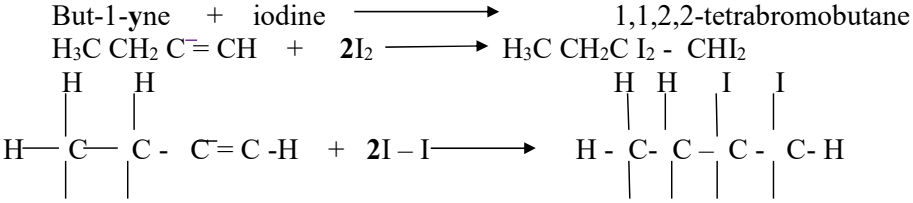
Ethyne + Bromine ⟶ 1,1,2,1-tetrabromoethane

2.Propyne reacts with chlorine to form 1,1,2,2-tetrachloropropane.
Chemical equation





3(a)But-1-yne undergo halogenation to form 1,1,2,2-tetraiodobutane with iodine
Chemical equation



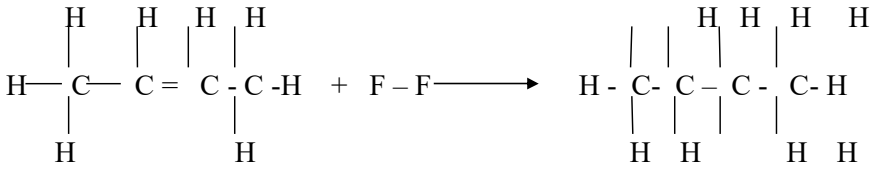
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H H H H H I I

(b) But-2-yne undergo halogenation to form 2,2,3,3-tetrafluorobutane with fluorine But-2-yne + Fluorine 2,2,3,3-tetrafluorobutane

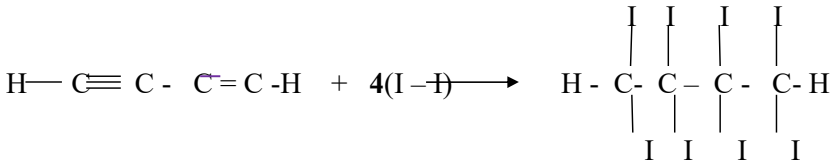
$H_3C\ C \equiv C\ -CH_2 + \mathbf{2}F_2 \longrightarrow H_3C\ CF_2CF_2\ -\ CH_3$



4. But-1,3-diyne should undergo halogenation to form 1,1,2,3,3,4,4 octaiodobutane. The reaction uses **four** moles of iodine molecules/**eight** iodine atoms to break the two(2) triple double bonds at carbon "1" and "2".

But-1,3-diene + iodine $\longrightarrow$ 1,2,3,4-tetraiodobutane

$H\ C \equiv C\ C = C\ H + \mathbf{4}I_2 \longrightarrow H\ C\ I_2\ C\ I_2\ C\ I_2\ C\ H\ I_2$



(iii) Reaction with hydrogen halides.

Hydrogen halides reacts with alkyne to form a halogenoalkene then halogenoalkane. The triple bond in the alkyne break and form a double then single bond.

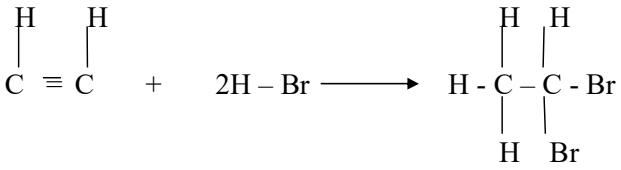
The main compound is one which the **hydrogen** atom bond at the carbon with **more hydrogen** .

Examples

1. Ethyne reacts with hydrogen bromide to form bromoethane.

Chemical equation

$H\ C = C\ H + 2HBr \longrightarrow H_3\ C\ -\ CH\ Br_2$



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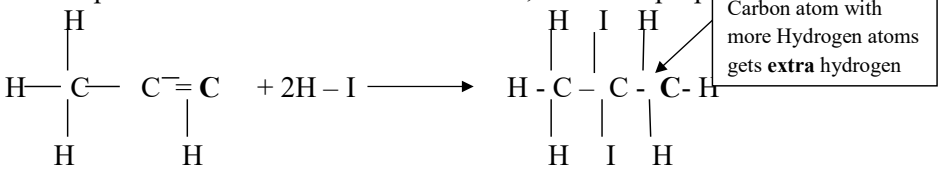
Ethyne + Bromine ⟶ 1,1-dibromoethane

2. Propyne reacts with hydrogen iodide to form 2,2-diiodopropane (as the main product)

Chemical equation

$H_3C\,C\equiv CH + 2HI \longrightarrow H_3C\,CHI_2 - CH_3$

Propene + Chlorine ⟶ 2,2-dichloropropane

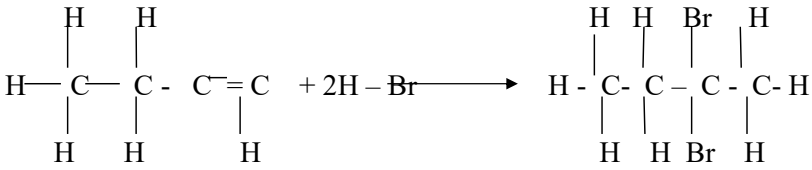


3. Both But-1-yne and But-2-yne reacts with hydrogen bromide to form 2,2-dibromobutane

Chemical equation

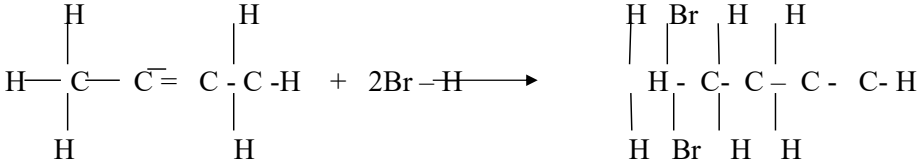
But-1-ene + hydrogen bromide ⟶ 2,2-dibromobutane

$H_3C\,CH_2\,C = CH + 2HBr \longrightarrow H_3C\,CH_2CHBr\,-CH_3$



But-2-yne + Hydrogen bromide ⟶ 2,2-dibromobutane

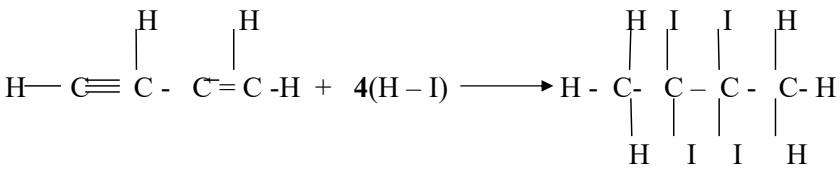
$H_3C\,C\equiv C\,-CH_3 + 2HBr \longrightarrow H_3C\,CBr_2CH_2 - CH_3$



4. But-1,3-diene react with hydrogen iodide to form 2,3- diiodobutane. The reaction uses **four** moles of hydrogen iodide molecules/**four** iodine atoms and two hydrogen atoms to break the two double bonds.

But-1,3-diyne + iodine ⟶ 2,2,3,3-tetraiodobutane

$H\,C\equiv C\,C\equiv C\,H + \mathbf{4}HI \longrightarrow H_3C\,C\,I_2\,C\,I_2\,CH_3$



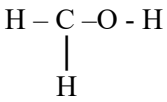
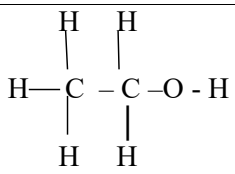
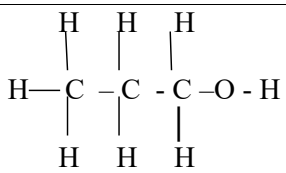
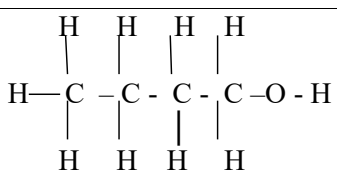
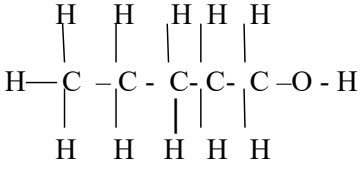
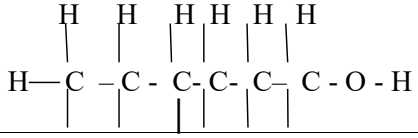
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B.ALKANOLS(Alcohols)

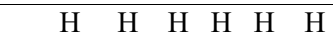
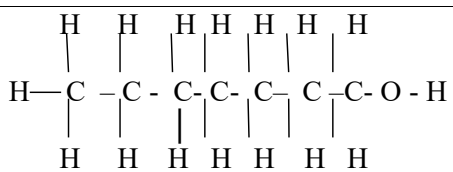
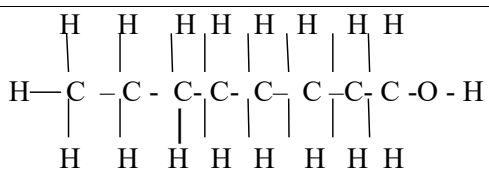
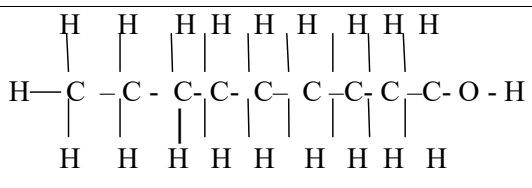
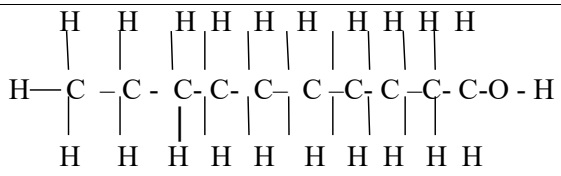
(A) INTRODUCTION.

Alkanols belong to a homologous series of organic compounds with a general formula $\mathbf{C_nH_{2n+1}OH}$ and thus **-OH** as the functional group .The 1st ten alkanols include

n	General / molecular formular	Structural formula	IUPAC name
1	CH_3OH		Methan**ol**
2	$CH_3\ CH_2OH$ $C_2H_5\ OH$		Ethan**ol**
3	$CH_3\ (CH_2)_2OH$ $C_3H_7\ OH$		Propan**ol**
4	$CH_3\ (CH_2)_3OH$ $C_4H_9\ OH$		Butan**ol**
5	$CH_3(CH_2)_4OH$ $C_5H_{11}\ OH$		Pentan**ol**
6	$CH_3(CH_2)_5OH$ $C_6H_{13}\ OH$		Hexan**ol**

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7	$CH_3(CH_2)_6OH$ C_7H_{15} OH		Heptan**ol**
8	$CH_3(CH_2)_7OH$ C_8H_{17} OH		Octan**ol**
9	$CH_3(CH_2)_8OH$ C_9H_{19} OH		Nonan**ol**
10	$CH_3(CH_2)_9OH$ $C_{10}H_{21}$ OH		Decan**ol**

Alkanols like Hydrocarbons(alkanes/alkenes/alkynes) form a homologous series where:
(i)general name is derived from the alkane name then ending with "**-ol**"
(ii)the members have –OH as the fuctional group
(iii)they have the same general formula represented by R-OH where R is an alkyl group.
(iv) each member differ by $–CH_2$ group from the next/previous.
(v)they show a similar and gradual change in their physical properties e.g. boiling and melting points.
(vi)they show similar and gradual change in their chemical properties.

B. ISOMERS OF ALKANOLS.
Alkanols exhibit both structural and position isomerism. The isomers are named by using the following basic guidelines:

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(i)Like alkanes , identify the **longest** carbon chain to be the parent name.
(ii)Identify the position of the **-OH** functional group to give it the **smallest /lowest** position.
(iii) Identify the type and position of the **side** branches.

Practice examples of isomers of alkanols

(i)Isomers of propanol C_3H_7OH

$CH_3CH_2CH_2OH$ - Propan-1-ol

CH_3CHCH_3 - Propan-2-ol

Propan-2-ol and Propan-1-ol are position isomers because only the position of the –OH functional group changes.

(ii)Isomers of Butanol C_4H_9OH

$CH_3\ CH_2\ CH_3\ CH_2\ OH$ Butan-1-ol

$CH_3\ CH_2\ CH\ CH_3$ Butan-2-ol

2-methylpropan-2-ol

Butan-2-ol and Butan-1-ol are position isomers because only the position of the -OH functional group changes.

2-methylpropan-2-ol is both a structural and position isomers because both the position of the functional group and the arrangement of the atoms in the molecule changes.

(iii)Isomers of Pentanol $C_5H_{11}OH$

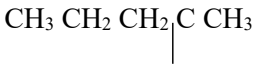
$CH_3\ CH_2\ CH_2CH_2CH_2\ OH$ Pentan-1-ol (Position isomer)

$CH_3\ CH_2\ CH\ CH_3$ Pentan-2-ol (Position isomer)

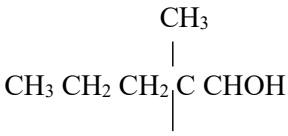
$CH_3\ CH_2\ CH\ CH_2\ CH_3$ Pentan-3-ol (Position isomer)

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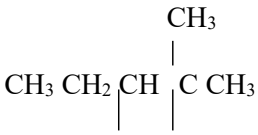
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2-methylbutan-2-ol (Position /structural isomer)

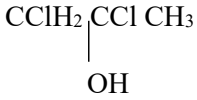


2,2-dimethylbutan-1-ol (Position /structural isomer)

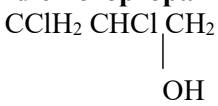


2,3-dimethylbutan-1-ol (Position /structural isomer)

(iv)1,2-dichloropropan-2-ol

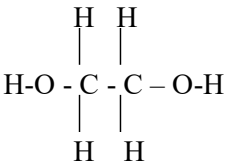


(v)1,2-dichloropropan-1-ol



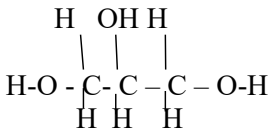
(vi) Ethan1,2-diol

$HOCH_2CH_2OH$



(vii) Propan1,2,3-triol

$HOCH_2CHOHCH_2OH$



C. LABORATORY PREPARATION OF ALKANOLS.

For decades the world over, people have been fermenting grapes juice, sugar, carbohydrates and starch to produce ethanol as a social drug for relaxation.

In large amount, drinking of ethanol by mammals /human beings causes mental and physical lack of coordination.

Prolonged intake of ethanol causes permanent mental and physical lack of coordination because it damages vital organs like the liver.

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Fermentation is the reaction where sugar is converted to alcohol/alkanol using biological catalyst/enzymes in **yeast.**
It involves **three** processes:
(i)Conversion of starch to maltose using the enzyme **diastase.**

$(C_6H_{10}O_5)n\ (s) + H_2O(l)$ --diastase enzyme --> $C_{12}H_{22}O_{11}(aq)$
(Starch) (Maltose)

(ii)Hydrolysis of Maltose to glucose using the enzyme **maltase.**

$C_{12}H_{22}O_{11}(aq) + H_2O(l)$ -- maltase enzyme -->2 $C_6H_{12}O_6(aq)$
(Maltose) (glucose)

(iii)Conversion of glucose to ethanol and carbon(IV)oxide gas using the enzyme **zymase.**

$C_6H_{12}O_6(aq)$ -- zymase enzyme --> 2 $C_2H_5OH(aq)$ + $2CO_2(g)$
(glucose) (Ethanol)

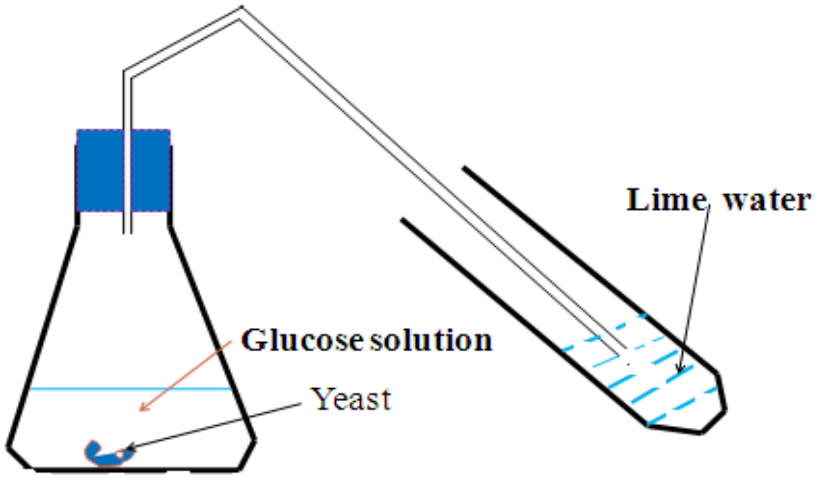
At concentration greater than 15% by volume, the ethanol produced kills the yeast enzyme stopping the reaction.
To increases the concentration, fractional distillation is done to produce spirits (e.g. Brandy=40% ethanol).
Methanol is much more poisonous /toxic than ethanol.
Taken large quantity in small quantity it causes instant blindness and liver, killing the consumer victim within hours.

School laboratory preparation of ethanol from fermentation of glucose

Measure 100cm3 of pure water into a conical flask.
Add about five spatula end full of glucose.
Stir the mixture to dissolve.
Add about one spatula end full of yeast.
Set up the apparatus as below.

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Preserve the mixture for about **three** days.

D.PHYSICAL AND CHEMICAL PROPERTIES OF ALKANOLS

Use the prepared sample above for the following experiments that shows the characteristic properties of alkanols

(a)Role of yeast

Yeast is a single cell fungus which contains the enzyme maltase and zymase that catalyse the fermentation process.

(b) Observations in lime water.

A white precipitate is formed that dissolve to a colourless solution later. Lime water/Calcium hydroxide reacts with carbon(IV)0xide produced during the fermentation to form insoluble calcium carbonate and water.

More carbon (IV)0xide produced during fermentation react with the insoluble calcium carbonate and water to form soluble calcium hydrogen carbonate.

$$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s)$$

$$H_2O(l) + CO_2(g) + CaCO_3(s) \rightarrow Ca(HCO_3)_2(aq)$$

(c)Effects on litmus paper

Experiment

Take the prepared sample and test with both blue and red litmus papers.
Repeat the same with pure ethanol and methylated spirit.
Sample Observation table

Substance/alkanol	Effect on litmus paper
Prepared sample	Blue litmus paper remain blue Red litmus paper remain red
Absolute ethanol	Blue litmus paper remain blue

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	Red litmus paper remain red
Methylated spirit	Blue litmus paper remain blue Red litmus paper remain red

Explanation

Alkanols are neutral compounds/solution that have characteristic sweet smell and taste.

They have no effect on both blue and red litmus papers.

(d)Solubility in water.

Experiment

Place about 5cm3 of prepared sample into a clean test tube Add equal amount of distilled water.

Repeat the same with pure ethanol and methylated spirit.

Observation

No layers formed between the two liquids.

Explanation

Ethanol is **miscible** in water.Both ethanol and water are polar compounds .

The solubility of alkanols decrease with increase in the alkyl chain/molecular mass.

The alkyl group is insoluble in water while –OH functional group is soluble in water.

As the molecular chain becomes **longer** ,the effect of the **alkyl** group **increases** as the effect of the functional group **decreases**.

e)Melting/boiling point.

Experiment

Place pure ethanol in a long boiling tube .Determine its boiling point.

Observation

Pure ethanol has a boiling point of 78°C at sea level/one atmosphere pressure.

Explanation

The melting and boiling point of alkanols increase with increase in molecular chain/mass .

This is because the intermolecular/van-der-waals forces of attraction between the molecules increase.

More heat energy is thus required to weaken the longer chain during melting and break during boiling.

f)Density

Density of alkanols increase with increase in the intermolecular/van-der-waals forces of attraction between the molecule, making it very close to each other.

This reduces the volume occupied by the molecule and thus increase the their mass per unit volume (density).

Summary table showing the trend in physical properties of alkanols

Alkanol	Melting point (^{o}C)	Boiling point (^{o}C)	Density gcm^{-3}	Solubility in water

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Methanol	-98	65	0.791	soluble
Ethanol	-117	78	0.789	soluble
Propanol	-103	97	0.803	soluble
Butanol	-89	117	0.810	Slightly soluble
Pentanol	-78	138	0.814	Slightly soluble
Hexanol	-52	157	0.815	Slightly soluble
Heptanol	-34	176	0.822	Slightly soluble
Octanol	-15	195	0.824	Slightly soluble
Nonanol	-7	212	0.827	Slightly soluble
Decanol	6	228	0.827	Slightly soluble

g)Burning

Experiment

Place the prepared sample in a watch glass. Ignite. Repeat with pure ethanol and methylated spirit.

Observation/Explanation

Fermentation produce ethanol with a lot of water(about a ratio of 1:3)which prevent the alcohol from igniting.

Pure ethanol and methylated spirit easily catch fire / highly flammable.

They burn with an almost colourless non-sooty/non-smoky **blue** flame to form **carbon(IV) oxide** (in excess air/oxygen)or **carbon(II) oxide** (limited air) and **water**.

Ethanol is thus a **saturated** compound like alkanes.

Chemica equation

$C_2H_5OH(l) + 3O_2(g) \rightarrow 3H_2O(l) + 2CO_2(g)$ (excess air)

$C_2H_5OH(l) + 2O_2(g) \rightarrow 3H_2O(l) + 2CO(g)$ (limited air)

$2CH_3OH(l) + 3O_2(g) \rightarrow 4H_2O(l) + 2CO_2(g)$ (excess air)

$2\ CH_3OH(l) + 2O_2(g) \rightarrow 4H_2O(l) + 2CO(g)$ (limited air)

$2C_3H_7OH(l) + 9O_2(g) \rightarrow 8H_2O(l) + 6CO_2(g)$ (excess air)

$C_3H_7OH(l) + 3O_2(g) \rightarrow 4H_2O(l) + 3CO(g)$ (limited air)

$2C_4H_9OH(l) + 13O_2(g) \rightarrow 20H_2O(l) + 8CO_2(g)$ (excess air)

$C_4H_9OH(l) + 3O_2(g) \rightarrow 4H_2O(l) + 3CO(g)$ (limited air)

Due to its flammability, ethanol is used;

(i) as a fuel in spirit lamps

(ii) as gasohol when blended with gasoline

(h)Formation of alkoxides

Experiment

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Cut a very small piece of sodium. Put it in a beaker containing about 20cm3 of the prepared sample in a beaker.

Test the products with litmus papers. Repeat with absolute ethanol and methylated spirit.

Sample observations

Substance/alkanol	Effect of adding sodium
Fermentation prepared sample	(i)effervescence/fizzing/bubbles (ii)colourless gas produced that extinguish burning splint with explosion/ "Pop" sound (iii)colourless solution formed (iv)blue litmus papers remain blue (v)red litmus papers turn blue
Pure/absolute ethanol/methylated spirit	(i)**slow** effervescence/fizzing/bubbles (ii)colourless gas **slowly** produced that extinguish burning splint with explosion/ "Pop" sound (iii)colourless solution formed (iv)blue litmus papers remain blue (v)red litmus papers turn blue

Explanations

Sodium/potassium reacts slowly with alkanols to form basic solution called **alkoxides** and producing **hydrogen** gas.

If the alkanol has some water the metals react faster with the water to form **soluble hydroxides/alkalis** i.e.

Sodium + Alkanol -> Sodium **alk**oxides + Hydrogen gas

Potassium + Alkanol -> Potassium **alk**oxides + Hydrogen gas

Sodium + Water -> Sodium **hydr**oxides + Hydrogen gas

Potassium + Water -> Potassium **hydr**oxides + Hydrogen gas

Examples

1.Sodium metal reacts with ethanol to form sodium **eth**oxide

Sodium metal reacts with water to form sodium **Hydr**oxide

$2CH_3CH_2OH(l) + 2Na(s) \rightarrow 2CH_3CH_2ONa\ (aq) + H_2\ (s)$

$2H_2O(l) + 2Na(s) \rightarrow 2NaOH\ (aq) + H_2\ (s)$

2.Potassium metal reacts with ethanol to form Potassium **eth**oxide

Potassium metal reacts with water to form Potassium **Hydr**oxide

$2CH_3CH_2OH(l) + 2K(s) \rightarrow 2CH_3CH_2OK\ (aq) + H_2\ (s)$

$2H_2O(l) + 2K(s) \rightarrow 2KOH\ (aq) + H_2\ (s)$

3.Sodium metal reacts with propanol to form sodium **prop**oxide

Sodium metal reacts with water to form sodium **Hydr**oxide

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$2CH_3CH_2\ CH_2OH(l) + 2Na(s) \rightarrow 2CH_3CH_2\ CH_2ONa\ (aq) + H_2\ (s)$
$2H_2O(l) + 2Na(s) \rightarrow 2NaOH\ (aq) + H_2\ (s)$

4.Potassium metal reacts with propanol to form Potassium **prop**oxide
Potassium metal reacts with water to form Potassium **Hydr**oxide
$2CH_3CH_2\ CH_2OH(l) + 2K(s) \rightarrow 2CH_3CH_2\ CH_2OK\ (aq) + H_2\ (s)$
$2H_2O(l) + 2K(s) \rightarrow 2KOH\ (aq) + H_2\ (s)$

5.Sodium metal reacts with butanol to form sodium **but**oxide
Sodium metal reacts with water to form sodium **Hydr**oxide
$2CH_3CH_2\ CH_2\ CH_2OH(l) + 2Na(s) \rightarrow 2CH_3CH_2\ CH_2\ CH_2ONa\ (aq) + H_2\ (s)$
$2H_2O(l) + 2Na(s) \rightarrow 2NaOH\ (aq) + H_2\ (s)$

6.Sodium metal reacts with pentanol to form sodium **pent**oxide
Sodium metal reacts with water to form sodium **Hydr**oxide
$2CH_3CH_2\ CH_2\ CH_2\ CH_2OH(l) + 2Na(s) \rightarrow 2CH_3CH_2\ CH_2\ CH_2\ CH_2ONa\ (aq) + H_2\ (s)$
$2H_2O(l) + 2Na(s) \rightarrow 2NaOH\ (aq) + H_2\ (s)$

(i)Formation of Esters/Esterification

Experiment
Place 2cm3 of ethanol in a boiling tube.
Add equal amount of ethanoic acid.To the mixture add carefully 2drops of concentrated sulphuric(VI)acid.
Warm/Heat gently.
Pour the mixture into a beaker containing about 50cm3 of cold water.
Smell the products.
Repeat with methanol
Sample observations

Substance/alkanol	**Effect on adding equal amount of ethanol/concentrated sulphuric(VI)acid**
Absolute ethanol	Sweet fruity smell
Methanol	Sweet fruity smell

Explanation
Alkanols react with alkanoic acids to form a group of homologous series of sweet smelling compounds called esters and water. This reaction is catalyzed by concentrated sulphuric(VI)acid in the laboratory.

Alkanol + Alkanoic acid $\xrightarrow{\text{Conc. } H_2SO_4}$ Ester + water

Naturally esterification is catalyzed by sunlight. Each ester has a characteristic smell derived from the many possible combinations of alkanols and alkanoic acids that create a variety of known natural(mostly in fruits) and synthetic(mostly in juices) esters .
Esters derive their names from the alkanol first then alkanoic acids. The alkanol "becomes" an **alkyl** group and the alkanoic acid "becomes" **alkanoate** hence **alkylalkanoate**. e.g.

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Ethanol	+	Ethanoic acid	->	Ethylethanoate	+	Water
Ethanol	+	Propanoic acid	->	Ethylpropanoate	+	Water
Ethanol	+	Methanoic acid	->	Ethylmethanoate	+	Water
Ethanol	+	butanoic acid	->	Ethylbutanoate	+	Water
Propanol	+	Ethanoic acid	->	Propylethanoate	+	Water
Methanol	+	Ethanoic acid	->	Methyethanoate	+	Water
Methanol	+	Decanoic acid	->	Methyldecanoate	+	Water
Decanol	+	Methanoic acid	->	Decylmethanoate	+	Water

During the formation of the ester, the "O" joining the alkanol and alkanoic acid comes from the alkanol.

R_1 -CO**OH** + R_2 –O**H** -> **R_1 -COO –R_2** + **H_2O**

e.g.

1. Ethanol reacts with ethanoic acid to form the ester ethyl ethanoate and water.

Ethanol + Ethanoic acid --Conc. H_2SO_4 -->Ethylethanoate + Water

C_2H_5OH (l) + CH_3COOH(l) --Conc. H_2SO_4 --> $CH_3COO\ C_2H_5$(**aq**) +H_2O(l)

CH_3CH_2OH (l)+ CH_3COOH(l) --Conc. H_2SO_4 --> $CH_3COOCH_2CH_3$(**aq**) +H_2O(l)

2. Ethanol reacts with propanoic acid to form the ester ethylpropanoate and water.

Ethanol + Propanoic acid --Conc. H_2SO_4 -->Ethylethanoate + Water

C_2H_5OH (l)+ $CH_3\ CH_2COOH$(l) --Conc. H_2SO_4 -->$CH_3CH_2COO\ C_2H_5$(**aq**) +H_2O(l)

CH_3CH_2OH (l)+ $CH_3\ CH_2COOH$(l) --Conc. H_2SO_4 --> $CH_3\ CH_2COOCH_2CH_3$(**aq**) +H_2O(l)

3. Methanol reacts with ethanoic acid to form the ester methyl ethanoate and water.

Methanol + Ethanoic acid --Conc. H_2SO_4 -->Methylethanoate + Water

CH_3OH (l) + CH_3COOH(l) --Conc. H_2SO_4 --> $CH_3COO\ CH_3$(**aq**) +H_2O(l)

4. Methanol reacts with propanoic acid to form the ester methyl propanoate and water.

Methanol + propanoic acid --Conc. H_2SO_4 -->Methylpropanoate + Water

CH_3OH (l)+ $CH_3\ CH_2COOH$(l) --Conc. H_2SO_4 --> $CH_3\ CH_2COO\ CH_3$(**aq**) +H_2O(l)

5. Propanol reacts with propanoic acid to form the ester propylpropanoate and water.

Propanol + Propanoic acid --Conc. H_2SO_4 -->Ethylethanoate + Water

C_3H_7OH (l)+ $CH_3\ CH_2COOH$(l) --Conc. H_2SO_4 -->$CH_3CH_2COO\ C_3H_7$(**aq**) +H_2O(l)

$CH_3CH_2\ CH_2OH$ (l)+ $CH_3\ CH_2COOH$(l) --Conc. H_2SO_4 --> $CH_3\ CH_2COOCH_2\ CH_2CH_3$(**aq**) +$H_2O$(l)

(j)Oxidation

Experiment

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Place 5cm3 of absolute ethanol in a test tube.Add three drops of acidified potassium manganate(VII).Shake thoroughly for one minute/warm.Test the solution mixture using pH paper. Repeat by adding acidified potassium dichromate(VII).

Sample observation table

Substance/alkanol	Adding acidified $KMnO_4/K_2Cr_2O_7$	pH of resulting solution/mixture	Nature of resulting solution/mixture
Pure ethanol	(i)Purple colour of $KMnO_4$decolorized (ii) Orange colour of $K_2Cr_2O_7$turns green.	pH= 4/5/6 pH = 4/5/6	Weakly acidic Weakly acidic

Explanation

Both acidified $KMnO_4$ and $K_2Cr_2O_7$ are oxidizing agents(add oxygen to other compounds. They oxidize alkan**o**ls to a group of homologous series called alkan**a**ls then further oxidize them to alkan**oic** acids.The oxidizing agents are themselves reduced hence changing their colour:

(i) Purple $KMnO_4$ is reduced to colourless Mn^{2+}

(ii)Orange $K_2Cr_2O_7$is reduced to green Cr^{3+}

The pH of alkanoic acids show they have few H^+ because they are weak acids i.e

Alkanol + [O] -> Alkanal + [O] -> alkanoic acid

NB The [O] comes from the oxidizing agents acidified $KMnO_4$ or $K_2Cr_2O_7$

Examples

1.When ethanol is warmed with three drops of acidified $KMnO_4$ there is decolorization of $KMnO_4$

Ethanol + [O] -> Ethanal + [O] -> Ethanoic acid

$CH_3CH_2OH + [O] \rightarrow CH_3CH_2O + [O] \rightarrow CH_3COOH$

2.When methanol is warmed with three drops of acidified $K_2Cr_2O_7$,the orange colour of acidified $K_2Cr_2O_7$ changes to green.

methanol + [O] -> methanal + [O] -> methanoic acid

$CH_3OH + [O] \rightarrow CH_3O + [O] \rightarrow HCOOH$

3.When propanol is warmed with three drops of acidified $K_2Cr_2O_7$,the orange colour of acidified $K_2Cr_2O_7$ changes to green.

Propanol + [O] -> Propanal + [O] -> Propanoic acid

$CH_3CH_2\ CH_2OH + [O] \rightarrow CH_3CH_2\ CH_2O + [O] \rightarrow CH_3\ CH_2COOH$

4.When butanol is warmed with three drops of acidified $K_2Cr_2O_7$,the orange colour of acidified $K_2Cr_2O_7$ changes to green.

Butanol + [O] -> Butanal + [O] -> Butanoic acid

$CH_3CH_2\ CH_2\ CH_2OH + [O] \rightarrow CH_3CH_2\ CH_2CH_2O + [O] \rightarrow CH_3\ CH_2COOH$

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Air slowly oxidizes ethanol to dilute ethanoic acid commonly called **vinegar.** If beer is not tightly corked, a lot of carbon(IV)oxide escapes and there is slow oxidation of the beer making it "flat".

(k)Hydrolysis /Hydration and Dehydration

I. Hydrolysis/Hydration is the reaction of a compound/substance with water.

Alkenes react with water vapour/steam at high temperatures and high pressures in presence of phosphoric acid catalyst to form alkanols.i.e.

Alkenes + Water - H_3PO_4 catalyst-> Alkanol

Examples

(i)Ethene is mixed with steam over a phosphoric acid catalyst at 300°C temperature and 60 atmosphere pressure to form ethanol

Ethene + water ---60 atm/300°C/ H_3PO_4 --> Ethanol

$H_2C=CH_2$ (g) + $\mathbf{H_2O}$(l) --60 atm/300°C/ H_3PO_4 --> $CH_3C\mathbf{H_2O}H$(l)

This is the main method of producing large quantities of ethanol instead of fermentation

(ii) Propene + water ---60 atm/300°C/ H_3PO_4 --> Propanol

$CH_3C=CH_2$ (g) + $\mathbf{H_2O}$(l) --60 atm/300°C/ H_3PO_4 --> $CH_3CH_2C\mathbf{H_2O}H$(l)

(iii) Butene + water ---60 atm/300°C/ H_3PO_4 --> Butanol

$CH_3CH_2C=CH_2$ (g) + $\mathbf{H_2O}$(l) --60 atm/300°C/ H_3PO_4 --> $CH_3CH_2CH_2C\mathbf{H_2O}H$(l)

II. Dehydration is the process which concentrated sulphuric(VI)acid (**dehydrating agent**) removes water from a compound/substances.

Concentrated sulphuric(VI)acid dehydrates alkanols to the corresponding alkenes at about 180°C. i.e

Alkanol --Conc. H_2SO_4/180°C--> Alkene + Water

Examples

1. At 180°C and in presence of Concentrated sulphuric(VI)acid, ethanol undergoes dehydration to form ethene.

Ethanol ---180°C/ H_2SO_4 --> Ethene + Water

$CH_3C\mathbf{H_2O}H$(l) --180°C/ H_2SO_4 --> $H_2C=CH_2$ (g) + $\mathbf{H_2O}$(l)

2. Propanol undergoes dehydration to form propene.

Propanol ---180°C/ H_2SO_4 --> Propene + Water

$CH_3CH_2C\mathbf{H_2O}H$(l) --180°C/ H_2SO_4 --> $CH_3CH=CH_2$ (g) + $\mathbf{H_2O}$(l)

3. Butanol undergoes dehydration to form Butene.

Butanol ---180°C/ H_2SO_4 --> Butene + Water

$CH_3CH_2CH_2C\mathbf{H_2O}H$(l) --180°C/ H_2SO_4 --> $CH_3CH_2C=CH_2$ (g) + $\mathbf{H_2O}$(l)

3. Pentanol undergoes dehydration to form Pentene.

Pentanol ---180°C/ H_2SO_4 --> Pentene + Water

$CH_3CH_2CH_2CH_2C\mathbf{H_2O}H$(l)--180°C/ H_2SO_4-->$CH_3CH_2CH_2C=CH_2$ (g)+$\mathbf{H_2O}$(l)

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(l)Similarities of alkanols with Hydrocarbons

I. Similarity with alkanes

Both alkanols and alkanes burn with a **blue non-sooty flame** to form carbon(IV)oxide(in excess air/oxygen)/carbon(II)oxide(in limited air) and water. This shows they are saturated with high C:H ratio. e.g.

Both ethanol and ethane ignite and burns in air with a **blue non-sooty flame** to form carbon(IV)oxide(in excess air/oxygen)/carbon(II)oxide(in limited air) and water.

$CH_2\ CH_2OH(l) + 3O_2(g)$ -Excess air-> $2CO_2(g) + 3H_2O(l)$

$CH_2\ CH_2OH(l) + 2O_2(g)$ -Limited air-> $2CO(g) + 3H_2O(l)$

$CH_3\ CH_3(g) + 3O_2(g)$ -Excess air-> $2CO_2(g) + 3H_2O(l)$

$2CH_3\ CH_3(g) + 5O_2(g)$ -Limited air-> $4CO(g) + 6H_2O(l)$

II. Similarity with alkenes/alkynes

Both alkanols(R-OH) and alkenes/alkynes(with = C = C = double and – C = C- triple) bond:

(i)decolorize acidified $KMnO_4$

(ii)turns Orange acidified $K_2Cr_2O_7$ to green.

Alkan**ol**s(R-**OH**) are oxidized to alkan**a**ls(R-**O**) ant then alkan**oic** acids(R-**OOH**).

Alkenes are oxidized to alkanols with duo/double functional groups.

Examples

1.When ethanol is warmed with three drops of acidified $K_2Cr_2O_7$ the orange of acidified $K_2Cr_2O_7$ turns to green. Ethanol is oxidized to ethanol and then to ethanoic acid.

Ethan**ol** + [O] -> Ethan**a**l + [O] -> Ethan**oic** acid

CH_3CH_2OH + [O] -> CH_3CH_2O + [O] -> CH_3COOH

2.When ethene is bubbled in a test tube containing acidified $K_2Cr_2O_7$,the orange of acidified $K_2Cr_2O_7$ turns to green. Ethene is oxidized to ethan-1,2-diol.

Ethene + [O] -> Ethan-1,2-diol.

$H_2C=CH_2$ + [O] -> $HOCH_2\text{-}CH_2OH$

III. Differences with alkenes/alkynes

Alkanols do not decolorize bromine and chlorine water.

Alkenes decolorizes bromine and chlorine water to form halogenoalkanols

Example

When ethene is bubbled in a test tube containing bromine water,the bromine water is decolorized. Ethene is oxidized to bromoethanol.

Ethene + Bromine water -> Bromoethanol.

$H_2C=CH_2$ + HOBr -> $BrCH_2\text{-}CH_2OH$

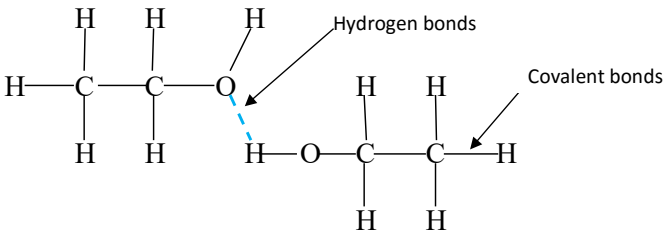
IV. Differences in melting and boiling point with Hydrocarbons

Alkanos have higher melting point than the corresponding hydrocarbon (alkane/alkene/alkyne)

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This is because most alkanols exist as **dimer.**A dimer is a molecule made up of two other molecules joined usually by van-der-waals forces/hydrogen bond or dative bonding.
Two alkanol molecules form a dimer joined by hydrogen bonding.
Example
In Ethanol the oxygen atom attracts/pulls the shared electrons in the covalent bond more to itself than Hydrogen.
This creates a partial negative charge ($^{\delta-}$) on oxygen and partial positive charge($^{\delta+}$) on hydrogen.
Two ethanol molecules attract each other at the partial charges through Hydrogen bonding forming a **dimme**r.



Dimerization of alkanols means more energy is needed to break/weaken the Hydrogen bonds before breaking/weakening the intermolecular forces joining the molecules of all organic compounds during boiling/melting.
E.**USES OF SOME ALKANOLS**
(a)Methanol is used as industrial alcohol and making methylated spirit
(b)Ethanol is used:

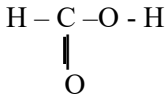
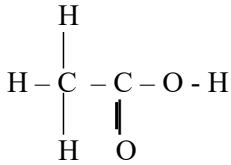
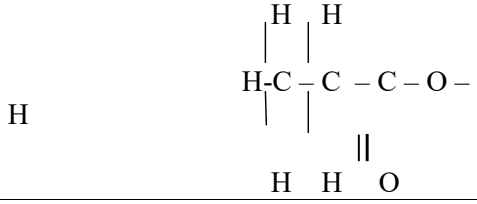
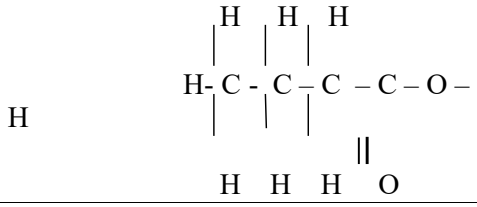
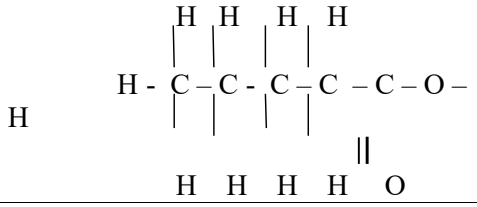
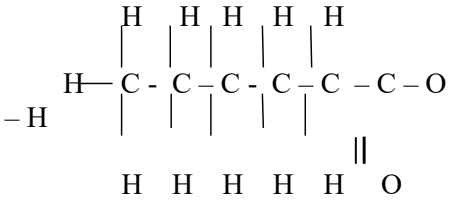
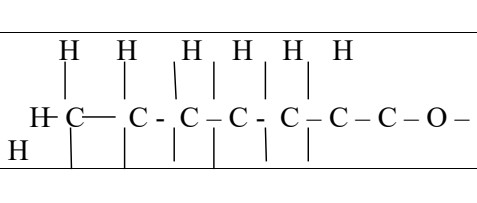
1. as alcohol in alcoholic drinks e.g Beer, wines and spirits.
2. as antiseptic to wash woulds
3. in manufacture of vanishes, ink ,glue and paint because it is volatile and thus easily evaporate
4. as a fuel when blended with petrol to make gasohol.

B.ALKANOIC ACIDS (Carboxylic acids)
(A) INTRODUCTION.

Alkanoic acids belong to a homologous series of organic compounds with a general formula $\mathbf{C_nH_{2n+1}COOH}$ and thus **-COOH** as the functional group .The 1st ten alkanoic acids include:

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n	General /molecular formular	Structural formula	IUPAC name
0	HCOOH		Methanoic acid
1	CH_3 COOH		Ethanoic acid
2	CH_3 CH_2 COOH C_2 H_5 COOH		Propanoic acid
3	CH_3 CH_2 CH_2 COOH C_3 H_7 COOH		Butanoic acid
4	$CH_3CH_2CH_2CH_2$ COOH C_4 H_9 COOH		Pentanoic acid
5	CH_3CH_2 $CH_2CH_2CH_2$ COOH C_5 H_{11} COOH		Hexanoic acid
6	CH_3CH_2 CH_2 $CH_2CH_2CH_2$ COOH C_6 H_{13} COOH		Pentanoic acid

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		H H H H H H O (C=O)	

Alkanoic acids like alkanols /alkanes/alkenes/alkynes form a homologous series where:

(i)the general name of an alkanoic acids is derived from the alkane name then ending with "**–oic**" acid as the table above shows.

(ii) the members have R-**COOH**/R-C(=O)-O-H as the functional group.

(iii)they have the same general formula represented by R-COOH where R is an alkyl group.

(iv)each member differ by $–CH_2-$ group from the next/previous.

(v)they show a similar and gradual change in their physical properties e.g. boiling and melting point.

(vi)they show similar and gradual change in their chemical properties.

(vii) since they are acids they show similar properties with mineral acids.

(B) ISOMERS OF ALKANOIC ACIDS.

lkanoic acids exhibit both structural and position isomerism. The isomers are named by using the following basic guidelines

(i)Like alkanes. identify the longest carbon chain to be the parent name.

(ii)Identify the position of the -C(=O)-O-H functional group to give it the smallest /lowest position.

(iii)Identify the type and position of the side group branches.

Practice examples on isomers of alkanoic acids

1.Isomers of butanoic acid C_3H_7COOH

$CH_3\ CH_2\ CH_2\ COOH$

Butan-1-oic acid

$H_2C(CH_3)-C-COOH$ 2-methylpropan-1-oic acid

2-methylpropan-1-oic acid and Butan-1-oic acid are structural isomers because the position of the functional group does not change but the arrangement of the atoms in the molecule does.

2.Isomers of pentanoic acid C_4H_9COOH

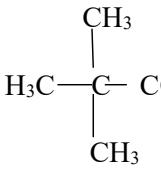
$CH_3CH_2CH_2CH_2\ COOH$ pentan-1-oic acid

CH_3

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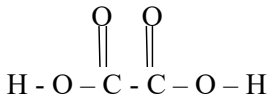
$CH_3CH_2CH\ COOH$ 2-methylbutan-1-oic acid



2,2-dimethylpropan-1-oic acid

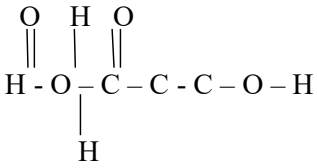
3.Ethan-1,2-dioic acid

HOOC- COOH //



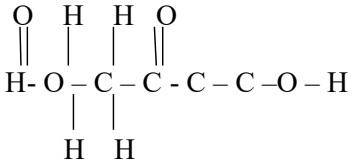
4.Propan-1,3-dioic acid

HOOC- CH_2COOH //



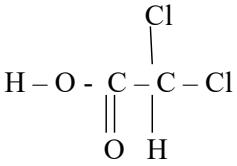
5.Butan-1,4-dioic acid

$HOOC\ CH_2\ CH_2\ COOH$



6.2,2-dichloroethan-1,2-dioic acid

$HOOCCHCl_2$



(C) LABORATORY AND INDUSTRIAL PREPARATIONOF ALKANOIC ACIDS.

In a school laboratory, alkanoic acids can be prepared by adding an oxidizing agent ($H^+/KMnO_4$ or $H^+/K_2Cr_2O_7$)to the corresponding alkanol then warming. The oxidation converts the alkanol first to an alkanal the alkanoic acid.

NB Acidified $KMnO_4$ is a stronger oxidizing agent than acidified $K_2Cr_2O_7$

General equation:

R- CH_2 – **OH** + [O] $\xrightarrow{H^+/KMnO_4}$ R- CH –**O** + $H_2O(l)$

(alkanol) (alkanal)

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R- CH – **O** + [O] --H^+/$KMnO_4$--> R- C –**OOH**
(alkanal) (alkanoic acid)

Examples

1.Ethan**o**l on warming in acidified $KMnO_4$ is oxidized to ethan**a**l then ethan**oic** acid .

CH_3- CH_2 – **OH** + [O] --H^+/$KMnO_4$--> CH_3- CH –**O** + H_2O(l)
(ethanol) (ethanal)

CH_3- CH – **O** + [O] --H^+/$KMnO_4$--> CH_3- C –**OOH**
(ethanal) (ethanoic acid)

2Propan**o**l on warming in acidified $KMnO_4$ is oxidized to propan**a**l then propan**oic** acid

CH_3- CH_2 CH_2 – **OH** + [O] --H^+/$KMnO_4$--> CH_3- CH_2 CH –**O** + H_2O(l)
(propanol) (propanal)

CH_3- CH – **O** + [O] --H^+/$KMnO_4$--> CH_3- C –**OOH**
(propanal) (propanoic acid)

Industrially,large scale manufacture of alkanoic acid like ethanoic acid is obtained from:

(a)Alk**e**nes reacting with steam at high temperatures and pressure in presence of phosphoric(V)acid catalyst and undergo hydrolysis to form alkanols. i.e.

Alkenes + Steam/water -- H_2PO_4 Catalyst--> Alkanol

The alkanol is then oxidized by air at 5 atmosphere pressure with Manganese (II)sulphate(VI) catalyst to form the alkanoic acid.

Alkanol + Air -- $MnSO_4$ Catalyst/5 atm pressure--> Alkanoic acid

Example

Ethene is mixed with steam over a phosphoric(V)acid catalyst,300oC temperature and 60 atmosphere pressure to form ethanol.

CH_2=CH_2 + H_2O -> CH_3 CH_2OH
(Ethene) (Ethanol)

This is the industrial large scale method of manufacturing ethanol

Ethanol is then oxidized by air at 5 atmosphere pressure with Manganese (II)sulphate(VI) catalyst to form the ethanoic acid.

CH_3 CH_2OH + [O] -- $MnSO_4$ Catalyst/5 atm pressure--> CH_3 COOH
(Ethanol) (Ethanoic acid)

(b)Alk**y**nes react with liquid water at high temperatures and pressure in presence of Mercury(II)sulphate(VI)catalyst and 30% concentrated sulphuric(VI)acid to form alkan**a**ls.

Alk**y**ne + Water -- Mercury(II)sulphate(VI)catalyst--> Alkan**a**l

The alkan**a**l is then oxidized by air at 5 atmosphere pressure with Manganese (II) sulphate(VI) catalyst to form the alkan**oic** acid.

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Alkan**al** + air/oxygen -- Manganese(II)sulphate(VI)catalyst--> Alkan**oic acid**

Example

Eth**y**ne react with liquid water at high temperature and pressure with Mercury (II) sulphate (VI)catalyst and 30% concentrated sulphuric(VI)acid to form ethan**a**l.

$CH \equiv CH + H_2O \xrightarrow{HgSO_4} CH_3CH_2O$

(Ethyne) (Ethan**a**l)

This is another industrial large scale method of manufacturing ethanol from large quantities of ethyne found in natural gas.

Ethan**a**l is then oxidized by air at 5 atmosphere pressure with Manganese (II)sulphate(VI) catalyst to form the ethanoic acid.

$CH_3CH_2O + [O] \xrightarrow{MnSO_4\ Catalyst/5\ atm\ pressure} CH_3COOH$

(Ethan**a**l) (Oxygen from air) (Ethanoic acid)

(D) PHYSICAL AND CHEMICAL PROPERTIES OF ALKANOIC ACIDS.

I.Physical properties of alkanoic acids

The table below shows some physical properties of alkanoic acids

Alkanol	Melting point(°C)	Boiling point(°C)	Density(gcm^{-3})	Solubility in water
Methanoic acid	18.4	101	1.22	soluble
Ethanoic acid	16.6	118	1.05	soluble
Propanoic acid	-2.8	141	0.992	soluble
Butanoic acid	-8.0	164	0.964	soluble
Pentanoic acid	-9.0	187	0.939	Slightly soluble
Hexanoic acid	-11	205	0.927	Slightly soluble
Heptanoic acid	-3	223	0.920	Slightly soluble
Octanoic acid	11	239	0.910	Slightly soluble
Nonanoic acid	16	253	0.907	Slightly soluble
Decanoic acid	31	269	0.905	Slightly soluble

From the table note the following:

(i) Melting and boiling point decrease as the carbon chain increases due to increase in intermolecular forces of attraction between the molecules requiring more energy to separate the molecules.

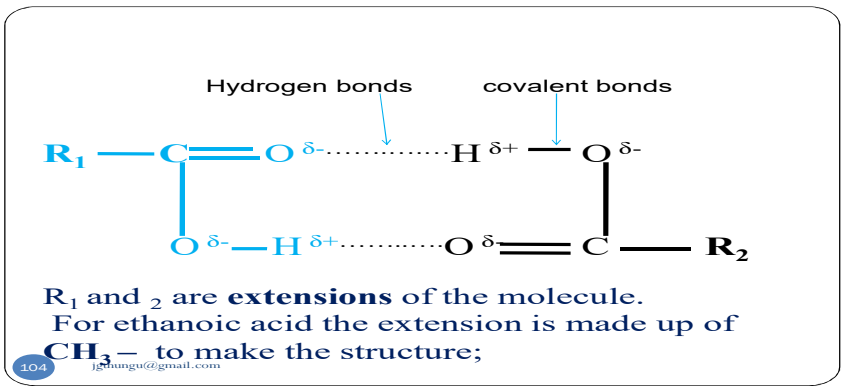
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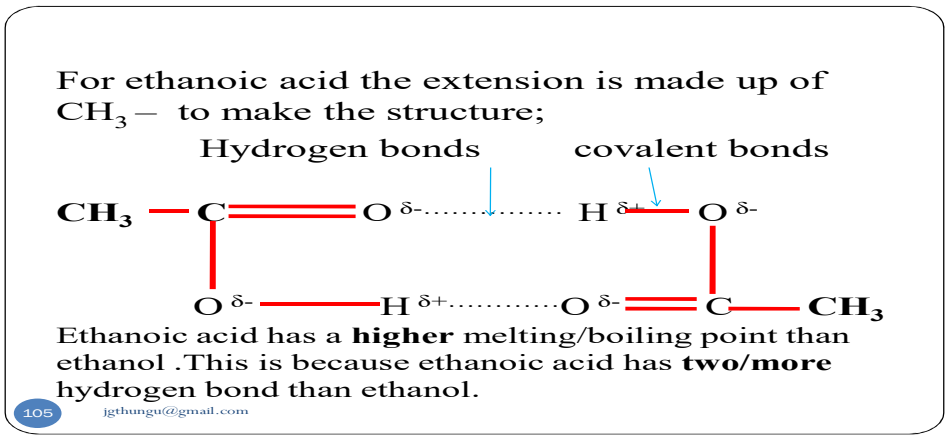
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(ii) The density decreases as the carbon chain increases as the intermolecular forces of attraction increases between the molecules making the molecule very close reducing their volume in unit mass.

(iii) Solubility decreases as the carbon chain increases as the soluble –COOH end is shielded by increasing insoluble alkyl/hydrocarbon chain.

(iv) Like alkanols ,alkanoic acids exist as dimmers due to the hydrogen bonds within the molecule. i.e..





II Chemical properties of alkanoic acids

The following experiments shows the main chemical properties of ethanoic (alkanoic) acid.

(a)Effect on litmus papers

Experiment

Dip both blue and red litmus papers in ethanoic acid. Repeat with a solution of succinic acid, citric acid, oxalic acid, tartaric acid and dilute nitric(V)acid.

Sample observations

Solution/acid	Observations/effect on litmus papers	Inference

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Ethanoic acid	Blue litmus paper turn red Red litmus paper remain red	H_3O^+/H^+(aq)ion
Succinic acid	Blue litmus paper turn red Red litmus paper remain red	H_3O^+/H^+(aq)ion
Citric acid	Blue litmus paper turn red Red litmus paper remain red	H_3O^+/H^+(aq)ion
Oxalic acid	Blue litmus paper turn red Red litmus paper remain red	H_3O^+/H^+(aq)ion
Tartaric acid	Blue litmus paper turn red Red litmus paper remain red	H_3O^+/H^+(aq)ion
Nitric(V)acid	Blue litmus paper turn red Red litmus paper remain red	H_3O^+/H^+(aq)ion

Explanation

All acidic solutions contains H^+/H_3O^+(aq) ions. The H^+ /H_3O^+ (aq) ions is responsible for turning blue litmus paper/solution to red

(b)pH

Experiment

Place 2cm3 of ethaoic acid in a test tube. Add 2 drops of universal indicator solution and determine its pH. Repeat with a solution of succinic acid, citric acid, oxalic acid, tartaric acid and dilute sulphuric (VI)acid.

Sample observations

Solution/acid	pH	Inference
Ethanoic acid	4/5/6	Weakly acidic
Succinic acid	4/5/6	Weakly acidic
Citric acid	4/5/6	Weakly acidic
Oxalic acid	4/5/6	Weakly acidic
Tartaric acid	4/5/6	Weakly acidic
Sulphuric(VI)acid	**1/2/3**	**Strongly acidic**

Explanations

Alkanoic acids are weak acids that partially/partly dissociate to release few H^+ ions in solution. The pH of their solution is thus 4/5/6 showing they form weakly acidic solutions when dissolved in water.

All alkanoic acid dissociate to releases the **"H"** at the functional group in -COO**H** to form the **alkanoate ion;** $-COO^-$

Mineral acids(Sulphuric(VI)acid, Nitric(V)acid and Hydrochloric acid) are strong acids that wholly/fully dissociate to release many H^+ ions in solution. The pH of their solution is thus 1/2/3 showing they form strongly acidic solutions when dissolved in water.i.e

Examples

1. $CH_3COOH(aq) \longleftarrow CH_3COO^-(aq) + H^+(aq)$

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(ethanoic acid) ⇢ (ethanoate ion) (few H^+ ion)

2. $CH_3CH_2COOH(aq) \rightleftharpoons CH_3CH_2COO^-(aq) + H^+(aq)$
   (propanoic acid) (propanoate ion) (few H^+ ion)
3. $CH_3CH_2CH_2COOH(aq) \rightleftharpoons CH_3CH_2CH_2COO^-(aq) + H^+(aq)$
   (Butanoic acid) (butanoate ion) (few H^+ ion)
4. $HOOH(aq) \rightleftharpoons HOO^-(aq) + H^+(aq)$
   (methanoic acid) (methanoate ion) (few H^+ ion)
5. $H_2SO_4(aq) \rightarrow SO_4^{2-}(aq) + 2H^+(aq)$
   (sulphuric(VI) acid) (sulphate(VI) ion) (**many** H^+ ion)
6. $HNO_3(aq) \rightarrow NO_3^-(aq) + H^+(aq)$
   (nitric(V) acid) (nitrate(V) ion) (**many** H^+ ion)

(c)Reaction with metals

Experiment

Place about 4cm3 of ethanoic acid in a test tube. Put about 1cm length of polished magnesium ribbon. Test any gas produced using a burning splint. Repeat with a solution of succinic acid, citric acid, oxalic acid, tartaric acid and dilute sulphuric (VI) acid.

Sample observations

Solution/acid	Observations	Inference
Ethanoic acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that burn with "pop" sound/explosion	$H_3O^+/H^+(aq)$ion
Succinic acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that burn with "pop" sound/explosion	$H_3O^+/H^+(aq)$ion
Citric acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that burn with "pop" sound/explosion	$H_3O^+/H^+(aq)$ion
Oxalic acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that burn with "pop" sound/explosion	$H_3O^+/H^+(aq)$ion
Tartaric acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that burn with "pop" sound/explosion	$H_3O^+/H^+(aq)$ion
Nitric(V)acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that burn with "pop" sound/explosion	$H_3O^+/H^+(aq)$ion

Explanation

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Metals higher in the reactivity series displace the hydrogen in all acids to evolve/produce hydrogen gas and form a salt. Alkanoic acids react with metals with metals to form alkanoates salt and produce/evolve hydrogen gas .Hydrogen extinguishes a burning splint with a pop sound/explosion. Only the "H"in the functional group -COO**H** is /are displaced and not in the alkyl hydrocarbon chain.

Alkanoic acid + Metal -> Alkanoate + Hydrogen gas. i.e.

Examples

1. For a monovalent metal with monobasic acid

$2R-COO\mathbf{H} + 2M \rightarrow 2R\text{-}COO\mathbf{M} + 2H_2(g)$

2.For a divalent metal with monobasic acid

$2R-COO\mathbf{H} + M \rightarrow (R\text{-}COO)_2\mathbf{M} + H_2(g)$

3.For a divalent metal with dibasic acid

$\mathbf{H}OOC\text{-}R\text{-}COO\mathbf{H} + M \rightarrow \mathbf{M}OOC\text{-}R\text{-}COO\mathbf{M} + H_2(g)$

4.For a monovalent metal with dibasic acid

$\mathbf{H}OOC\text{-}R\text{-}COO\mathbf{H} + 2M \rightarrow \mathbf{M}OOC\text{-}R\text{-}COO\mathbf{M} + H_2(g)$

5 For mineral acids

(i)Sulphuric(VI)acid is a dibasic acid

$\mathbf{H_2}SO_4(aq) + 2M \rightarrow M_2SO_4(aq) + H_2(g)$

$\mathbf{H_2}SO_4(aq) + M \rightarrow MSO_4(aq) + H_2(g)$

(ii)Nitric(V) and hydrochloric acid are monobasic acid

$\mathbf{H}NO_3(aq) + 2M \rightarrow 2MNO_3(aq) + H_2(g)$

$\mathbf{H}NO_3(aq) + M \rightarrow M(NO_3)_2(aq) + H_2(g)$

Examples

1.Sodium reacts with ethanoic acid to form sodium ethanoate and produce. hydrogen gas.

Caution: This reaction is explosive.

$CH_3COO\mathbf{H}(aq) + Na(s) \rightarrow CH_3COO\mathbf{Na}(aq) + H_2(g)$

(Ethanoic acid) (Sodium ethanoate)

2.Calcium reacts with ethanoic acid to form calcium ethanoate and produce. hydrogen gas.

$2CH_3COO\mathbf{H}(aq) + Ca(s) \rightarrow (CH_3COO)_2\mathbf{Ca}(aq) + H_2(g)$

(Ethanoic acid) (Calcium ethanoate)

3.Sodium reacts with ethan-1,2-dioic acid to form sodium ethan-1,2-dioate and produce. hydrogen gas.

$\mathbf{H}OOC\text{-}COO\mathbf{H} + 2Na \rightarrow \mathbf{Na}OOC\text{-}COO\mathbf{Na} + H_2(g)$

(ethan-1,2-dioic acid) (sodium ethan-1,2-dioate)

Commercial name of ethan-1,2-dioic acid is oxalic acid. The salt is sodium oxalate.

4.Magnesium reacts with ethan-1,2-dioic acid to form magnesium ethan-1,2-dioate and produce. hydrogen gas.

$\mathbf{H}OOC\text{-}R\text{-}COO\mathbf{H} + Mg \rightarrow (OOC\text{-}COO)\mathbf{Mg} + H_2(g)$

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(ethan-1,2-dioic acid) (magnesium ethan-1,2-dioate)

5.Magnesium reacts with

(i)Sulphuric(VI)acid to form Magnesium sulphate(VI)

$\mathbf{H_2}SO_4(aq) + Mg \rightarrow MgSO_4(aq) + H_2(g)$

(ii)Nitric(V) and hydrochloric acid are monobasic acid

$2\mathbf{H}NO_3(aq) + Mg \rightarrow M(NO_3)_2(aq) + H_2(g)$

(d)Reaction with hydrogen carbonates and carbonates

Experiment

Place about 3cm3 of ethanoic acid in a test tube. Add about 0.5g/ ½ spatula end full of sodium hydrogen carbonate/sodium carbonate. Test the gas produced using lime water. Repeat with a solution of succinic acid, citric acid, oxalic acid, tartaric acid and dilute sulphuric (VI) acid.

Sample observations

Solution/acid	Observations	Inference
Ethanoic acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that forms a white precipitate with lime water	$H_3O^+/H^+(aq)$ion
Succinic acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that forms a white precipitate with lime water	$H_3O^+/H^+(aq)$ion
Citric acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that forms a white precipitate with lime water	$H_3O^+/H^+(aq)$ion
Oxalic acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that forms a white precipitate with lime water	$H_3O^+/H^+(aq)$ion
Tartaric acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that forms a white precipitate with lime water	$H_3O^+/H^+(aq)$ion
Nitric(V)acid	(i)effervescence, fizzing, bubbles (ii)colourless gas produced that forms a white precipitate with lime water	$H_3O^+/H^+(aq)$ion

All acids react with hydrogen carbonate/carbonate to form salt ,water and evolve/produce bubbles of carbon(IV)oxide and water.

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Carbon(IV)oxide forms a white precipitate when bubbled in lime water/extinguishes a burning splint.
Alkanoic acids react with hydrogen carbonate/carbonate to form alkanoates ,water and evolve/produce bubbles of carbon(IV)oxide and water.
Alkanoic acid + hydrogen carbonate -> alkanoate + water + carbon(IV)oxide

Alkanoic acid + carbonate -> alkanoate + water + carbon(IV)oxide
Examples

1. Sodium hydrogen carbonate reacts with ethanoic acid to form sodium ethanoate ,water and carbon(IV)oxide gas.
   CH_3COOH (aq) + $NaHCO_3$ (s) -> CH_3COONa (aq) + $H_2O(l)$ + CO_2 (g)
   (Ethanoic acid) (Sodium ethanoate)
2. Sodium carbonate reacts with ethanoic acid to form sodium ethanoate ,water and carbon(IV)oxide gas.
   $2CH_3COOH$ (aq) + Na_2CO_3 (s) -> $2CH_3COONa$ (aq) + $H_2O(l)$ + CO_2 (g)
   (Ethanoic acid) (Sodium ethanoate)
3. Sodium carbonate reacts with ethan-1,2-dioic acid to form sodium ethanoate ,water and carbon(IV)oxide gas.

HOOC-COO**H**+ Na_2CO_3 (s) -> **Na**OOC - COO**Na** + $H_2O(l)$ + CO_2 (g)
(ethan-1,2-dioic acid) (sodium ethan-1,2-dioate)

4. Sodium hydrogen carbonate reacts with ethan-1,2-dioic acid to form sodium ethanoate ,water and carbon(IV)oxide gas.

HOOC-COO**H**+ $2NaHCO_3$ (s) -> **Na**OOC - COO**Na** + $H_2O(l)$ + $2CO_2$ (g)
(ethan-1,2-dioic acid) (sodium ethan-1,2-dioate)

(e)Esterification
Experiment
Place 4cm3 of ethanol acid in a boiling tube.
Add equal volume of ethanoic acid. To the mixture, add 2 drops of concentrated sulphuric(VI)acid **carefully**. Warm/heat gently on Bunsen flame.
Pour the mixture into a beaker containing 50cm3 of water. Smell the products.
Repeat with a solution of succinic acid, citric acid, oxalic acid, tartaric acid and dilute sulphuric (VI) acid.
Sample observations

Solution/acid	Observations
Ethanoic acid	Sweet fruity smell
Succinic acid	Sweet fruity smell
Citric acid	Sweet fruity smell
Oxalic acid	Sweet fruity smell
Tartaric acid	Sweet fruity smell
Dilute	**No sweet fruity smell**

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sulphuric(VI)acid	

Explanation

Alkanols react with alkanoic acid to form the sweet smelling homologous series of esters and water.The reaction is catalysed by concentrated sulphuric(VI)acid in the laboratory but naturally by sunlight /heat.Each ester has a characteristic smell derived from the many possible combinations of alkanols and alkanoic acids.

Alkanol + Alkanoic acids -> Ester + water

Esters derive their names from the alkanol first then alkanoic acids. The alkanol "becomes" an **alkyl** group and the alkanoic acid "becomes" **alkanoate** hence **alkylalkanoate**. e.g.

Ethanol + Ethanoic acid -> Ethylethanoate + Water
Ethanol + Propanoic acid -> Ethylpropanoate + Water
Ethanol + Methanoic acid -> Ethylmethanoate + Water
Ethanol + butanoic acid -> Ethylbutanoate + Water
Propanol + Ethanoic acid -> Propylethanoate + Water
Methanol + Ethanoic acid -> Methyethanoate + Water
Methanol + Decanoic acid -> Methyldecanoate + Water
Decanol + Methanoic acid -> Decylmethanoate + Water

During the formation of the ester, the "O" joining the alkanol and alkanoic acid comes from the alkanol.

R_1 -CO**OH** + R_2 –O**H** -> **R_1 -COO –R_2** + **H_2O**

Examples

1. Ethanol reacts with ethanoic acid to form the ester ethyl ethanoate and water.

Ethanol + Ethanoic acid --Conc. H_2SO_4 -->Ethylethanoate + Water

C_2H_5OH (l) + CH_3COOH(l) --Conc. H_2SO_4 --> $CH_3COO\ C_2H_5$(**aq**) +H_2O(l)

CH_3CH_2OH (l)+ CH_3COOH(l) --Conc. H_2SO_4 --> $CH_3COOCH_2CH_3$(**aq**) +H_2O(l)

2. Ethanol reacts with propanoic acid to form the ester ethylpropanoate and water.

Ethanol + Propanoic acid --Conc. H_2SO_4 -->Ethylethanoate + Water

C_2H_5OH (l)+ $CH_3\ CH_2COOH$(l) --Conc. H_2SO_4 -->$CH_3CH_2COO\ C_2H_5$(**aq**) +H_2O(l)

CH_3CH_2OH (l)+ $CH_3\ CH_2COOH$(l) --Conc. H_2SO_4 -->

$CH_3\ CH_2COOCH_2CH_3$(**aq**) +H_2O(l)

3. Methanol reacts with ethanoic acid to form the ester methyl ethanoate and water.

Methanol + Ethanoic acid --Conc. H_2SO_4 -->Methylethanoate + Water

CH_3OH (l) + CH_3COOH(l) --Conc. H_2SO_4 --> $CH_3COO\ CH_3$(**aq**) +H_2O(l)

4. Methanol reacts with propanoic acid to form the ester methyl propanoate and water.

Methanol + propanoic acid --Conc. H_2SO_4 -->Methylpropanoate + Water

CH_3OH (l)+ $CH_3\ CH_2COOH$(l) --Conc. H_2SO_4 --> $CH_3\ CH_2COO\ CH_3$(**aq**) +H_2O(l)

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5. Propanol reacts with propanoic acid to form the ester propylpropanoate and water.

Propanol + Propanoic acid --Conc. H_2SO_4 -->Ethylethanoate + Water

C_3H_7OH (l)+ CH_3 CH_2COOH(l) --Conc. H_2SO_4 -->CH_3CH_2COO C_3H_7(**aq**) +H_2O(l)

CH_3CH_2 CH_2OH (l)+ CH_3 CH_2COOH(l) --Conc. H_2SO_4 --> CH_3 CH_2COOCH_2 CH_2CH_3(**aq**) +H_2O(l)

C. DETERGENTS

Detergents are cleaning agents that improve the cleaning power /properties of water.A detergent therefore should be able to:

(i)dissolve substances which water cannot e.g grease ,oil, fat

(ii)be washed away after cleaning.

There are two types of detergents:

(a)Soapy detergents

(b)Soapless detergents

(a)SOAPY DETERGENTS

Soapy detergents usually called soap is long chain salt of organic alkanoic acids.Common soap is sodium octadecanoate .It is derived from reacting concentrated sodium hydroxide solution with octadecanoic acid(18 carbon alkanoic acid) i.e.

Sodium hydroxide + octadecanoic acid -> Sodium octadecanoate + water

NaOH(aq) + CH_3 $(CH_2)_{16}$ COOH(aq) -> CH_3 $(CH_2)_{16}$ COO^- $\mathbf{Na}^+$ (aq) +H_2 O(l)

Commonly ,soap can thus be represented ;

R- COO^- $\mathbf{Na}^+$ where;

R is a long chain alkyl group and -COO^- $\mathbf{Na}^+$ is the alkanoate ion.

In a school laboratory and at industrial and domestic level,soap is made by reacting concentrated sodium hydroxide solution with esters from (animal) **fat** and **oil**. The process of making soap is called **saponification.** During saponification ,the ester is **hydrolyzed** by the alkali to form sodium salt /soap and **glycerol/propan-1,2,3-triol** is produced.

Fat/oil(ester)+sodium/potassium hydroxide->sodium/potassium salt(soap)+ glycerol

Fats/Oils are esters with fatty acids and glycerol parts in their structure;

$C_{17}H_{35}COO\mathbf{CH_2}$

$C_{17}H_{35}COO\mathbf{CH}$

$C_{17}H_{35}COO\mathbf{CH_2}$

When boiled with concentrated sodium hydroxide solution NaOH;

(i)NaOH ionizes/dissociates into $\mathbf{Na}^+$ and $\mathbf{OH}^-$ ions

(ii)fat/oil split into **three** $C_{17}H_{35}COO^-$ and **one** CH_2 CH CH_2

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(iii) the three Na^+ combine with the three $C_{17}H_{35}COO^-$ to form the salt $C_{17}H_{35}COO^-\ Na^+$

(iv)the three OH^-ions combine with the $CH_2\ CH\ CH_2$ to form an alkanol with three functional groups CH_2 **OH** CH **OH** CH_2 **OH(propan-1,2,3-triol)**

$$\begin{array}{l} C_{17}H_{35}COOCH_2 \\ | \\ C_{17}H_{35}COOCH \\ | \\ C_{17}H_{35}COOCH_2 \end{array} + NaOH \rightarrow 3\ C_{17}H_{35}COO^-\ Na^+ + \begin{array}{l} CH_2OH \\ | \\ CHOH \\ | \\ CH_2OH \end{array}$$

Ester **Alkali** **Soap** **glycerol**

Generally:

$$\begin{array}{l} C_nH_{2n+1}COOCH_2 \\ | \\ C_nH_{2n+1}COOCH \\ | \\ C_nH_{2n+1}COOCH_2 \end{array} + NaOH \rightarrow 3\ C_nH_{2n+1}COO^-\ Na^+ + \begin{array}{l} CH_2OH \\ | \\ CHOH \\ | \\ CH_2OH \end{array}$$

Ester **Alkali** **Soap** **glycerol**

$$\begin{array}{l} R\text{ - }COOCH_2 \\ | \\ R\text{ - }COOCH \\ | \\ R\text{- }COOCH_2 \end{array} + NaOH \rightarrow 3R\text{-}COO^-\ Na^+ + \begin{array}{l} CH_2OH \\ | \\ CHOH \\ | \\ CH_2OH \end{array}$$

Ester **Alkali** **Soap** **glycerol**

During this process a little sodium chloride is added to **precipitate** the soap by reducing its solubility. This is called **salting out**.

The soap is then added colouring agents ,perfumes and herbs of choice.

School laboratory preparation of soap

Place about 40 g of fatty (animal fat)beef/meat in 100cm3 beaker .Add about 15cm3 of 4.0M sodium hydroxide solution. Boil the mixture for about 15minutes.Stir the mixture .Add about 5.0cm3 of distilled water as you boil to make up for evaporation. Boil for about another 15minutes.Add about four spatula end full of pure sodium chloride crystals. Continue stirring for another five minutes. Allow to cool. Filter of /decant and wash off the residue with distilled water .Transfer the clean residue into a dry beaker. Preserve.

The action of soap

Soapy detergents:

(i)act by reducing the surface tension of water by forming a thin layer on top of the water.

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(ii)is made of a **non-polar** alkyl /hydrocarbon tail and a **polar** $-COO^-Na^+$ head. The non-polar alkyl /hydrocarbon tail is **hydrophobic** (water hating) and thus does not dissolve in water .It dissolves in non-polar solvent like grease, oil and fat. The polar $-COO^-Na^+$ head is **hydrophilic** (water loving)and thus dissolve in water. When washing with soapy detergent, the non-polar tail of the soapy detergent surround/dissolve in the dirt on the garment /grease/oil while the polar head dissolve in water.

Through **mechanical agitation**/stirring/sqeezing/rubbing/beating/kneading, some grease is dislodged/lifted of the surface of the garment. It is immediately surrounded by more soap molecules It float and spread in the water as tiny droplets that scatter light in form of emulsion making the water cloudy and shinny. It is removed from the garment by rinsing with fresh water. The repulsion of the soap head prevent /ensure the droplets do not mix. Once removed, the dirt molecules cannot be redeposited back because it is surrounded by soap molecules.

Advantages and disadvantages of using soapy detergents

Soapy detergents are biodegradable. They are acted upon by bacteria and rot. They thus do not cause environmental pollution.

Soapy detergents have the disadvantage in that:

(i)they are made from fat and oils which are better eaten as food than make soap.

(ii)forms an insoluble precipitate with hard water called **scum.** Scum is insoluble calcium octadecanoate and Magnesium octadecanoate formed when soap reacts with Ca^{2+} and Mg^{2+} present in hard water.

Chemical equation

$2C_{17}H_{35}COO^-\ \mathbf{Na^+}(aq) + Ca^{2+}(aq) \rightarrow \mathbf{(}C_{17}H_{35}COO^-\mathbf{)}Ca^{2+}(s) + \mathbf{2Na^+}(aq)$

(insoluble Calcium octadecanote/scum)

$2C_{17}H_{35}COO^-\ \mathbf{Na^+}(aq) + Mg^{2+}(aq) \rightarrow \mathbf{(}C_{17}H_{35}COO^-\mathbf{)}Mg^{2+}(s) + \mathbf{2Na^+}(aq)$

(insoluble Magnesium octadecanote/scum)

This causes wastage of soap.

Potassium soaps are better than Sodium soap. Potassium is more expensive than sodium and thus its soap is also more expensive.

(b)SOAPLESS DETERGENTS

Soapless detergent usually called detergent is a long chain salt fromed from by-products of fractional distillation of crude oil.Commonly used soaps include:

(i)washing agents

(ii)toothpaste

(iii)emulsifiers/wetting agents/shampoo

Soapless detergents are derived from reacting:

(i)concentrated sulphuric(VI)acid with a long chain alkanol e.g. Octadecanol(18 carbon alkanol) to form alkyl hydrogen sulphate(VI)

Alkanol + Conc sulphuric(VI)acid -> alkyl hydrogen sulphate(VI) + Water

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$R–OH + H_2SO_4 \rightarrow R–O-SO_3H + H_2O$

(ii)the alkyl hydrogen sulphate(VI) is then neutralized with sodium/potassium hydroxide to form sodium/potassium alkyl hydrogen sulphate(VI)

Sodium/potassium alkyl hydrogen sulphate(VI) is the soapless detergent.

alkyl hydrogen sulphate(VI) + Potassium/sodium hydroxide -> Sodium/potassium alkyl hydrogen sulphate(VI) + Water

$R–O-SO_3H + NaOH \rightarrow R–O-SO_3^- Na^+ + H_2O$

Example

Step I : Reaction of Octadecanol with Conc.H_2SO_4

$C_{17}H_{35}CH_2OH\,(aq) + H_2SO_4 \rightarrow C_{17}H_{35}CH_2\text{-}\mathbf{O\text{-}\ SO_3^-\ H^+}\,(aq) + H_2O\,(l)$

octadecanol + sulphuric(VI)acid -> Octadecyl hydrogen sulphate(VI) + water

Step II: Neutralization by an alkali

$C_{17}H_{35}CH_2\text{-}\mathbf{O\text{-}\ SO_3^-\ H^+}\,(aq) + NaOH \rightarrow C_{17}H_{35}CH_2\text{-}\mathbf{O\text{-}\ SO_3^-\ Na^+}\,(aq) + H_2O\,(l)$

Octadecyl hydrogen sulphate(VI) + sodium/potassium hydroxide -> sodium/potassium octadecyl hydrogen sulphate(VI)+Water

School laboratory preparation of soapless detergent

Place about 20g of olive oil in a 100cm3 beaker. Put it in a trough containing ice cold water.

Add dropwise carefully 18M concentrated sulphuric(VI)acid stirring continuously into the olive oil until the oil turns brown.Add 30cm3 of 6M sodium hydroxide solution.Stir.This is a soapless detergent.

The action of soapless detergents

The action of soapless detergents is similar to that of soapy detergents.The soapless detergents contain the hydrophilic head and a long hydrophobic tail. i.e.

vvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvv-COO^-Na^+

vvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvvv-$O\text{-}SO_3^- Na^+$

(long hydrophobic /non-polar alkyl tail) (hydrophilic/polar/ionic head)

The tail dissolves in fat/grease/oil while the ionic/polar/ionic head dissolves in water.

The tail stick to the dirt which is removed by the attraction of water molecules and the polar/ionic/hydrophilic head by mechanical agitation /squeezing/kneading/ beating/rubbing/scrubbing/scatching.

The suspended dirt is then surrounded by detergent molecules and repulsion of the anion head preventing the dirt from sticking on the material garment.

The tiny droplets of dirt emulsion makes the water cloudy. On rinsing the cloudy emulsion is washed away.

Advantages and disadvantages of using soapless detergents

Soapless detergents are non-biodegradable unlike soapy detergents.

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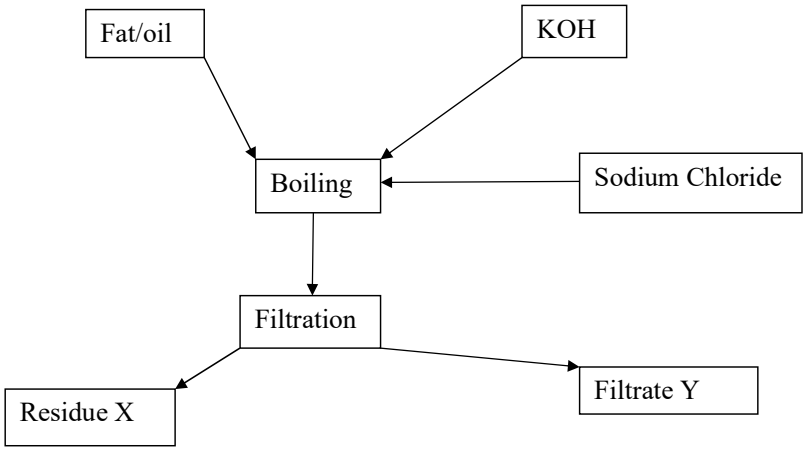
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They persist in water during sewage treatment by causing foaming in rivers ,lakes and streams leading to marine /aquatic death.
Soapless detergents have the advantage in that they:

(i)do not form scum with hard water.
(ii)are cheap to manufacture/buying
(iii)are made from petroleum products but soapis made from fats/oil for human consumption.

Sample revision questions

1. Study the scheme below



(a)Identify the process

Saponification

(b)Fats and oils are esters. Write the formula of the a common structure of ester

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$C_{17}H_{35}COOCH_2$
|
$C_{17}H_{35}COOCH$
|
$C_{17}H_{35}COOCH_2$

(c)Write a balanced equation for the reaction taking place during boiling

$$\begin{array}{l} C_{17}H_{35}COOCH_2 \\ | \\ C_{17}H_{35}COOCH \\ | \\ C_{17}H_{35}COOCH_2 \end{array} + 3NaOH \rightarrow 3\ C_{17}H_{35}COO^-Na^+ + \begin{array}{l} CH_2OH \\ | \\ CHOH \\ | \\ CH_2OH \end{array}$$

Ester Alkali Soap glycerol

(d)Give the IUPAC name of:

(i)Residue X

Potassium octadecanoate

(ii)Filtrate Y

Propan-1,2,3-triol

(e)Give one use of fitrate Y

Making paint

(f)What is the function of sodium chloride

To reduce the solubility of the soap hence helping in precipitating it out

(g)Explain how residue X helps in washing.

Has a non-polar hydrophobic tail that dissolves in dirt/grease /oil/fat

Has a polar /ionic hydrophilic head that dissolves in water.

From mechanical agitation,the dirt is plucked out of the garment and surrounded by the tail end preventing it from being deposited back on the garment.

(h)State one:

(i)advantage of continued use of residue X on the environment

Is biodegradable and thus do not pollute the environment

(ii)disadvantage of using residue X

Uses fat/oil during preparation/manufacture which are better used for human consumption.

(i)Residue X was added dropwise to some water.The number of drops used before lather forms is as in the table below.

	Water sample		
	A	**B**	**C**
Drops of residue X	**15**	**2**	**15**

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Drops of residue X in boiled water	**2**	**2**	**15**

(i)State and explain which sample of water is:

I. Soft

Sample B .Very little soap is used and no effect on amount of soap even on boiling/heating.

II. Permanent hard

Sample C . A lot of soap is used and no effect on amount of soap even on boiling/heating. Boiling does not remove permanent hardness of water.

III. Temporary hard

Sample A . A lot of soap is used before boiling. Very little soap is used on boiling/heating. Boiling remove temporary hardness of water.

(ii)Write the equation for the reaction at water sample C.

Chemical equation

$2C_{17}H_{35}COO^{-}\,\mathbf{K}^{+}(aq) + CaSO_4(aq) \rightarrow (C_{17}H_{35}COO^{-})Ca^{2+}(s) + \mathbf{K}_2SO_4(aq)$

(insoluble Calcium octadecanote/scum)

Ionic equation

$2C_{17}H_{35}COO^{-}\,\mathbf{K}^{+}(aq) + Ca^{2+}(aq) \rightarrow (C_{17}H_{35}COO^{-})Ca^{2+}(s) + \mathbf{2K}^{+}(aq)$

(insoluble Calcium octadecanote/scum)

Chemical equation

$2C_{17}H_{35}COO^{-}\,\mathbf{K}^{+}(aq) + MgSO_4(aq) \rightarrow (C_{17}H_{35}COO^{-})Mg^{2+}(s) + \mathbf{K}_2SO_4(aq)$

(insoluble Calcium octadecanote/scum)

Ionic equation

$2C_{17}H_{35}COO^{-}\,\mathbf{K}^{+}(aq) + Mg^{2+}(aq) \rightarrow (C_{17}H_{35}COO^{-})Mg^{2+}(s) + \mathbf{2K}^{+}(aq)$

(insoluble Magnesium octadecanote/scum)

(iii)Write the equation for the reaction at water sample A before boiling.

Chemical equation

$2C_{17}H_{35}COO^{-}\,\mathbf{K}^{+}(aq) + Ca(HCO_3)(aq) \rightarrow (C_{17}H_{35}COO^{-})Ca^{2+}(s) + 2KHCO_3(aq)$

(insoluble Calcium octadecanote/scum)

Ionic equation

$2C_{17}H_{35}COO^{-}\,\mathbf{K}^{+}(aq) + Ca^{2+}(aq) \rightarrow (C_{17}H_{35}COO^{-})Ca^{2+}(s) + 2K^{+}(aq)$

(insoluble Calcium octadecanote/scum)

Chemical equation

$2C_{17}H_{35}COO^{-}\,\mathbf{K}^{+}(aq) + Mg(HCO_3)(aq) \rightarrow (C_{17}H_{35}COO^{-})Mg^{2+}(s) + 2KHCO_3(aq)$

(insoluble Calcium octadecanote/scum)

Ionic equation

$2C_{17}H_{35}COO^{-}\,\mathbf{K}^{+}(aq) + Mg^{2+}(aq) \rightarrow (C_{17}H_{35}COO^{-})Mg^{2+}(s) + 2K^{+}(aq)$

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(insoluble Magnesium octadecanote/scum)

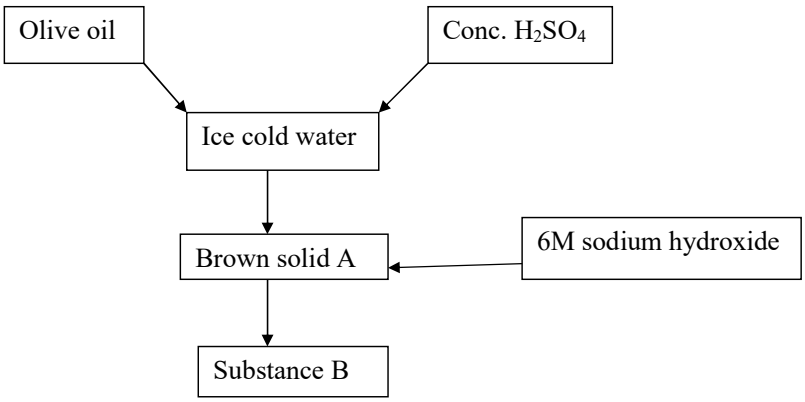
(iv)Explain how water becomes hard

Natural or rain water flowing /passing through rocks containing calcium (chalk, gypsum, limestone)and magnesium compounds (dolomite)dissolve them to form soluble Ca^{2+} and Mg^{2+} ions that causes water hardness.

(v)State two useful benefits of hard water

-Used in bone and teeth formation
-Coral polyps use hard water to form coral reefs
-Snails use hard water to make their shells

2.Study the scheme below and use it to answer the questions that follow.



(a)Identify :

(i)brown solid A

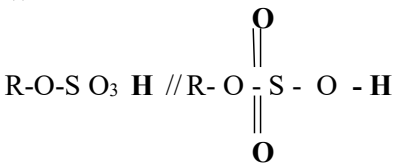
Alkyl hydrogen sulphate(VI)

(ii)substance B

Sodium alkyl hydrogen sulphate(VI)

(b)Write a general formula of:

(i)Substance A.



(ii)Substance B



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(c)State one

(i) advantage of continued use of substance B

-Does not form scum with hard water

-Is cheap to make

-Does not use food for human as a raw material.

(ii)disadvantage of continued use of substance B.

Is non-biodegradable therefore do not pollute the environment

(d)Explain the action of B during washing.

Has a non-polar hydrocarbon long tail that dissolves in dirt/grease/oil/fat.

Has a polar/ionic hydrophilic head that dissolves in water

Through mechanical agitation the dirt is plucked /removed from the garment and surrounded by the tail end preventing it from being deposited back on the garment.

(e) Ethene was substituted for olive oil in the above process. Write the equation and name of the new products A and B.

Product A

Ethene + Sulphuric(VI)acid -> Ethyl hydrogen sulphate(VI)

$H_2C=CH_2 + H_2SO_4 \rightarrow H_3C - CH_2 - O\text{-}SO_3H$

Product B

Ethyl hydrogen sulphate(VI) + sodium hydroxide -> sodium Ethyl hydrogen sulphate(VI) + Water

$H_3C - CH_2 - O\text{-}SO_3H + \mathbf{Na}OH \rightarrow H_3C - CH_2 - O\text{-}SO_3^-\mathbf{Na}^+ + H_2O$

(f)Ethanol can also undergo similar reactions forming new products A and B.Show this using a chemical equation.

Product A

Ethanol + Sulphuric(VI)acid ->Ethyl hydrogen sulphate(VI) + water

$H_3C\text{-}CH_2OH + H_2SO_4 \rightarrow H_3C - CH_2 - O\text{-}SO_3H + H_2O$

Product B

Ethyl hydrogen sulphate(VI) + sodium hydroxide -> sodium Ethyl hydrogen sulphate(VI) + Water

$H_3C - CH_2 - O\text{-}SO_3H + \mathbf{Na}OH \rightarrow H_3C - CH_2 - O\text{-}SO_3^-\mathbf{Na}^+ + H_2O$

3.Below is part of a detergent

$\mathbf{H_3C - (CH_2)_{16} - O - SO_3^- K^+}$

(a)Write the formular of the polar and non-polar end

Polar end

$H_3C - (CH_2)_{16} -$

Non-polar end

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– O - SO_3^- K^+

(b)Is the molecule a soapy or saopless detergent?

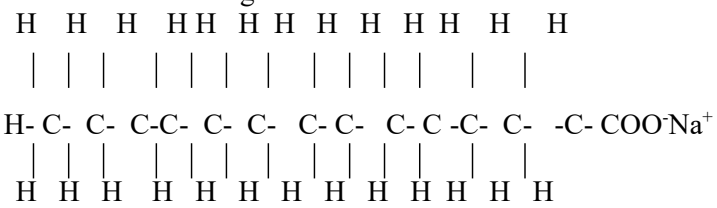
Soapless detergent

(c)State one advantage of using the above detergent

-does not form scum with hard water

-is cheap to manufacture

4.The structure of a detergent is



a) Write the molecular formula of the detergent. (1mk)

$CH_3(CH_2)_{12}COO^-Na^+$

b) What type of detergent is represented by the formula? (1mk)

Soapy detergent

c) When this type of detergent is used to wash linen in hard water, spots (marks) are left on the linen. Write the formula of the substance responsible for the spots

$(CH_3(CH_2)_{12}COO^-)_2Ca^{2+}$ / $CH_3(CH_2)_{12}COO^-)_2Mg^{2+}$

D. POLYMERS AND FIBRES

Polymers and fibres are giant molecules of organic compounds. Polymers and fibres are formed when **small** molecules called monomers join together to form **large** molecules called polymers at high temperatures and pressures. This process is called polymerization.

Polymers and fibres are either:

(a)**Natural** polymers and fibres

(b)**Synthetic** polymers and fibres

Natural polymers and fibres are found in living things(plants and animals) Natural polymers/fibres include:

-proteins/polypeptides making amino acids in animals

-cellulose that make cotton,wool,paper and silk

-Starch that come from glucose

-Fats and oils

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-Rubber from latex in rubber trees.

Synthetic polymers and fibres are man-made. They include:

-polyethene
-polychloroethene
-polyphenylethene(polystyrene)
-Terylene(Dacron)
-Nylon-6,6
-Perspex(artificial glass)

Synthetic polymers and fibres have the following characteristic advantages over natural polymers

1. They are light and portable
2. They are easy to manufacture.
3. They can easily be molded into shape of choice.
4. They are resistant to corrosion, water, air , acids, bases and salts.
5. They are comparatively cheap, affordable, colourful and aesthetic

Synthetic polymers and fibres however have the following disadvantages over natural polymers

1. They are non-biodegradable and hence cause environmental pollution during disposal
2. They give out highly poisonous gases when burnt like chlorine/carbon(II)oxide
3. Some on burning produce Carbon(IV)oxide. Carbon(IV)oxide is a green house gas that cause global warming.
4. Compared to some metals, they are poor conductors of heat,electricity and have lower tensile strength.

To reduce environmental pollution from synthetic polymers and fibres, the followitn methods of disposal should be used:

1.Recycling: Once produced all synthetic polymers and fibres should be recycled to a new product. This prevents accumulation of the synthetic polymers and fibres in the environment.

2.Production of biodegradable synthetic polymers and fibres that **rot** away.

There are two types of polymerization:

(a)addition polymerization
(b)condensation polymerization

(a)addition polymerization

Addition polymerization is the process where a small unsaturated monomer (alkene) molecule join together to form a large saturated molecule. Only alkenes undergo addition polymerization.

Addition polymers are named from the alkene/monomer making the polymer and adding the prefix "**poly**" before the name of monomer to form a **polyalkene**

During addition polymerization

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(i)the double bond in alkenes break
(ii)free radicals are formed
(iii)the free radicals collide with each other and join to form a larger molecule. The more collisions the larger the molecule.

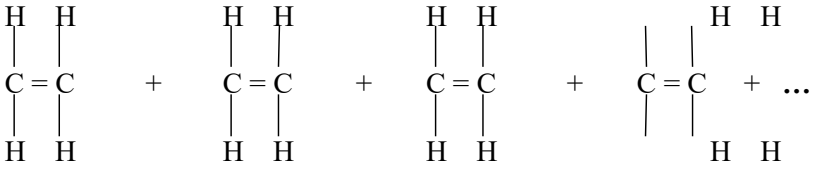
Examples of addition polymerization

1.Formation of Polyethene

Polyethene is an addition polymer formed when ethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

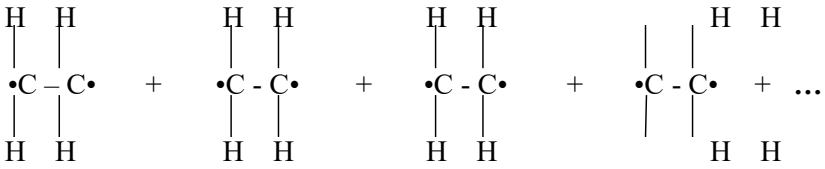
During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting paticles)



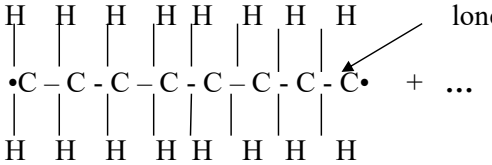
Ethene + Ethene + Ethene + Ethene + ...

(ii)the double bond joining the ethane molecule break to free readicals



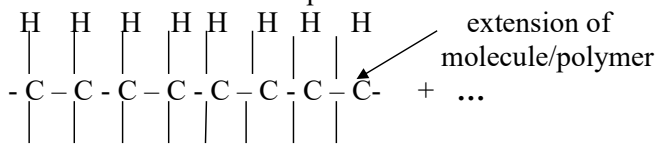
Ethene radical + Ethene radical + Ethene radical + Ethene radical + ...

(iii)the free radicals collide with each other and join to form a larger molecule



Lone pair of electrons can be used to join more monomers to form longer polyethene.

Polyethene molecule can be represented as:

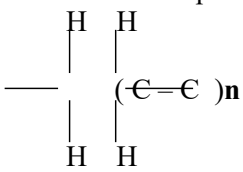


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H H H H H H H H

Since the molecule is a repetition of one monomer, then the polymer is:



Where n is the number of monomers in the polymer. The number of monomers in the polymer can be determined from the molar mass of the polymer and monomer from the relationship:

Number of monomers/repeating units in monomer = Molar mass polymer
Molar mass monomer

Examples

Polythene has a molar mass of 4760.Calculate the number of ethene molecules in the polymer(C=12.0, H=1.0)

Number of monomers/repeating units in polyomer = Molar mass polymer
Molar mass monomer

=> Molar mass ethene (C_2H_4)= 28 Molar mass polyethene = 4760

Substituting 4760 = 170 ethene molecules
28

The **commercial** name of polyethene is **polythene**. It is an elastic, tough, transparent and durable plastic. Polythene is used:

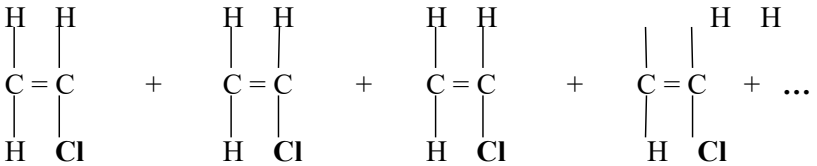
(i)in making plastic bag
(ii)bowls and plastic bags
(iii)packaging materials

2.Formation of Polychlorethene

Polychloroethene is an addition polymer formed when chloroethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)

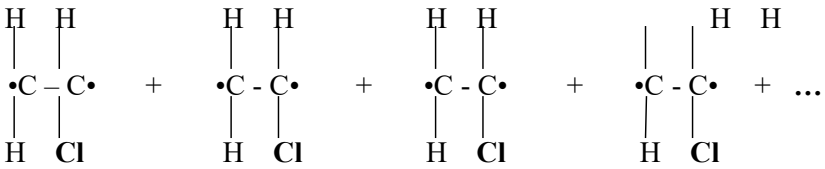


chloroethene + chloroethene + chloroethene + chloroethene + …

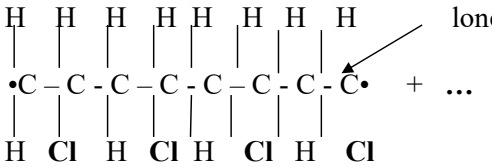
(ii)the double bond joining the chloroethene molecule break to free radicals

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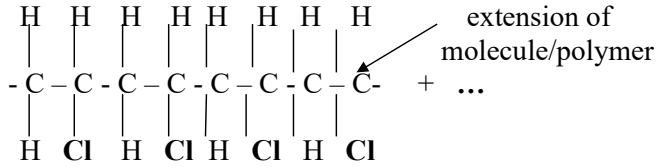


(iii)the free radicals collide with each other and join to form a larger molecule

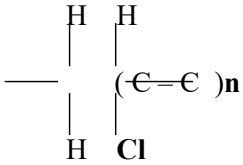


Lone pair of electrons can be used to join more monomers to form longer polychloroethene.

Polychloroethene molecule can be represented as:



Since the molecule is a repetition of one monomer, then the polymer is:



Examples

Polychlorothene has a molar mass of 4760.Calculate the number of chlorethene molecules in the polymer(C=12.0, H=1.0,Cl=35.5)

Number of monomers/repeating units in monomer = Molar mass polymer

Molar mass monomer

=> Molar mass ethene (C_2H_3Cl)= 62.5 Molar mass polyethene = 4760

Substituting 4760 = 77.16 => 77 polychloroethene molecules**(whole number)**

62.5

The **commercial** name of polychloroethene is **polyvinylchloride(PVC)**. It is a tough, non-transparent and durable plastic. PVC is used:

(i)in making plastic rope

(ii)water pipes

(iii)crates and boxes

3.Formation of Polyphenylethene

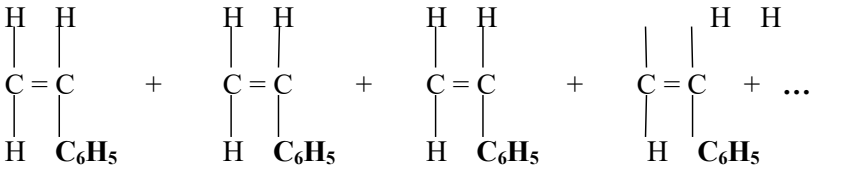
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Polyphenylethene is an addition polymer formed when phenylethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

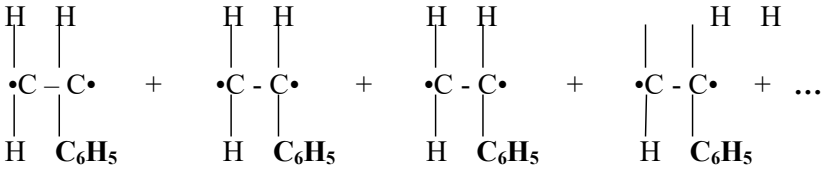
During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)

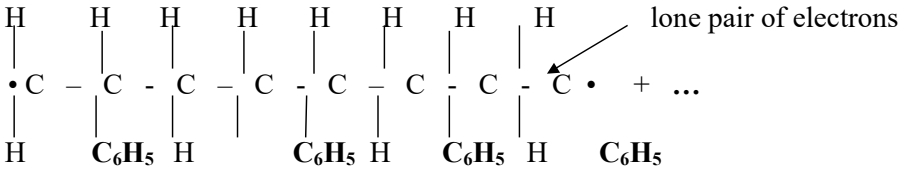


phenylethene + phenylethene + phenylethene + phenylethene + …

(ii)the double bond joining the phenylethene molecule break to free radicals

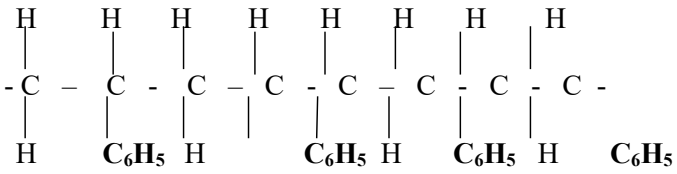


(iii)the free radicals collide with each other and join to form a larger molecule

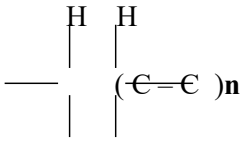


Lone pair of electrons can be used to join more monomers to form longer polyphenylethene.

Polyphenylethene molecule can be represented as:



Since the molecule is a repetition of one monomer, then the polymer is:



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H C_6H_5

Examples

Polyphenylthene has a molar mass of 4760.Calculate the number of phenylethene molecules in the polymer(C=12.0, H=1.0,)

Number of monomers/repeating units in monomer = Molar mass polymer
Molar mass monomer

=> Molar mass ethene (C_8H_8)= 104 Molar mass polyethene = 4760

Substituting 4760 = 45.7692 =>45 polyphenylethene molecules**(whole number)**
104

The **commercial** name of polyphenylethene is **polystyrene**. It is a very light durable plastic. Polystyrene is used:

(i)in making packaging material for carrying delicate items like computers, radion,calculators.

(ii)ceiling tiles

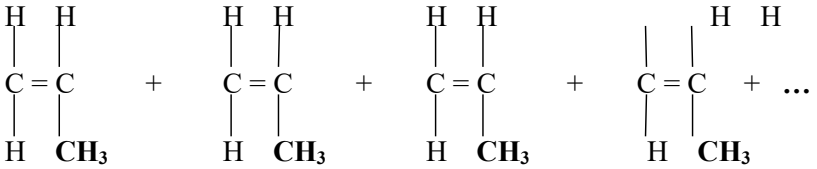
(iii)clothe linings

4.Formation of Polypropene

Polypropene is an addition polymer formed when propene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

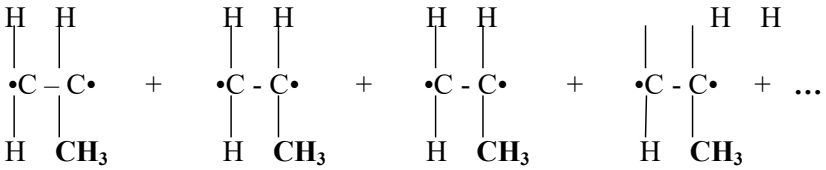
During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)



propene + propene + propene+ propene + …

(ii)the double bond joining the phenylethene molecule break to free radicals

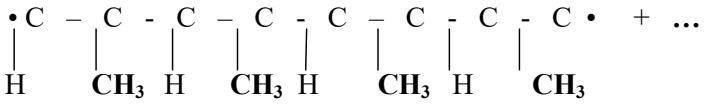


(iii)the free radicals collide with each other and join to form a larger molecule

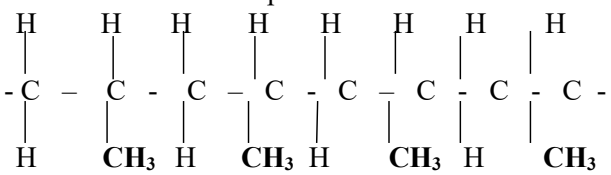


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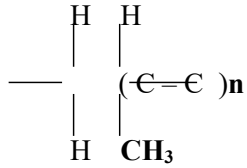
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Lone pair of electrons can be used to join more monomers to form longer propene. propene molecule can be represented as:



Since the molecule is a repetition of one monomer, then the polymer is:



Examples

Polypropene has a molar mass of 4760.Calculate the number of propene molecules in the polymer(C=12.0, H=1.0,)

Number of monomers/repeating units in monomer = Molar mass polymer
Molar mass monomer

=> Molar mass propene (C_3H_8)= 44 Molar mass polyethene = 4760

Substituting 4760 = 108.1818 =>108 propene molecules**(whole number)**
44

The **commercial** name of polyphenylethene is **polystyrene**. It is a very light durable plastic. Polystyrene is used:

(i)in making packaging material for carrying delicate items like computers, radion,calculators.

(ii)ceiling tiles

(iii)clothe linings

5.Formation of Polytetrafluorothene

Polytetrafluorothene is an addition polymer formed when tetrafluoroethene molecule/monomer join together to form a large molecule/polymer at high temperatures and pressure.

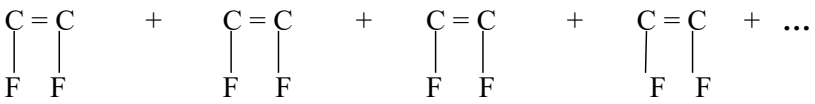
During polymerization:

(i)many molecules are brought nearer to each other by the high pressure(which reduces the volume occupied by reacting particles)



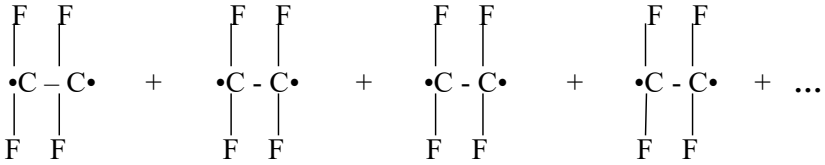
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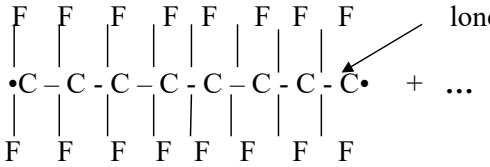


tetrafluoroethene + tetrafluoroethene+ tetrafluoroethene+ tetrafluoroethene + …

(ii)the double bond joining the tetrafluoroethene molecule break to free radicals

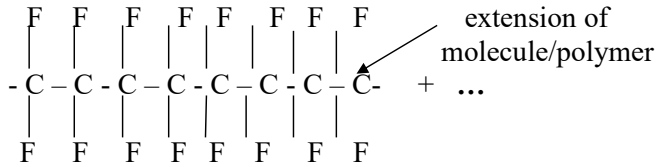


(iii)the free radicals collide with each other and join to form a larger molecule

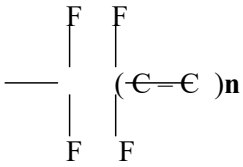


Lone pair of electrons can be used to join more monomers to form longer polytetrafluoroethene.

polytetrafluoroethene molecule can be represented as:



Since the molecule is a repetition of one monomer, then the polymer is:



Examples

Polytetrafluorothene has a molar mass of 4760.Calculate the number of tetrafluoroethene molecules in the polymer(C=12.0, ,F=19)

Number of monomers/repeating units in monomer = Molar mass polymer

Molar mass monomer

=> Molar mass ethene (C_2F_4)= 62.5 Molar mass polyethene = 4760

Substituting 4760 = 77.16 => 77 polychloroethene molecules**(whole number)**

62.5

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The **commercial** name of polytetrafluorethene(**P.T.F.E**) is **Teflon(P.T.F.E)**. It is a tough, non-transparent and durable plastic. PVC is used:

(i)in making plastic rope
(ii)water pipes
(iii)crates and boxes

5.Formation of rubber from Latex

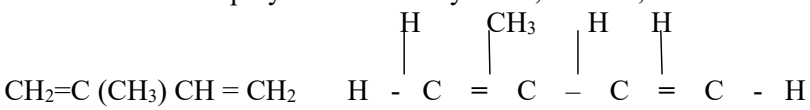
Natural rubber is obtained from rubber trees.

During harvesting an incision is made on the rubber tree to produce a milky white substance called **latex.**

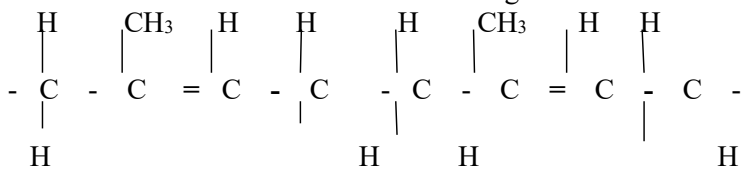
Latex is a mixture of rubber and lots of water.

The latex is then added an acid to coagulate the rubber.

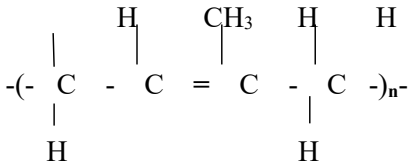
Natural rubber is a polymer of 2-methylbut-1,3-diene ;



During natural polymerization to rubber, one double C=C bond break to self add to another molecule. The double bond remaining move to carbon "2" thus;

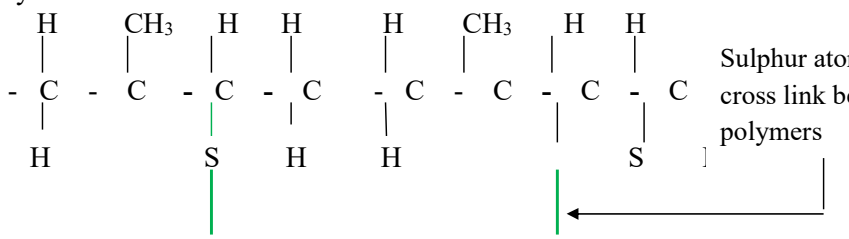


Generally the structure of rubber is thus;



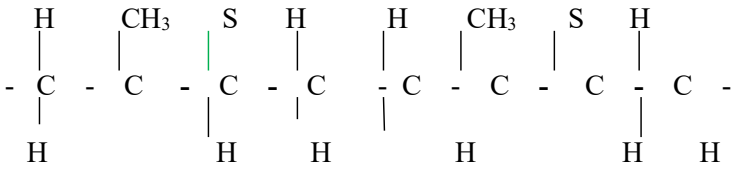
Pure rubber is soft and sticky. It is used to make erasers, car tyres. Most of it is vulcanized. Vulcanization is the process of heating rubber with sulphur to make it harder/tougher.

During vulcanization the sulphur atoms form a cross link between chains of rubber molecules/polymers. This decreases the number of C=C double bonds in the polymer.



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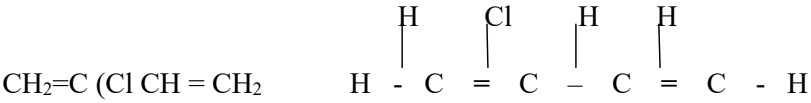


Vulcanized rubber is used to make **tyres**, **shoes** and **valves**.

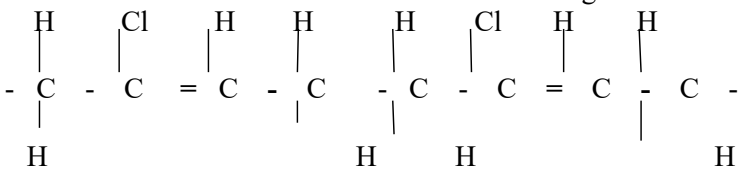
6.Formation of synthetic rubber

Synthetic rubber is able to resist action of oil,abrasion and organic solvents which rubber cannot.

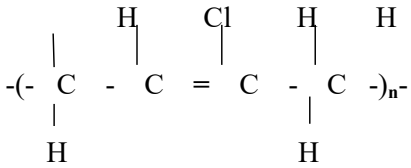
Common synthetic rubber is a polymer of 2-chlorobut-1,3-diene ;



During polymerization to synthetic rubber, one double C=C bond is broken to self add to another molecule. The double bond remaining move to carbon "2" thus;



Generally the structure of rubber is thus;



Rubber is thus strengthened through vulcanization and manufacture of synthetic rubber.

(b)Condensation polymerization

Condensation polymerization is the process where two or more small monomers join together to form a larger molecule by elimination/removal of a simple molecule. (usually water).

Condensation polymers acquire a different name from the monomers because the two monomers are two different compounds

During condensation polymerization:

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(i)the two monomers are brought together by high pressure to reduce distance between them.

(ii)monomers realign themselves at the functional group.

(iii)from each functional group an element is removed so as to form simple molecule (of usually H_2O/HCl)

(iv)the two monomers join without the simple molecule of H_2O/HCl

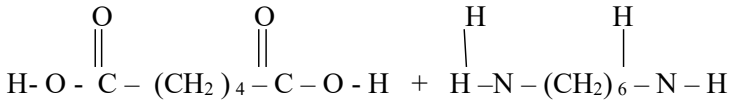
Examples of condensation polymerization

1.Formation of Nylon-6,6

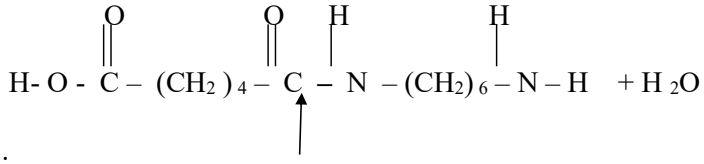
Method 1: Nylon-6,6 can be made from the condensation polymerization of hexan-1,6-dioic acid with hexan-1,6-diamine.Amines are a group of homologous series with a general formula R-NH_2 and thus -NH_2 as the functional group.

During the formation of Nylon-6,6:

(i)the two monomers are brought together by high pressure to reduce distance between them and realign themselves at the functional groups.



(iii)from each functional group an element is removed so as to form a molecule of H_2O and the two monomers join at the linkage .



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Polymer bond linkage

Nylon-6,6 derive its name from the two monomers each with **six** carbon chain

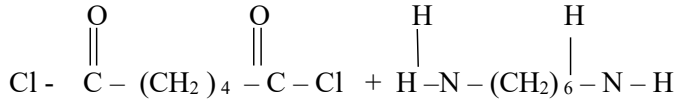
Method 2: Nylon-6,6 can be made from the condensation polymerization of hexan-1,6-dioyl dichloride with hexan-1,6-diamine.

Hexan-1,6-dioyl dichloride belong to a group of homologous series with a general formula R-OCl and thus -OCl as the functional group.

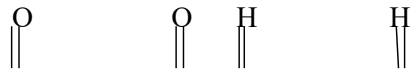
The R-OCl is formed when the "OH" in R-O**OH**/alkanoic acid is replaced by Cl/chlorine/Halogen

During the formation of Nylon-6,6:

(i)the two monomers are brought together by high pressure to reduce distance between them and realign themselves at the functional groups.

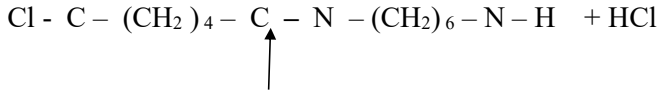


(iii)from each functional group an element is removed so as to form a molecule of HCl and the two monomers join at the linkage .



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Polymer bond linkage

The two monomers each has **six** carbon chain hence the name "nylon-6,6"
The commercial name of Nylon-6,6 is **Nylon** It is a a tough, elastic and durable plastic. It is used to make **clothes**, **plastic ropes** and **carpets.**

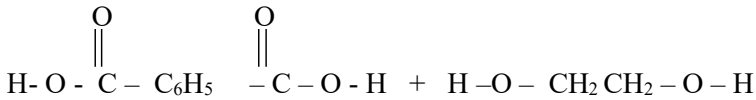
2.Formation of Terylene

Method 1: Terylene can be made from the condensation polymerization of ethan-1,2-diol with benzene-1,4-dicarboxylic acid.

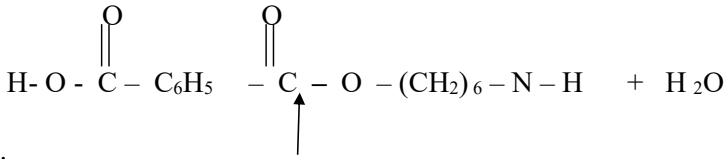
Benzene-1,4-dicarboxylic acid a group of homologous series with a general formula R-COOH where R is a ring of six carbon atom called Benzene ring .The functional group is -COOH.

During the formation of Terylene:

(i)the two monomers are brought together by high pressure to reduce distance between them and realign themselves at the functional groups.



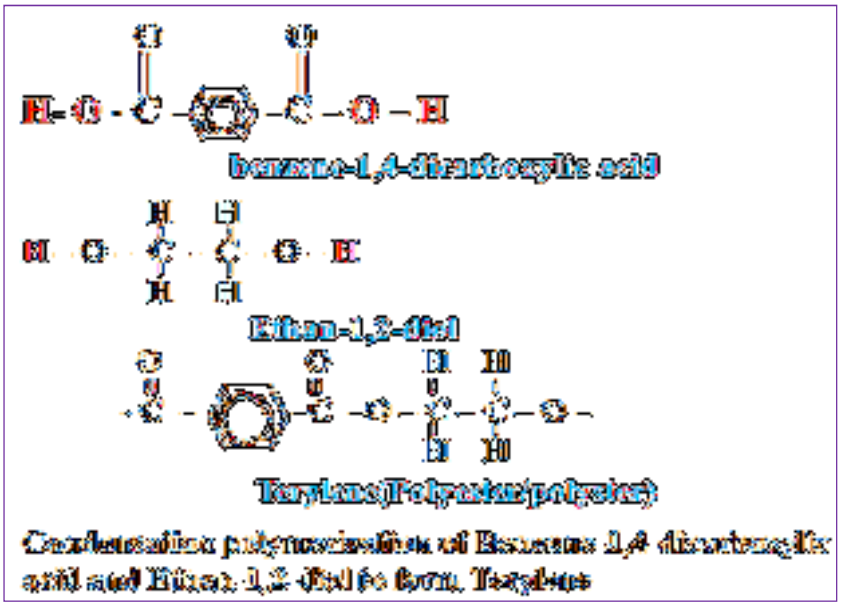
(iii)from each functional group an element is removed so as to form a molecule of H_2O and the two monomers join at the linkage .



Polymer bond linkage of terylene

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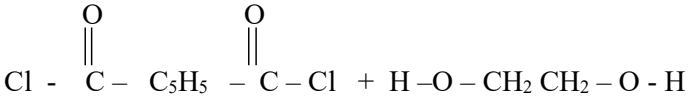
Method 2: Terylene can be made from the condensation polymerization of benzene-1,4-dioyl dichloride with ethan-1,2-diol.

Benzene-1,4-dioyl dichloride belong to a group of homologous series with a general formula R-OCl and thus -OCl as the functional group and R as a benzene ring.

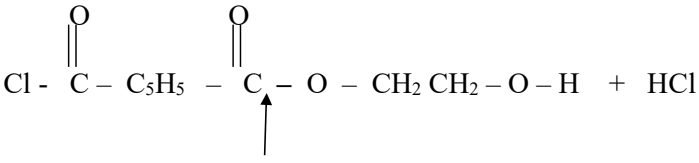
The R-OCl is formed when the "OH" in R-O**OH** is replaced by Cl/chlorine/Halogen

During the formation of Terylene

(i)the two monomers are brought together by high pressure to reduce distance between them and realign themselves at the functional groups.



(iii)from each functional group an element is removed so as to form a molecule of HCl and the two monomers join at the linkage .

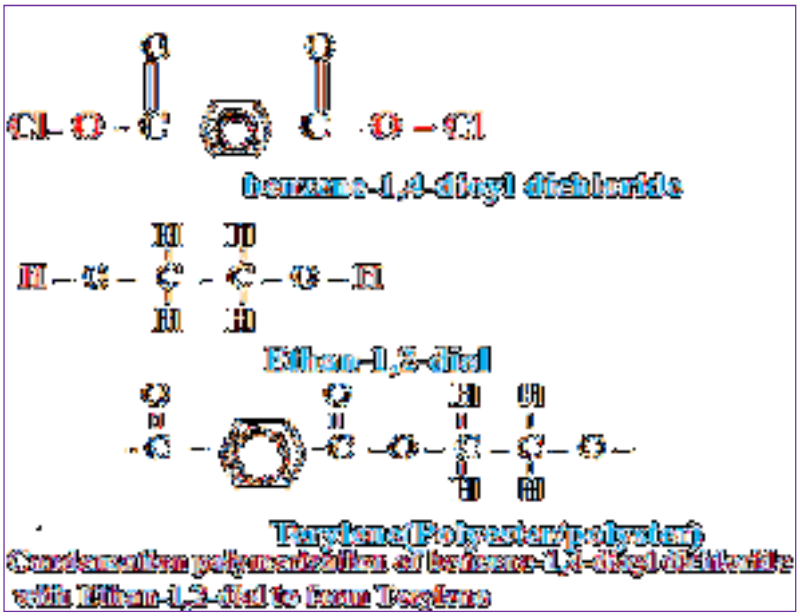


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Polymer bond linkage of terylene

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The commercial name of terylene is **Polyester /polyster** It is a a tough, elastic and durable plastic. It is used to make **clothes**, **plastic ropes and sails** and **plastic model kits.**

Practice questions Organic chemistry

1. A student mixed equal volumes of Ethanol and butanoic acid. He added a few drops of concentrated Sulphuric (VI) acid and warmed the mixture

(i) Name and write the formula of the main products

Name…….

Formula……..

(ii) Which homologous series does the product named in (i) above belong?

2. The structure of the monomer phenyl ethene is given below:-



a) Give the structure of the polymer formed when four of the monomers are added together

b) Give the name of the polymer formed in **(a)** above

3. Explain the environmental effects of burning plastics in air as a disposal method 4. Write chemical equation to represent the effect of heat on ammonium carbonate

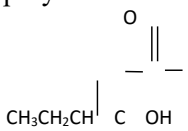
5. Sodium octadecanoate has a chemical formula $CH_3(CH_2)_6\ COO^-Na^+$, which is used as soap.

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Explain why a lot of soap is needed when washing with hard water

6. A natural polymer is made up of the monomer:



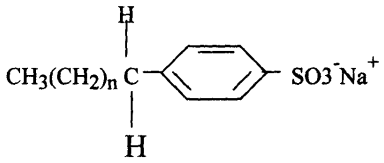
(a) Write the structural formula of the repeat unit of the polymer

(b) When 5.0×10^{-5} moles of the polymer were hydrolysed, 0.515g of the monomer were obtained.

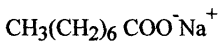
Determine the number of the monomer molecules in this polymer.

(C = 12; H = 1; N = 14; O =16)

7. The formula below represents active ingredients of two cleansing agents **A** and **B**

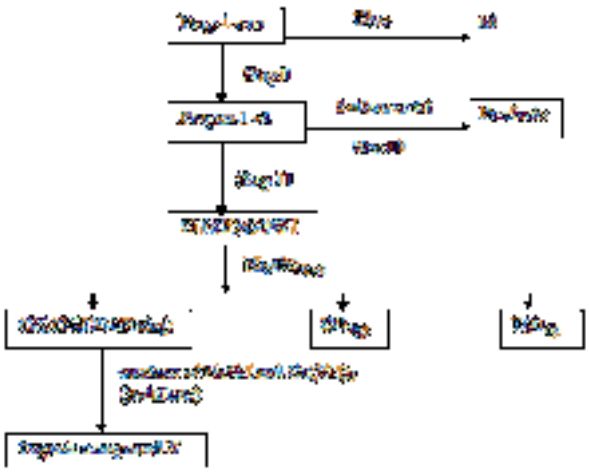


Agent A



Agent B

Which one of the cleansing agents would be suitable to be used in water containing magnesium hydrogen carbonate? Explain



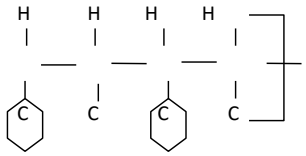
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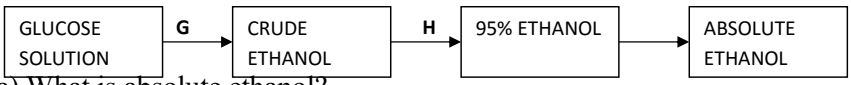
8. Study the polymer below and use it to answer the questions that follow:



(a) Give the name of the monomer and draw its structures
(b) Identify the type of polymerization that takes place
(c) State **one** advantage of synthetic polymers

9. Ethanol and Pentane are miscible liquids. Explain how water can be used to separate a mixture of ethanol and pentane

10.



(a) What is absolute ethanol?

(b) State **two** conditions required for process **G** to take place efficiently

11. (a) (i) The table below shows the volume of oxygen obtained per unit time when hydrogen peroxide was decomposed in the presence of manganese (IV) Oxide. Use it to answer the questions that follow:-

Time in seconds	**Volume of Oxygen evolved (cm^3)**
0	0
30	10
60	19
90	27
120	34
150	38
180	43
210	45
240	45
270	45

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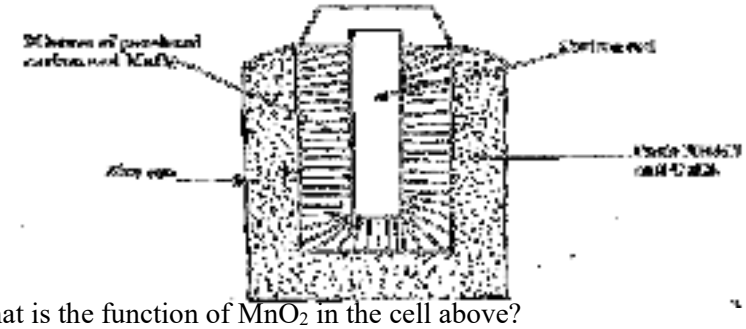
300	45

(i) Plot a graph of volume of oxygen gas against time

(ii) Determine the rate of reaction at time 156 seconds

(iii) From the graph, find the time taken for 18cm^3 of oxygen to be produced (iv) Write a chemical equation to show how hydrogen peroxide decomposes in the presence of manganese (IV) Oxide

(b) The diagram below shows how a Le'clanche (Dry cell) appears:-



(i) What is the function of MnO_2 in the cell above?

(ii) Write the equation of a reaction that occurs at the cathode

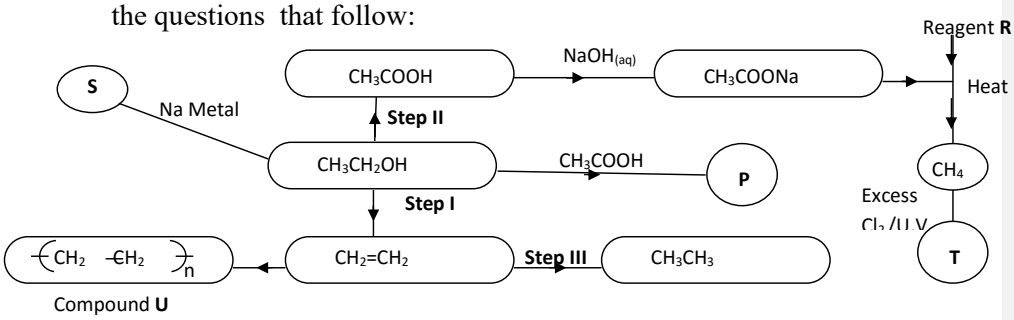
(iii) Calculate the mass of Zinc that is consumed when a current of 0.1amperes flows through the above cell for 30minutes (1F =96500c Zn =65)

12. (a) Give the IUPAC names of the following compounds:

(i) $CH_3COOCH_2CH_3$ *

(ii) $CH_2 = C - CHCH_3$ with a branch CH_3 on the middle carbon

(b) The structure below shows some reactions starting with ethanol. Study it and answer the questions that follow:



(i) Write the formula of the organic compounds **P** and **S**

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(ii) Name the type of reaction, the reagent(s) and condition for the reactions in the following steps :-

(I) Step I

(II) Step II

(III) Step III

(iii) Name reagent **R**

…… (iv) Draw the structural formula of **T** and give its name

(v) (I) Name compound **U**……..

(II) If the relative molecular mass of **U** is 42000, determine the value of n (**C**=12, **H**=1)

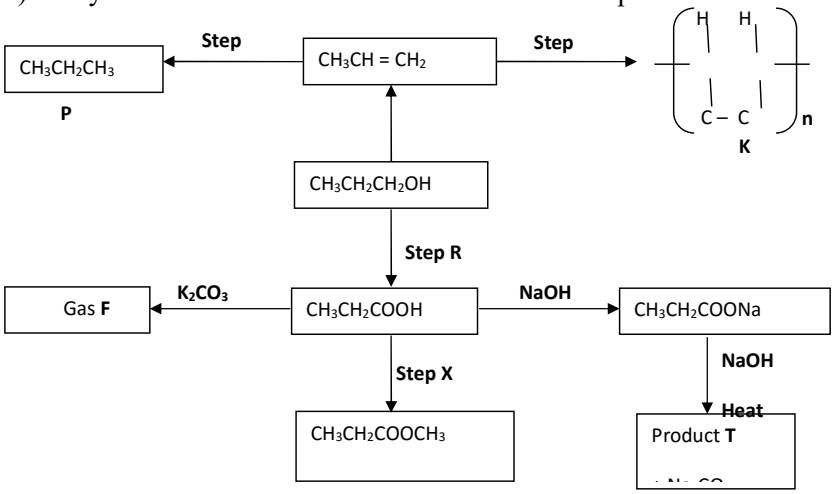
(c) State why C_2H_4 burns with a more smoky flame than C_2H_6

13. a) State **two** factors that affect the properties of a polymer

b) Name the compound with the formula below :

$CH_3CH_2CH_2ONa$

c) Study the scheme below and use it to answer the questions that follow:-



i) Name the following compounds:-

I. Product **T** …… II. **K** ……

ii) State **one** common physical property of substance **G**

iii) State the type of reaction that occurred in step **J**

iv) Give **one** use of substance **K**

v) Write an equation for the combustion of compound **P**

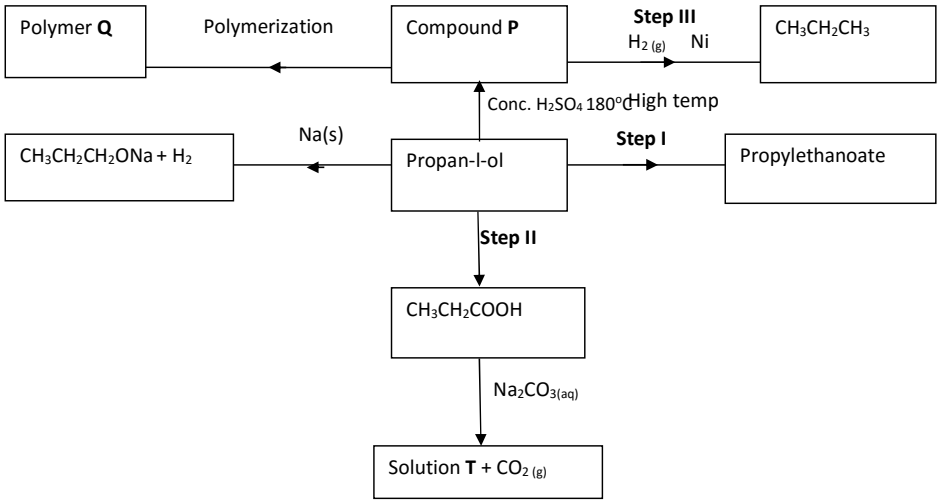
vi) Explain how compounds CH_3CH_2COOH and $CH_3CH_2CH_2OH$ can be distinguished chemically

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vii) If a polymer **K** has relative molecular mass of 12,600, calculate the value of **n** (H=1 C =12)

14. Study the scheme given below and answer the questions that follow:-



(a) (i) Name compound **P**

……

(ii) Write an equation for the reaction between CH_3CH_2COOH and Na_2CO_3

(b) State **one** use of polymer **Q**

(c) Name **one** oxidising agent that can be used in **step II**

……..

(d) A sample of polymer **Q** is found to have a molecular mass of 4200. Determine the number of

monomers in the polymer (H = 1, C = 12)

(e) Name the type of reaction in **step I**

……..

(f) State **one** industrial application of **step III**

(g)State how burning can be used to distinguish between propane and propyne. Explain your

answer

(h) 1000cm^3 of ethene (C_2H_4) burnt in oxygen to produce Carbon (II) Oxide and water vapour.

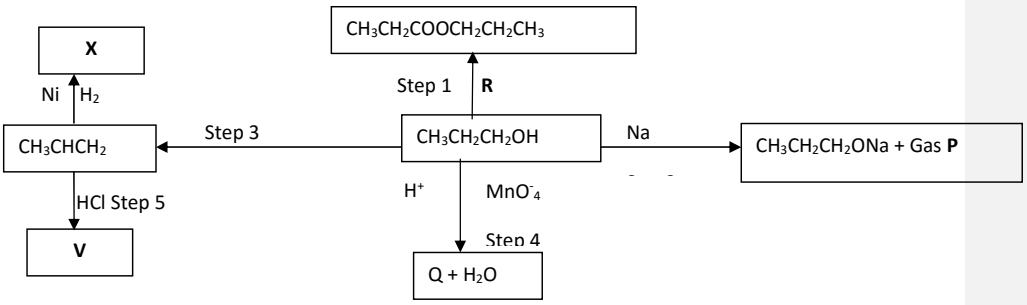
Calculate the minimum volume of air needed for the complete combustion of ethene

(Air contains 20% by volume of oxygen)

15. (a) Study the schematic diagram below and answer the questions that follow:-

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(i) Identify the following:

Substance **Q**

.....

Substance

R.....

Gas

P.....

(ii) Name:

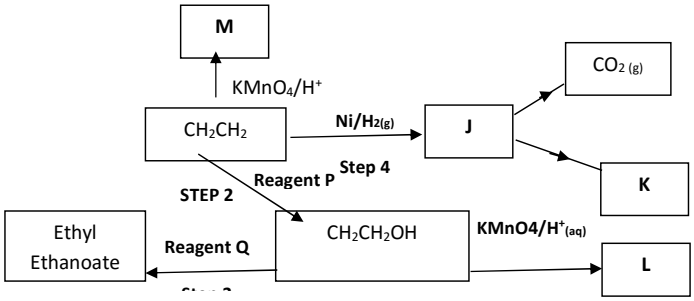
Step 1.....

Step 4.....

(iii) Draw the structural formula of the major product of step **5**

(iv) State the condition and reagent in step **3**

16. Study the flow chart below and answer the questions that follow



(a) (i) Name the following organic compounds:

M……..……..

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L……..

(ii) Name the process in step:

Step 2 ……..…

Step 4 ……..…

(iii) Identify the reagent **P** and **Q**

(iv) Write an equation for the reaction between $CH_3CH_2CH_2OH$ and sodium

17. a) Give the names of the following compounds:

i) $CH_3CH_2CH_2CH_2OH$

……

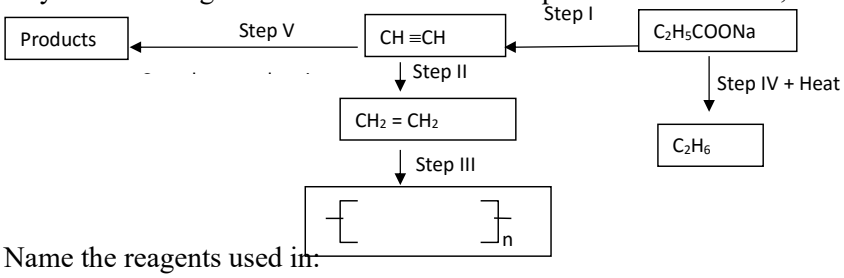
ii) CH_3CH_2COOH

……

iii) $CH_3C – O- CH_2CH_3$

……

18. Study the scheme given below and answer the questions that follow;



i) Name the reagents used in:

Step I:

……

Step II

……

Step III

……

ii) Write an equation to show products formed for the complete combustion of CH = CH

iii) Explain **one** disadvantage of continued use of items made from the compound formed in step III

19. A hydrated salt has the following composition by mass. Iron 20.2 %, oxygen 23.0%, sulphur 11.5%, water 45.3%

i) Determine the formula of the hydrated salt (Fe=56, S=32, O=16, H=11)

ii) 6.95g of the hydrated salt in **c(i)** above were dissolved in distilled water and the total volume made to 250cm^3 of solution. Calculate the concentration of the resulting salt solution in moles per litre. (Given that the molecula mass of the salt is 278)

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20. Write an equation to show products formed for the complete combustion of CH = CH

iii) Explain **one** disadvantage of continued use of items made from the compound formed in step III

21. Give the IUPAC name for each of the following organic compounds;

i) $CH_3 - CH(OH) - CH_2 - CH_3$

ii) $CH_3 - CH(C_2H_5) - CH_2 - CH_2 - CH_3$

iii) $CH_3COOCH_2CH_2CH_3$

22. The structure below represents a cleansing agent.

$$R - S(=O)(=O) - O^-Na^+$$

a) State the type of cleansing agent represented above

b) State **one** advantage and one disadvantage of using the above cleansing agent.

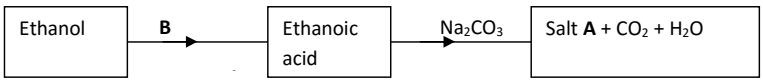
23. The structure below shows part of polymer .Use it to answer the questions that follow.

$$-CH(CH_3) - CH_2 - CH(CH_3) - CH_2 - CH(CH_3) - CH_2-$$

a) Derive the structure of the monomer

b) Name the type of polymerization represented above

24. The flow chart below represents a series of reactions starting with ethanoic acid:-



(a) Identify substances **A** and **B**

(b) Name the process **I**

25. a) Write an equation showing how ammonium nitrate may be prepared starting with ammonia gas

(b) Calculate the maximum mass of ammonium nitrate that can be prepared using 5.3kg of ammonia (H=1, N=14, O=16)

26. (a) What is meant by the term, esterification?

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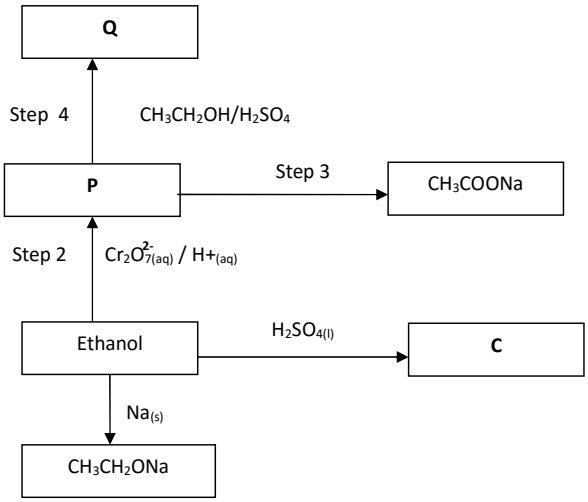
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(b) Draw the structural formulae of **two** compounds that may be reacted to form ethylpropanoate

27. (a) Draw the structure of pentanoic acid

(b) Draw the structure and give the name of the organic compound formed when ethanol reacts with pentanoic acid in presence of concentrated sulphuric acid

28. The scheme below shows some reactions starting with ethanol. Study it and answer the questions that follow:-



(i) Name and draw the structure of substance **Q**

(ii) Give the names of the reactions that take place in **steps 2** and **4**

(iii) What reagent is necessary for reaction that takes place in step 3

29. Substances **A** and **B** are represented by the formulae **ROH** and **RCOOH** respectively.

They belong to two different homologous series of organic compounds. If both A and B react with potassium metal:

(a) Name the common product produced by both

(b) State the observation made when each of the samples **A** and **B** are reacted with sodium hydrogen carbonate

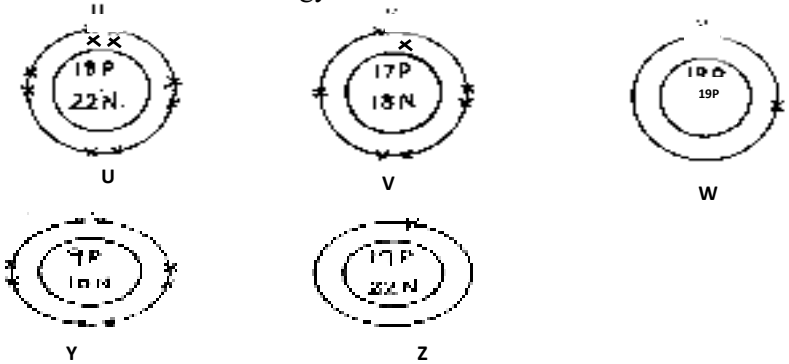
(i) **A**

(ii) **B**

30. Below are structures of particles. Use it to answer questions that follow. In each case only electrons in the outermost energy level are shown

key

P = Proton



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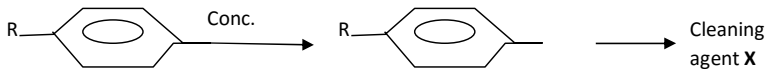
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N = Neutron
X = Electron

(a) Identify the particle which is an anion

31. Plastics and rubber are extensively used to cover electrical wires.
(a) What term is used to describe plastic and rubbers used in this way?
(b) Explain why plastics and rubbers are used this way

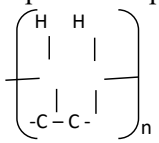
32. The scheme below represents the manufacture of a cleaning agent **X**



(a) Draw the structure of **X** and state the type of cleaning agent to which **X** belong (b) State **one** disadvantage of using **X** as a cleaning agent

33. **Y** grams of a radioactive isotope take 120days to decay to 3.5grams. The half-life period of the isotope is 20days
(a) Find the initial mass of the isotope
(b) Give **one** application of radioactivity in agriculture

34. The structure below represents a polymer. Study and answer the questions that follow:-

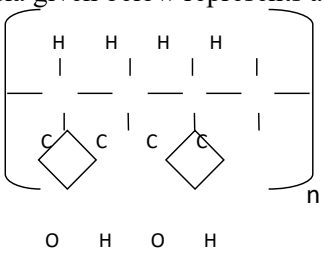


(i) Name the polymer above..... (ii) Determine the value of **n** if giant molecule had relative molecular mass of 4956

35. $RCOO^-Na^+$ and $RCH_2OSO_3^-Na^+$ are two types of cleansing agents;
i) Name the class of cleansing agents to which each belongs
ii) Which one of these agents in **(i)** above would be more suitable when washing with water from the Indian ocean. Explain

iii) Both sulphur (IV) oxide and chlorine are used bleaching agents. Explain the difference in their bleaching properties

36. The formula given below represents a portion of a polymer



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(a) Give the name of the polymer

(b) Draw the structure of the monomer used to manufacture the polymer

THE MOLE

(a)Gas laws

1. Matter is made up of small particle in accordance to Kinetic Theory of matter:
Naturally, there are basically **three** states of matter: **Solid, Liquid** and **gas:**
(i)A solid is made up of particles which are very closely packed with a definite/fixed shape and fixed/definite volume /occupies definite space. It has a very high density.
(ii) A liquid is made up of particles which have some degree of freedom. It thus has no definite/fixed shape. It takes the shape of the container it is put. A liquid has fixed/definite volume/occupies definite space.
(iii)A gas is made up of particles free from each other. It thus has no definite /fixed shape. It takes the shape of the container it is put. It has no fixed/definite volume/occupies every space in a container.

2.Gases are affected by **physical conditions**. There are **two** physical conditions:
(i)**Temperature**
(ii)**Pressure**

3. The SI unit of temperature is **Kelvin**(**K**).
Degrees Celsius/Centigrade(**^{o}C**) are also used.
The two units can be interconverted from the relationship:

^{o}C + 273= K
K -273 = ^{o}C

Practice examples

1. Convert the following into Kelvin.

(i) O **^{o}C**
^{o}C + 273 = K substituting : O ^{o}C + 273 = **273 K**

(ii) -273 **^{o}C**
^{o}C + 273 = K substituting : -273°C + 273 = **0 K**

(iii) 25 **^{o}C**
^{o}C + 273 = K substituting : 25 ^{o}C + 273 = **298 K**

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(iv) 100 **°C**

°C + 273 = K substituting : 100 °C + 273 = **373 K**

2. Convert the following into degrees Celsius/Centigrade(°C).

(i) 10 K

K -273 = °C substituting: 10 – 273 = **-263 °C**

(ii) (i) 1 K

K -273 = °C substituting: 1 – 273 = **-272 °C**

(iii) 110 K

K -273 = °C substituting: **110 – 273 = -163 °C**

(iv) -24 K

K -273 = °C substituting: **-24 – 273 = -297 °C**

The **standard** temperature is **273K** = **0 °C**.

The **room** temperature is assumed to be **298K** = **25°C**

4. The SI unit of pressure is Pascal(**Pa**) / Newton per metre squared (**Nm^{-2}**) .

Millimeters' of mercury(**mmHg**) ,centimeters of mercury(**cmHg**) and **atmospheres** are also commonly used.

The units are **not** interconvertible but Pascals(Pa) are equal to Newton per metre squared(Nm^{-2}).

The **standard** pressure is the **atmospheric** pressure.

Atmospheric pressure is **equal** to about:

(i)101325 Pa

(ii)101325 Nm^{-2}

(iii)760 mmHg

(iv)76 cmHg

(v)one atmosphere.

5. Molecules of gases are always in continuous random motion at high speed. This motion is affected by the physical conditions of temperature and pressure.

Physical conditions change the volume occupied by gases in a **closed** system.

The effect of physical conditions of temperature and pressure was investigated and expressed in both Boyles and Charles laws.

6. Boyles law states that

"**the volume of a fixed mass of a gas is inversely proportional to the pressure at constant/fixed temperature** "

Mathematically:

$$\text{Volume } \alpha \frac{1}{\text{Pressure}} \quad \textbf{(Fixed /constant Temperature)}$$

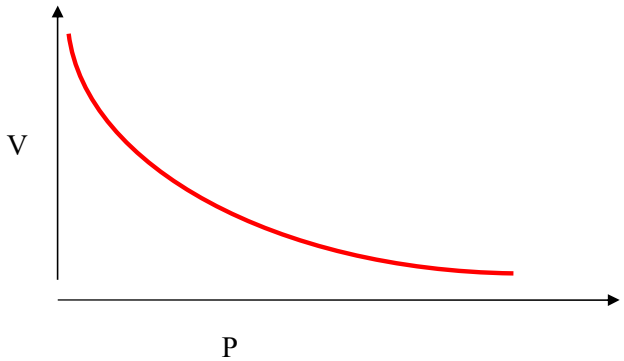
$$V \alpha \frac{1}{P} \quad \textbf{(Fixed /constant T)} \text{ ie } \textbf{PV = Constant}(k)$$

From Boyles law , an **increase** in pressure of a gas cause a **decrease** in volume. i.e **doubling** the pressure cause the volume to be **halved**.

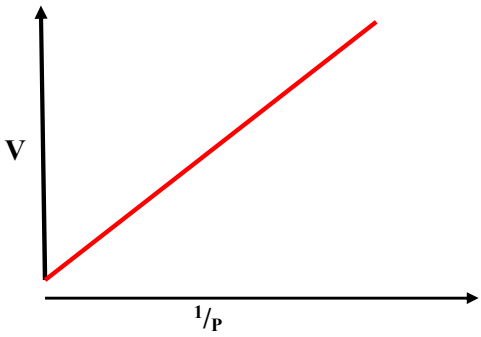
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Graphically therefore a plot of volume(**V**) **against** pressure (**P**) produces a **curve**.



Graphically a plot of volume(**V**) **against** inverse/reciprocal of pressure (**1/p**) produces a **straight line**



For **two** gases then $\mathbf{P_1 V_1 = P_2 V_2}$

P_1 = Pressure of gas 1
V_1 = Volume of gas 1
P_2 = Pressure of gas 2
V_2 = Volume of gas 2

Practice examples:

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1. A fixed mass of gas at 102300Pa pressure has a volume of 25cm3.Calculate its volume if the pressure is doubled.

Working

$P_1 V_1 = P_2 V_2$ Substituting :$102300 \times 25 = (102300 \times \mathbf{2}) \times V_2$

$V_2 = \frac{102300 \times 25}{(102300 \times 2)} = \mathbf{12.5cm3}$

2. Calculate the pressure which must be applied to a fixed mass of 100cm3 of Oxygen for its volume to triple at 100000Nm^{-2}.

$P_1 V_1 = P_2 V_2$ Substituting :$100000 \times 100 = P_2 \times (100 \times \mathbf{3})$

$V_2 = \frac{100000 \times 100}{(100 \times 3)} = \mathbf{33333.3333\ Nm^{-2}}$

3.A 60cm3 weather ballon full of Hydrogen at atmospheric pressure of 101325Pa was released into the atmosphere. Will the ballon reach stratosphere where the pressure is 90000Pa?

$P_1 V_1 = P_2 V_2$ Substituting :$101325 \times 60 = 90000 \times V_2$

$V_2 = \frac{101325 \times 60}{90000} = \mathbf{67.55\ cm3}$

The new volume at 67.55 cm3 **exceed** ballon capacity of 60.00 cm3.It will **burst** before reaching destination.

7.Charles law states that"**the volume of a fixed mass of a gas is directly proportional to the absolute temperature at constant/fixed pressure** "

Mathematically:

Volume α Pressure **(Fixed /constant pressure)**

$V \alpha T$ **(Fixed /constant P)** ie $\frac{V}{T}$ **= Constant**(k)

From Charles law , an **increase** in temperature of a gas cause an **increase** in volume. i.e **doubling** the temperature cause the volume to be **doubled**.

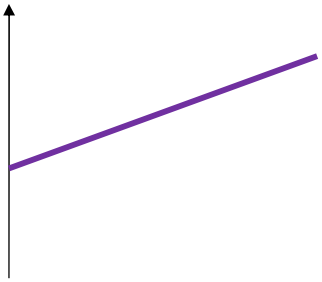
Gases expand/increase by $^1/_{273}$ by volume on heating.Gases contact/decrease by $^1/_{273}$ by volume on cooling at constant/fixed pressure.

The volume of a gas continue decreasing with decrease in temperature until at **-273°C /0 K** the volume is **zero**. i.e. there is no gas.

This temperature is called **absolute zero.** It is the **lowest** temperature at which a gas **can** exist.

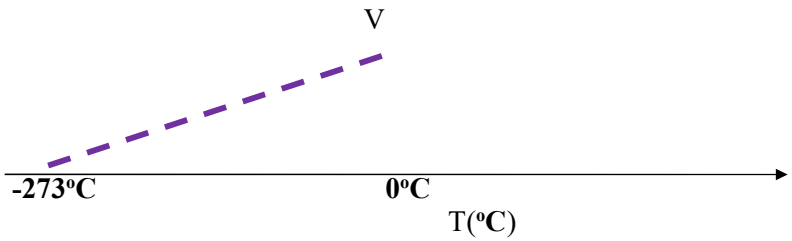
Graphically therefore a plot of volume(**V**) **against** Temperature(**T**) in:

(**i**)°C produces a **straight line** that is **extrapolated** to the absolute zero of -273°**C** .

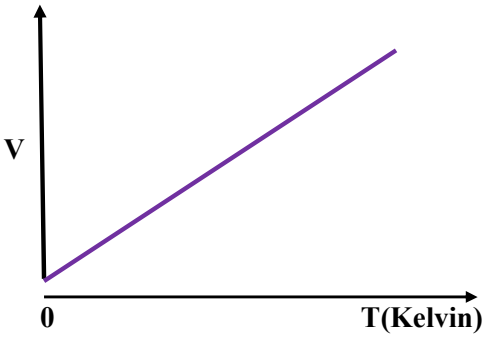


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(**ii**)Kelvin/K produces a **straight line** from absolute zero of **O Kelvin**



For **two** gases then $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

T_1 = Temperature **in Kelvin** of gas 1
V_1 = Volume of gas 1
T_2 = Temperature **in Kelvin** of gas 2
V_2 = Volume of gas 2

Practice examples:

1. 500cm3 of carbon(IV)oxide at 0°C was transfered into a cylinder at -4°C. If the capacity of the cylinder is 450 cm3,explain what happened.

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ substituting $\frac{500}{(0+273)} = \frac{V_2}{(-4+273)}$

$= \frac{500 \times (-4 \times 273)}{(0 + 273)} =$ **492.674cm3**

The capacity of cylinder (500cm3) is **less** than new volume(492.674cm3).
7.326cm3(500-492.674cm3)of carbon(IV)oxide gas did not fit into the cylinder.

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2. A mechanic was filling a deflated tyre with air in his closed garage using a hand pump. The capacity of the tyre was 40,000cm3 at room temperature. He rolled the tyre into the car outside. The temperature outside was 30°C.Explain what happens.

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ substituting $\frac{40000}{(\mathbf{25}+273)} = \frac{V_2}{(\mathbf{30}+273)}$

$= \frac{40000 \times (30 \times 273)}{(25 + 273)} = \mathbf{40671.1409cm3}$

The capacity of a tyre (40000cm3) is **less** than new volume(40671.1409cm3).
The tyre thus bursts.

3. A hydrogen gas balloon with 80cm3 was released from a research station at room temperature. If the temperature of the highest point it rose is -30°C , explain what happened.

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ substituting $\frac{80}{(\mathbf{25}+273)} = \frac{V_2}{(\mathbf{-30}+273)}$

$= \frac{80 \times (-30 \times 273)}{(25 + 273)} = \mathbf{65.2349cm3}$

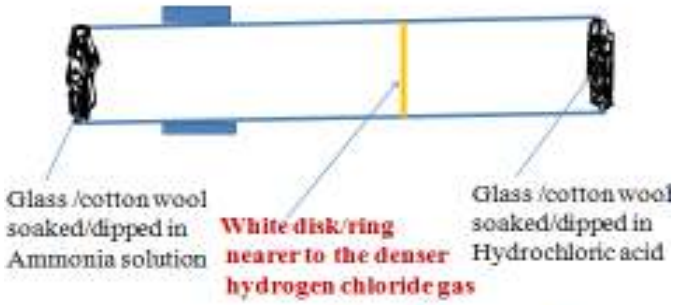
The capacity of balloon (80cm3) is **more** than new volume (65.2349cm3).
The balloon thus remained intact.

8. The continuous random motion of gases differ from gas to the other.The movement of molecules (of a gas) from region of high concentration to a region of low concentration is called **diffusion.**
The rate of diffusion of a gas depends on its density. i.e. **The higher the rate of diffusion, the less dense the gas**.
The density of a gas depends on its molar mass/relative molecular mass. i.e. **The** higher **the density the** higher **the molar mass/relative atomic mass and thus the** lower **the rate of diffusion.**
Examples
1.Carbon (IV)oxide(CO_2) has a molar mass of 44g.Nitrogen(N_2)has a molar mass of 28g. (N_2)is thus lighter/less dense than Carbon (IV)oxide(CO_2). N_2 diffuses faster than CO_2.
2.Ammonia(NH_3) has a molar mass of 17g.Nitrogen(N_2)has a molar mass of 28g. (N_2)is thus about **twice** lighter/less dense than Ammonia(NH_3). Ammonia(NH_3) diffuses twice faster than N_2.
3. Ammonia(NH_3) has a molar mass of 17g.Hydrogen chloride gas has a molar mass of 36.5g.Both gases on contact react to form **white fumes** of ammonium chloride .When a glass/cotton wool dipped in ammonia and another glass/cotton wool dipped in hydrochloric acid are placed at opposite ends of a glass tube, both gases diffuse towards each other. A white disk appears near to glass/cotton wool dipped in hydrochloric acid. This is because hydrogen chloride is heavier/denser than Ammonia and thus its rate of diffusion is lower .

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Diffusion of ammonia and hydrogen chloride



Chemical equation: $HCl(g) + NH_3(s) \rightarrow NH_3Cl(s)$

The rate of diffusion of a gas is in accordance to **Grahams law of diffusion**.
Grahams law states that:
"the rate of diffusion of a gas is inversely proportional to the square root of its density, at the same/constant/fixed temperature and pressure"
Mathematically
$R \propto \frac{1}{\sqrt{p}}$ and since density is proportional to mass then $R \propto \frac{1}{\sqrt{m}}$
For two gases then:
$\frac{R_1}{\sqrt{M_2}} = \frac{R_2}{\sqrt{M_1}}$ where: R_1 and R_2 is the **rate** of diffusion of **1**st and **2**nd gas.
M_1 and M_2 is the **molar mass** of **1**st and **2**nd gas.
Since rate is inverse of time. i.e. the higher the rate the less the time:
For two gases then:
$\frac{T_1}{\sqrt{M_1}} = \frac{T_2}{\sqrt{M_2}}$ where: T_1 and T_2 is the **time taken** for **1**st and **2**nd gas to diffuse.
M_1 and M_2 is the **molar mass** of **1**st and **2**nd gas.
Practice examples:
1. It takes 30 seconds for 100cm3 of carbon(IV)oxide to diffuse across a porous plate. How long will it take 150cm3 of nitrogen(IV)oxide to diffuse across the same plate under the same conditions of temperature and pressure. (C=12.0,N=14.0=16.0)

Molar mass CO_2=44.0 Molar mass NO_2=46.0

Method 1

100cm3 CO_2 takes 30seconds
150cm3 takes $\frac{150 \times 30}{100}$ = 45seconds

$\frac{T\ CO_2}{T\ NO_2} = \frac{\sqrt{\text{molar mass } CO_2}}{\sqrt{\text{molar mass } NO_2}} \Rightarrow \frac{45\text{seconds}}{T\ NO_2} = \frac{\sqrt{44.0}}{\sqrt{46.0}}$

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$T\ NO_2 = 45 seconds \times \frac{\sqrt{46.0}}{\sqrt{44.0}} = \mathbf{46.0114}$ seconds

Method 2

$100cm^3\ CO_2$ takes 30seconds

$1cm^3$ takes $\frac{100 \times 1}{30} = \mathbf{3.3333cm^3sec^{-1}}$

$\frac{R\ CO_2}{R\ NO_2} = \frac{\sqrt{\text{molar mass } NO_2}}{\sqrt{\text{molar mass } CO_2}} \Rightarrow \frac{3.3333cm^3sec^{-1}}{R\ NO_2} = \frac{\sqrt{46.0}}{\sqrt{44.0}}$

$R\ NO_2 = 3.3333cm^3sec^{-1} \times \frac{\sqrt{44.0}}{\sqrt{46.0}} = \mathbf{3.2601}cm^3sec^{-1}$

$3.2601cm^3$ takes 1seconds

$150cm^3$ take $\frac{150cm^3}{3.2601cm^3} = \mathbf{46.0109seconds}$

2. How long would 200cm3 of Hydrogen chloride take to diffuse through a porous plug if carbon(IV)oxide takes 200seconds to diffuse through.

Molar mass $CO_2 = 44g$ Molar mass $HCl = 36.5g$

$\frac{T\ CO_2}{T\ HCl} = \frac{\sqrt{\text{molar mass } CO_2}}{\sqrt{\text{molar mass } HCl}} \Rightarrow \frac{200 \text{ seconds}}{T\ HCl} = \frac{\sqrt{44.0}}{\sqrt{36.5}}$

$T\ HCl = 200seconds \times \frac{\sqrt{36.5}}{\sqrt{44.0}} = \mathbf{182.1588}$ seconds

3. Oxygen gas takes 250 seconds to diffuse through a porous diaphragm. Calculate the molar mass of gas Z which takes 227 second to diffuse.

Molar mass $O_2 = 32g$ Molar mass $Z = x\ g$

$\frac{T\ O_2}{T\ Z} = \frac{\sqrt{\text{molar mass } O_2}}{\sqrt{\text{molar mass } Z}} \Rightarrow \frac{250 \text{ seconds}}{227seconds} = \frac{\sqrt{32.0}}{\sqrt{x}}$

$\sqrt{x} = 227seconds \times \frac{\sqrt{32}}{250} = \mathbf{26.3828}$ grams

4. 25cm3 of carbon(II)oxide diffuses across a porous plate in 25seconds. How long will it take 75cm3 of Carbon(IV)oxide to diffuse across the same plate under the same conditions of temperature and pressure. (C=12.0,0=16.0)

Molar mass $CO_2 = 44.0$ Molar mass $CO = 28.0$

Method 1

$25cm^3\ CO$ takes 25seconds

$75cm^3$ takes $\frac{75 \times 25}{25} = 75seconds$

$\frac{T\ CO_2}{T\ CO} = \frac{\sqrt{\text{molar mass } CO_2}}{\sqrt{\text{molar mass } CO}} \Rightarrow \frac{T\ CO_2 seconds}{75} = \frac{\sqrt{44.0}}{\sqrt{28.0}}$

$T\ CO_2 = 75seconds \times \frac{\sqrt{44.0}}{\sqrt{28.0}} = \mathbf{94.0175}$ seconds

Method 2

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25cm3 CO_2 takes 25seconds
1cm3 takes $25 \times \frac{1}{25} = \mathbf{1.0cm3sec^{-1}}$

$\frac{R\ CO_2}{R\ CO} = \frac{\sqrt{\text{molar mass CO}}}{\sqrt{\text{molar mass } CO_2}} \Rightarrow \frac{x\ cm3sec^{-1}}{1.0cm3sec^{-1}} = \frac{\sqrt{28.0}}{\sqrt{44.0}}$

$R\ CO_2 = 1.0cm3sec^{-1} \times \frac{\sqrt{28.0}}{\sqrt{44.0}} = \mathbf{0.7977}cm3sec^{-1}$

0.7977cm3 takes 1 seconds
75cm3 takes $75cm3 \times \frac{}{0.7977cm3} = \mathbf{94.0203seconds}$

(b)Introduction to the mole, molar masses and Relative atomic masses

1. The mole is the **SI** unit of the **amount** of substance.
2. The number of particles e.g. atoms, ions, molecules, electrons, cows, cars are all measured in terms of moles.
3. The number of particles in one mole is called the **Avogadros Constant**. It is denoted "**L**".

The Avogadros Constant contain **6.023×10^{23}** particles. i.e.

1mole = 6.023×10^{23} particles = 6.023×10^{23}
2 moles = $2 \times 6.023 \times 10^{23}$ particles = 1.205×10^{24}
0.2 moles = $0.2 \times 6.023 \times 10\ 23$ particles = 1.205×10^{22}
0.0065 moles = $0.0065 \times 6.023 \times 10^{23}$ particles = 3.914×10^{21}

3. The mass of one mole of a substance is called **molar mass**. The molar mass of:

(i)an **element** has mass equal to relative **atomic** mass /RAM(in grams)of the element e.g.

Molar mass of carbon(C)= relative atomic mass = 12.0g
6.023×10^{23} particles of carbon = 1 mole =12.0 g
Molar mass of sodium(Na) = relative atomic mass = 23.0g
6.023×10^{23} particles of sodium = 1 mole =23.0 g
Molar mass of Iron (Fe) = relative atomic mass = 56.0g
6.023×10^{23} particles of iron = 1 mole =56.0 g

(ii)a **molecule** has mass equal to relative molecular mass /RMM (in grams)of the molecule. Relative molecular mass is the **sum** of the relative atomic masses of the elements making the molecule.

The number of atoms making a molecule is called **atomicity**. Most **gaseous** molecules are **diatomic (e.g. O_2, H_2, N_2, F_2, Cl_2, Br_2, I_2)**noble gases are **monoatomic(e.g. He, Ar, Ne, Xe)**,Ozone gas(**O_3**) is **triatomic** e.g.

Molar mass **Oxygen molecule(O_2)** =relative molecular mass =(16.0x 2)g =32.0g

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6.023 x10 ^{23} particles of Oxygen molecule = 1 mole = 32.0 g
Molar mass **chlorine molecule(Cl_2)** =relative molecular mass =(35.5x 2)g =71.0g
6.023 x10 ^{23} particles of chlorine molecule = 1 mole = 71.0 g
Molar mass **Nitrogen molecule(N_2)** =relative molecular mass =(14.0x 2)g =28.0g
6.023 x10 ^{23} particles of Nitrogen molecule = 1 mole = 28.0 g
(ii)a **compound** has mass equal to relative formular mass /RFM (in grams)of the molecule. Relative formular mass is the **sum** of the relative atomic masses of the elements making the compound. e.g.
(i)Molar mass **Water(H_2O)** = relative formular mass =[(1.0 x 2) + 16.0]g =18.0g
6.023 x10 ^{23} particles of Water molecule = 1 mole = 18.0 g
6.023 x10 ^{23} particles of Water molecule has:
- **2** x 6.023 x10 ^{23} particles of Hydrogen atoms
- **1** x 6.023 x10 ^{23} particles of Oxygen atoms

(ii)Molar mass **sulphuric(VI)acid(H_2SO_4)** = relative formular mass
=[(1.0 x 2) + 32.0 + (16.0 x 4)]g =98.0g
6.023 x10 ^{23} particles of sulphuric(VI)acid(H_2SO_4) = 1 mole = 98.0g
6.023 x10 ^{23} particles of sulphuric(VI)acid(H_2SO_4) has:
- **2** x 6.023 x10 ^{23} particles of **H**ydrogen atoms
- **1** x 6.023 x10 ^{23} particles of **S**ulphur atoms
- **4** x 6.023 x10 ^{23} particles of **O**xygen atoms

(iii)Molar mass **sodium carbonate(IV)(Na_2CO_3)** = relative formular mass
=[(23.0 x 2) + 12.0 + (16.0 x 3)]g =106.0g
6.023 x10 ^{23} particles of sodium carbonate(IV)(Na_2CO_3) = 1 mole = 106.0g
6.023 x10 ^{23} particles of sodium carbonate(IV)(Na_2CO_3) has:
- **2** x 6.023 x10 ^{23} particles of **S**odium atoms
- **1** x 6.023 x10 ^{23} particles of **C**arbon atoms
- **3** x 6.023 x10 ^{23} particles of **O**xygen atoms

(iv)Molar mass **Calcium carbonate(IV)($CaCO_3$)** = relative formular mass
=[(40.0+ 12.0 + (16.0 x 3)]g =100.0g.

6.023 x10 ^{23} particles of Calcium carbonate(IV)($CaCO_3$) = 1 mole = 100.0g
6.023 x10 ^{23} particles of Calcium carbonate(IV)($CaCO_3$) has:
- **1** x 6.023 x10 ^{23} particles of **C**alcium atoms
- **1** x 6.023 x10 ^{23} particles of **C**arbon atoms
- **3** x 6.023 x10 ^{23} particles of **O**xygen atoms

(v)Molar mass **Water(H_2O)** = relative formular mass
=[(2 x 1.0)+ 16.0]g =18.0g
6.023 x10 ^{23} particles of Water(H_2O) = 1 mole = 18.0g
6.023 x10 ^{23} particles of Water(H_2O) has:
- **2** x 6.023 x10 ^{23} particles of Hydrogen atoms

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-2 x 6.023 x10 ^{23} particles of **O**xygen atoms

Practice

1. Calculate the number of moles present in:

(i)0.23 g of Sodium atoms

Molar mass of Sodium atoms = 23g

Moles = mass in grams => 0.23g = **0.01moles**
Molar mass 23

(ii) 0.23 g of Chlorine atoms

Molar mass of Chlorine atoms = 35.5 g

Moles = mass in grams => 0.23g = **0.0065moles /6.5 x 10^{-3} moles**
Molar mass 35.5

(iii) 0.23 g of Chlorine molecules

Molar mass of Chlorine molecules =(35.5 x 2) = **71.0** g

Moles = mass in grams => 0.23g = **0.0032moles /3.2 x 10^{-3} moles**
Molar mass 71

(iv) 0.23 g of dilute sulphuric(VI)acid

Molar mass of H_2SO_4 = [(2 x 1) + 32 + (4 x14)] = **98.0**g

Moles = mass in grams => 0.23g = **0.0023moles /2.3 x 10^{-3} moles**
Molar mass 98

2. Calculate the number of atoms present in:(Avogadros constant L = 6.0 x 10 ^{23})

(i) 0.23 g of dilute sulphuric (VI)acid

Method I

Molar mass of H_2SO_4 = [(2 x 1) + 32 + (4 x14)] = **98.0**g

Moles = mass in grams => 0.23g = **0.0023moles /2.3 x 10^{-3} moles**
Molar mass 98

1 mole has 6.0 x 10 ^{23} atoms

2.3 x 10^{-3} moles has (2.3 x 10^{-3} x 6.0 x 10 ^{23}) = **1.38 x 10 ^{21}** atoms
1

Method II

Molar mass of H_2SO_4 = [(2 x 1) + 32 + (4 x14)] = **98.0**g

98.0g = 1 mole has 6.0 x 10 ^{23} atoms

0.23 g therefore has (0.23 g x 6.0 x 10 ^{23}) = **1.38 x 10 ^{21}** atoms
98

(ii)0.23 g of sodium carbonate(IV)decahydrate

Molar mass of $Na_2CO_3.10H_2O$=

[(2 x 23) + 12 + (3 x16) + (10 x 1.0) + (10 x 16)] = **276.0**g

Method I

Moles = mass in grams => 0.23g = **0.00083moles /**
Molar mass 276 **8.3 x 10^{-4} moles**

1 mole has 6.0 x 10 ^{23} atoms

8.3 x 10^{-4} moles has (8.3 x 10^{-4} moles x 6.0 x 10 ^{23}) = **4.98 x 10 ^{20}** atoms

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1

Method II

276.0g = 1 mole has 6.0×10^{23} atoms

0.23 g therefore has $\frac{(0.23\text{ g} \times 6.0 \times 10^{23})}{276.0}$ = **4.98×10^{20}** atoms

(iii)0.23 g of Oxygen gas

Molar mass of O_2 = (2 x16) = **32.0** g

Method I

Moles = $\frac{\text{mass in grams}}{\text{Molar mass}}$ => $\frac{0.23\text{g}}{32}$ = **0.00718moles / 7.18×10^{-3} moles**

1 mole has **2** x 6.0×10^{23} **atoms** in O_2

7.18×10^{-3}moles has $\frac{(7.18 \times 10^{-3}\text{moles} \times \mathbf{2} \times 6.0 \times 10^{23})}{1}$ =**8.616×10^{21}**atoms

Method II

32.0g = 1 mole has **2** x 6.0×10^{23} **atoms** in O_2

0.23 g therefore has $\frac{(0.23\text{ g} \times \mathbf{2} \times 6.0 \times 10^{23})}{32.0}$ = **8.616×10^{21}**atoms

(iv)0.23 g of Carbon(IV)oxide gas

Molar mass of CO_2 = [12 + (2 x16)] = **44.0** g

Method I

Moles = $\frac{\text{mass in grams}}{\text{Molar mass}}$ => $\frac{0.23\text{g}}{44}$ = **0.00522moles / 5.22×10^{-3} moles**

1 mole has **3** x 6.0×10^{23} **atoms** in CO_2

7.18×10^{-3}moles has $\frac{(5.22 \times 10^{-3}\text{moles} \times \mathbf{3} \times 6.0 \times 10^{23})}{1}$ =**9.396×10^{21}**atoms

Method II

44.0g = 1 mole has **3** x 6.0×10^{23} **atoms** in CO_2

0.23 g therefore has $\frac{(0.23\text{ g} \times 3 \times 6.0 \times 10^{23})}{44.0}$ = **9.409×10^{21}**atoms

(c)Empirical and molecular formula

1.The empirical formula of a compound is its simplest formula. It is the simplest whole number ratios in which atoms of elements combine to form the compound.

2.It is mathematically the lowest common multiple (LCM) of the atoms of the elements in the compound

3.Practically the empirical formula of a compound can be determined as in the following examples.

To determine the empirical formula of copper oxide

(a) Method 1:From copper to copper(II)oxide

Procedure.

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Weigh a clean dry covered crucible($\mathbf{M_1}$).Put two spatula full of copper **powder** into the crucible. Weigh again ($\mathbf{M_2}$).Heat the crucible on a strong Bunsen flame for five minutes. Lift the lid, and swirl the crucible carefully using a pair of tong. Cover the crucible and continue heating for another five minutes. Remove the lid and stop heating. Allow the crucible to cool. When cool replace the lid and weigh the contents again ($\mathbf{M_3}$).

Sample results

Mass of crucible($\mathbf{M_1}$)	15.6g
Mass of crucible + copper before heating($\mathbf{M_2}$)	18.4
Mass of crucible + copper after heating($\mathbf{M_3}$)	19.1

Sample questions

1. Calculate the mass of copper powder used.

Mass of crucible + copper before heating($\mathbf{M_2}$) = 18.4
Less Mass of crucible($\mathbf{M_1}$) = - 15.6g
Mass of copper **2.8 g**

2. Calculate the mass of Oxygen used to react with copper.

Method I

Mass of crucible + copper after heating($\mathbf{M_3}$) = 19.1g
Mass of crucible + copper before heating($\mathbf{M_2}$) = - 18.4g
Mass of Oxygen = **0.7 g**

Method II

Mass of crucible + copper after heating($\mathbf{M_3}$) = 19.1g
Mass of crucible = - 15.6g
Mass of copper(II)Oxide = 3.5 g
Mass of copper(II)Oxide = 3.5 g
Mass of copper = - 2.8 g
Mass of Oxygen = **0.7 g**

3. Calculate the number of moles of:

(i) copper used (Cu = 63.5)

number of moles of copper = $\frac{\text{mass used}}{\text{Molar mass}}$ => $\frac{2.8}{63.5}$ = **0.0441moles**

(ii) Oxygen used (O = 16.0)

number of moles of oxygen = $\frac{\text{mass used}}{\text{Molar mass}}$ => $\frac{0.7}{16.0}$ = **0.0441moles**

4. Determine the mole ratio of the reactants

Moles of copper = $\frac{0.0441\text{moles}}{0.0441\text{moles}}$ = $\frac{1}{1}$ => Mole ratio Cu: O = **1:1**
Moles of oxygen

5.What is the empirical, formula of copper oxide formed.

CuO (copper(II)oxide

6. State and explain the observations made during the experiment.

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Observation

Colour change from **brown** to **black**

Explanation

Copper powder is **brown**. On heating it reacts with oxygen from the air to form **black** copper(II)oxide

7. Explain why magnesium ribbon/shavings would be unsuitable in a similar experiment as the one above.

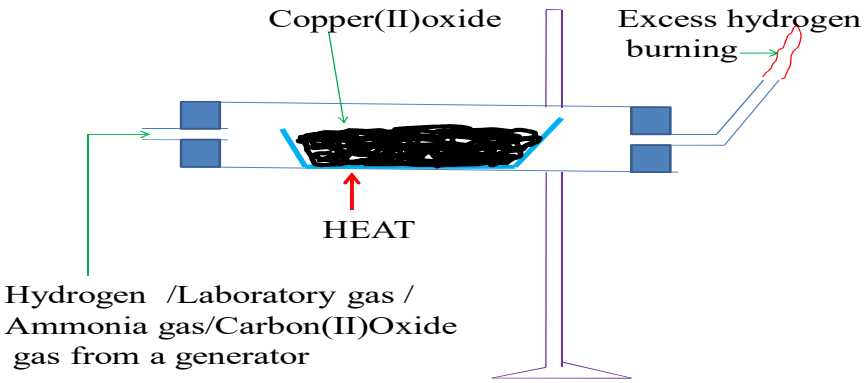
Hot magnesium generates enough heat energy to react with **both** Oxygen and Nitrogen in the air forming a white solid mixture of Magnesium **oxide** and magnesium **nitride**. This causes experimental mass errors.

(b) Method 2:From copper(II)oxide to copper

Procedure.

Weigh a clean dry porcelain boat ($\mathbf{M_1}$). Put two spatula full of copper(II)oxide **powder** into the crucible. Reweigh the porcelain boat ($\mathbf{M_2}$).Put the porcelain boat in a glass tube and set up the apparatus as below;

Determining empirical formula from copper(II)oxide **to copper**



Pass slowly(to prevent copper(II)oxide from being blown away)a stream of either dry Hydrogen /ammonia/laboratory gas/ carbon(II)oxide gas for about two minutes from a suitable generator.

When all the in the apparatus set up is driven out ,heat the copper(II)oxide strongly for about five minutes until there is no further change. Stop heating.

Continue passing the gases until the glass tube is cool.

Turn off the gas generator.

Carefully remove the porcelain boat form the combustion tube.

Reweigh ($\mathbf{M_3}$).

Sample results

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Mass of boat(M_1)	15.6g
Mass of boat before heating(M_2)	19.1
Mass of boat after heating(M_3)	18.4

Sample questions

1. Calculate the mass of copper(II)oxide used.

Mass of boat before heating(M_2) = 19.1
Mass of empty boat(M_1) = - 15.6g
Mass of copper(II)Oxide **3.5 g**

2. Calculate the mass of

(i) Oxygen.

Mass of boat before heating(M_2) = 19.1
Mass of boat after heating (M_3) = - 18.4g
Mass of oxygen = **0.7 g**

(ii)Copper

Mass of copper(II)Oxide = 3.5 g
Mass of oxygen = 0.7 g
Mass of oxygen = **2.8 g**

3. Calculate the number of moles of:

(i) Copper used (Cu = 63.5)

number of moles of copper $= \frac{\text{mass used}}{\text{Molar mass}} \Rightarrow \frac{2.8}{63.5} =$ **0.0441moles**

(ii) Oxygen used (O = 16.0)

number of moles of oxygen $= \frac{\text{mass used}}{\text{Molar mass}} \Rightarrow \frac{0.7}{16.0} =$ **0.0441moles**

4. Determine the mole ratio of the reactants

$\frac{\text{Moles of copper}}{\text{Moles of oxygen}} = \frac{0.0441\text{moles}}{0.0441\text{moles}} = \frac{1}{1} \Rightarrow$ Mole ratio Cu: O = **1:1**

5.What is the empirical, formula of copper oxide formed.

CuO (copper(II)oxide

6. State and explain the observations made during the experiment.

Observation

Colour change from **black** to **brown**

Explanation

Copper(II)oxide powder is **black**. On heating it is reduced by a suitable reducing agent to **brown** copper metal.

7. Explain why magnesium oxide would be unsuitable in a similar experiment as the one above.

Magnesium is high in the reactivity series. None of the above reducing agents is strong enough to reduce the oxide to the metal.

8. Write the equation for the reaction that would take place when the reducing agent is:

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(i) Hydrogen

$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$

(Black) (brown) (colourless liquid form on cooler parts)

(ii)Carbon(II)oxide

$CuO(s) + CO(g) \rightarrow Cu(s) + CO_2(g)$

(Black) (brown) (colourless gas, form white ppt with lime water)

(iii)Ammonia

$3CuO(s) + 2NH_3(g) \rightarrow 3Cu(s) + N_2(g) + 3H_2O(l)$

(Black) (brown) (colourless liquid form on cooler parts)

9. Explain why the following is necessary during the above experiment;

(i)A stream of dry hydrogen gas should be passed before heating copper (II) Oxide.

Air combine with hydrogen in presence of heat causing an explosion

(ii)A stream of dry hydrogen gas should be passed after heating copper (II) Oxide has been stopped.

Hot metallic copper can be re-oxidized back to copper(II)oxide

(iii) A stream of excess carbon (II)oxide gas should be ignited to burn

Carbon (II)oxide is highly poisonous/toxic. On ignition it burns to form less toxic carbon (IV)oxide gas.

10. State two sources of error in this experiment.

(i)All copper(II)oxide may not be reduced to copper.

(ii)Some copper(II)oxide may be blown out the boat by the reducing agent.

4.Theoreticaly the empirical formula of a compound can be determined as in the following examples.

(a)A oxide of copper contain 80% by mass of copper. Determine its empirical formula. (Cu = 63.5, 16.0)

% of Oxygen = 100% - % of Copper => 100- 80 = **20%** of Oxygen

Element	Copper	Oxygen
Symbol	Cu	O
Moles present = % composition / Molar mass	80 / 63.5	20 / 16
Divide by the smallest value	1.25 / 1.25	1.25 / 1.25
Mole ratios	1	1

Empirical formula is **CuO**

(b)1.60g of an oxide of Magnesium contain 0.84g by mass of Magnesium. Determine its empirical formula(Mg = 24.0, 16.0)

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Mass of Oxygen = 1.60 – 0.84 => **0.56 g** of Oxygen

Element	Magnesium	Oxygen
Symbol	Mg	O
Moles present = % composition / Molar mass	0.84 / 24	0.56 / 16
Divide by the smallest value	0.35 / 0.35	0.35 / 0.35
Mole ratios	1	1

Empirical formula is **MgO**

(c)An oxide of Silicon contain 47% by mass of Silicon. What is its empirical formula(Si = 28.0, 16.0)

Mass of Oxygen = 100 – 47 => **53%** of Oxygen

Element	Silicon	Oxygen
Symbol	Si	O
Moles present = % composition / Molar mass	47 / 28	53 / 16
Divide by the smallest value	1.68 / 1.68	3.31 / 1.68
Mole ratios	1	1.94 = 2

Empirical formula is **SiO_2**

(d)A compound contain 70% by mass of Iron and 30% Oxygen. What is its empirical formula(Fe = 56.0, 16.0)

Mass of Oxygen = 100 – 47 => **53%** of Oxygen

Element	Silicon	Oxygen
Symbol	Si	O
Moles present = % composition / Molar mass	47 / 28	53 / 16
Divide by the smallest value	1.68 / 1.68	3.31 / 1.68
Mole ratios	1	1.94 = 2

Empirical formula is **SiO_2**

2.During heating of a hydrated copper (II)sulphate(VI) crystals, the following readings were obtained:

Mass of evaporating dish =300.0g

Mass of evaporating dish + hydrated salt = 305.0g

Mass of evaporating dish + anhydrous salt = 303.2g

Calculate the number of water of crystallization molecules in hydrated copper (II)sulphate(VI)

(Cu =64.5, S = 32.0,O=16.0, H = 1.0)

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Working

Mass of Hydrated salt = 305.0g -300.0g = **5.0g**

Mass of anhydrous salt = 303.2 g -300.0g = **3.2 g**

Mass of water in hydrated salt = 5.0g -3.2 g = **1.8g**

Molar mass of water(H_2O) = **18.0g**

Molar mass of anhydrous copper (II)sulphate(VI) ($CuSO_4$) = **160.5g**

Element/compound	anhydrous copper (II) sulphate(VI)	Oxygen
Symbol	$CuSO_4$	O
Moles present = composition by mass / Molar mass	3,2 / 160.5	1.8 / 18
Divide by the smallest value	0.0199 / 0.0199	0.1 / 18
Mole ratios	1	5

The **empirical formula** of hydrated salt = $CuSO_4.\mathbf{5}H_2O$

Hydrated salt has **five/5 molecules** of water of crystallizations

4. The molecular formula is the actual number of each kind of atoms present in a **molecule** of a compound.

The empirical formula of an ionic compound is the same as the chemical formula but for simple molecular structured compounds, the empirical formula may not be the same as the chemical formula.

The molecular formula is a multiple of empirical formula .It is determined from the relationship:

(i) $\mathbf{n} = \frac{\text{Relative formular mass}}{\text{Relative empirical formula}}$

where **n** is a whole number.

(ii) Relative empirical formula x **n** = Relative formular mass where **n** is a whole number.

Practice sample examples

1. A hydrocarbon was found to contain 92.3% carbon and the remaining Hydrogen.

If the molecular mass of the compound is 78, determine the molecular formula(C=12.0, H =1.0)

Mass of Hydrogen = 100 – 92.3 => **7.7%** of Oxygen

Element	Carbon	Hydrogen
Symbol	C	H
Moles present = % composition / Molar mass	92.3 / 12	7.7 / 1
Divide by the smallest value	7.7	7.7

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	7.7	7.7
Mole ratios	1	1

Empirical formula is **CH**

The molecular formular is thus determined :

n = Relative formular mass = 78 = 6
Relative empirical formula 13

The molecular formula is (C H) x 6 = **C_6H_6**

2. A compound of carbon, hydrogen and oxygen contain 54.55% carbon, 9.09% and remaining 36.36% oxygen.
If its relative molecular mass is 88, determine its molecular formula(C=12.0, H =1.0, O= 16.0)

Element	Carbon	Hydrogen	Oxygen
Symbol	C	H	O
Moles present = % composition / Molar mass	54.55 / 12	9.09 / 1	36.36 / 16
Divide by the smallest value	4.5458 / 2.2725	9.09 / 2.2725	2.2725 / 2.2725
Mole ratios	2	4	1

Empirical formula is **C_2H_4O**

The molecular formula is thus determined :

n = Relative formular mass = 88 = 2
Relative empirical formula 44

The molecular formula is (C_2H_4O) x 2 = **$C_4H_8O_2$**.

4.A hydrocarbon burns completely in excess air to form 5.28 g of carbon (IV) oxide and 2,16g of water.
If the molecular mass of the hydrocarbon is 84, draw and name its molecular structure.

Since a hydrocarbon is a compound containing Carbon and Hydrogen only. Then:

Mass of carbon in CO_2 = Mass of C in CO_2 x mass of CO_2 =>
Molar mass of CO_2
12 x 5.28 = **1.44g**√
44

Mass of Hydrogen in H_2O = Mass of C in H_2O x mass of H_2O =>
Molar mass of H_2O
2 x 2.16 = **0.24g**√
18

Element	Carbon	Hydrogen
Symbol	C	H
Moles present = mass / Molar mass	1.44g / 12	0.24g√ / 1

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Divide by the smallest value	0.12 0.12	0.24 0.12
Mole ratios	1	2√

Empirical formula is **CH_2**√

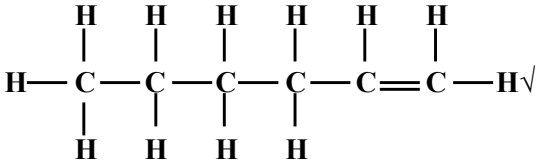
The molecular formular is thus determined :

n = Relative formular mass = 84 = **6**√
Relative empirical formula 14

The molecular formula is (CH_2) x 6 = **C_6H_{12}**. √

molecular name **Hexene√/Hex-1-ene (or any position isomer of Hexene)**

Molecular structure



5. Compound A contain 5.2% by mass of Nitrogen .The other elements present are Carbon, hydrogen and Oxygen. On combustion of 0.085g of A in excess Oxygen,0.224g of carbon(IV)oxide and 0.0372g of water was formed. Determine the empirical formula of A (N=14.0, O=16.0 , C=12.0 , H=1.0)

Mass of **N** in A = 5.2% x 0.085 = **0.00442 g**

Mass of C in A = $\frac{12}{44}$ x 0.224 = **0.0611g**

Mass of H in A = $\frac{2}{18}$ x 0.0372 = **0.0041g**

Mass of O in A = 0.085g – 0.004442g = **0.0806g** (Mass of C,H,O)

=> 0.0611g + 0.0041g = **0.0652g (Mass of C,H)**

0.0806g (Mass of C,H,O)- 0.0652g (Mass of C,H) = **0.0154 g**

Element	Nitrogen	Carbon	Hydrogen	Oxygen
Symbol	N	C	H	O
Moles present = $\frac{\text{mass}}{\text{Molar mass}}$	0.00442 g / 14	0.0611g / 12	0.0041g / 1	0.0154 g / 16
Divide by the smallest value	0.00032 0.00032	0.00509 0.00032	0.0041g 0.00032	0.00096 0.00032
Mole ratios	1	16	13	3

Empirical formula = $C_{16}H_{13}NO_3$

(d)Molar gas volume

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The volume occupied by one mole of all gases at the same temperature and pressure is a constant.It is:

(i) 24dm3/24litres/24000cm3 at **r**oom **t**emperature(25°C/298K)and **p**ressure(**r.t.p**).

i.e. 1mole of all gases =24dm3/24litres/24000cm3 at r.t.p

Examples

1mole of O_2 = 32g =6.0 x10^{23} particles**= 24dm3/24litres/24000cm3 at r.t.p**

1mole of H_2 = 2g =6.0 x10^{23} particles **=24dm3/24litres/24000cm3 at r.t.p**

1mole of CO_2 = 44g = 6.0 x10^{23} particles **=24dm3/24litres/24000cm3 at r.t.p**

1mole of NH_3 = 17g =6.0 x10^{23} particles **= 24dm3/24litres/24000cm3 at r.t.p**

1mole of CH_4 = 16g =6.0 x10^{23} particles **=24dm3/24litres/24000cm3 at r.t.p**

(**ii**)22.4dm3/22.4litres/22400cm3 at **s**tandard **t**emperature(0°C/273K) and **p**ressure(**s.t.p**)

i.e. 1mole of all gases =22.4dm3/22.4litres/22400cm3 at s.t.p

Examples

1mole of O_2 = 32g =6.0 x10^{23} particles**= 22.4dm3/22.4litres/22400cm3 at s.t.p**

1mole of H_2 = 2g =6.0 x10^{23} particles **= 22.4dm3/22.4litres/22400cm3 at s.t.p**

1mole of CO_2 = 44g = 6.0 x10^{23} particles **= 22.4dm3/22.4litres/22400cm3 at s.t.p**

1mole of NH_3 = 17g =6.0 x10^{23} particles**= 22.4dm3/22.4litres/22400cm3 at s.t.p**

1mole of CH_4 = 16g =6.0 x10^{23} particles **= 22.4dm3/22.4litres/22400cm3 at s.t.p**

The volume occupied by one mole of a gas at r.t.p or s.t.p is commonly called the **molar gas volume.** Whether the molar gas volume is at r.t.p or s.t.p must always be **specified**.

From the above therefore a less or more volume can be determined as in the examples below.

Practice examples

1. Calculate the number of particles present in:

(Avogadros constant =6.0 x10^{23}mole^{-1} **)**

(i) 2.24dm3 of Oxygen.

22.4dm3 -> 6.0 x10^{23}

2.24dm3 -> $\frac{2.24 \times 6.0 \times 10^{23}}{22.4}$

=6.0 x10^{22} **molecules** = 2 x 6.0 x$10^{22.}$ = 1.2 x10^{23} **atoms**

(ii) 2.24dm3 of Carbon(IV)oxide.

22.4dm3 -> 6.0 x10^{23}

2.24dm3 -> $\frac{2.24 \times 6.0 \times 10^{23}}{22.4}$

=6.0 x10^{22} **molecules** = (CO_2) = 3 x 6.0 x$10^{22.}$ = 1.8 x10^{23} **atoms**

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2. 0.135 g of a gaseous hydrocarbon X on complete combustion produces 0.41g of carbon(IV)oxide and 0.209g of water.0.29g of X occupy 120cm3 at room temperature and 1 atmosphere pressure .Name X and draw its molecular structure.(C=12.0,O= 16.O,H=1.0,1 mole of gas occupies 24dm3 at r.t.p)

Molar mass CO_2= 44 $gmole^{-1}$√ Molar mass H_2O = 18 $gmole^{-1}$√

Molar mass X = $\frac{0.29 \times (24 \times 1000)cm3}{120cm3}$ = 58 $gmole^{-1}$√

Since a hydrocarbon is a compound containing Carbon and Hydrogen only. Then:

Mass of carbon in CO_2 = $\frac{\text{Mass of C in } CO_2 \times \text{mass of } CO_2}{\text{Molar mass of } CO_2}$ =>

$\frac{12 \times 0.41}{44}$ = **0.1118g**√

Mass of Hydrogen in H_2O = $\frac{\text{Mass of C in } H_2O \times \text{mass of } H_2O}{\text{Molar mass of } H_2O}$ =>

$\frac{2 \times 0.209}{18}$ = **0.0232g**√

Element	Carbon	Hydrogen
Symbol	C	H
Moles present = $\frac{\%\text{ composition}}{\text{Molar mass}}$	$\frac{0.g118}{12}$	$\frac{0.0232g}{1}$√
Divide by the smallest value	0.0093 0.0093	0.0232 0.0093√
Mole ratios	1 x2	2.5x2
	2	5√

Empirical formula is **C_2H_5**√

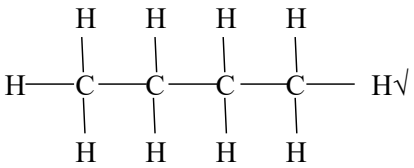
The molecular formular is thus determined :

n = $\frac{\text{Relative formular mass}}{\text{Relative empirical formula}}$ = $\frac{58}{29}$ = **2**√

The molecular formula is **(C_2H_5)** x 2 = **C_4H_{10}**.√

Molecule name **Butane**

Molecula structure



√

(e)**Gravimetric analysis**

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Gravimetric analysis is the relationship between reacting masses and the volumes and /or masses of products. All reactants are in mole **ratios** to their products in accordance to their stoichiometric equation. Using the mole ration of reactants and products any volume and/or mass can be determined as in the examples:

1. Calculate the volume of carbon(IV)oxide at r.t.p produced when 5.0 g of calcium carbonate is strongly heated.(Ca=40.0, C= 12.0,O = 16.0,1 mole of gas =22.4 at r.t.p)

Chemical equation

$CaCO_3(s)$ -> $CaO(s)$ + $CO_2(g)$

Mole ratios 1: 1: 1

Molar Mass $CaCO_3$ =100g

Method 1

100g $CaCO_3(s)$ -> 24dm3 $CO_2(g)$ at r.t.p

5.0 g $CaCO_3(s)$ -> $\frac{5.0\text{ g} \times 24\text{dm3}}{100\text{g}}$ = **1.2dm3/1200cm3**

Method 2

Moles of 5.0 g $CaCO_3(s)$ = $\frac{5.0\text{ g}}{100\text{ g}}$ = **0.05** moles

Mole ratio 1:1

Moles of $CO_2(g)$ = **0.05**moles

Volume of $CO_2(g)$ = 0.05 x 24000cm3 =**1200cm3 /1.2dm3**

2. 1.0g of an alloy of aluminium and copper were reacted with excess hydrochloric acid. If 840cm3 of hydrogen at s.t.p was produced, calculate the % of copper in the alloy.(Al =27.0,one mole of a gas at s.t.p =22.4dm3)

Chemical equation

Copper does not react with hydrochloric acid

$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$

Method 1

$3H_2(g)$ = 3 moles x (22.4 x 1000)cm3 => 2 x 27 g Al

840cm3 => $\frac{840\text{cm3} \times 2 \times 27}{3 \times 22.4 \times 1000}$ = **0.675**g of Aluminium

Total mass of alloy – mass of aluminium = mass of copper

=> 1.0g - 0.675g =**0.325**g of copper

% copper = $\frac{\text{mass of copper}}{\text{Mass of alloy}} \times 100\%$ = **32.5%**

Method 2

Mole ratio 2Al: $3H_2$ = 2:3

Moles of Hydrogen gas = $\frac{\text{volume of gas}}{\text{Molar gas volume}}$ => $\frac{840\text{cm3}}{22400\text{cm3}}$ = **0.0375moles**

Moles of Al = $^2/_3$ moles of H_2 => $^2/_3 \times 0.0375$moles = **0.025moles**

Mass of Al = moles x molar mass =>0.025moles x 27 = **0.675g**

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Total mass of alloy – mass of aluminium = mass of copper
=> 1.0g - 0.675g = **0.325** g of copper
% copper = mass of copper x100% = **32.5%**
Mass of alloy

(f)Gay Lussac's law
Gay Lussacs law states that **"when gases combine/react they do so in simple volume ratios to each other and to their gaseous products at constant/same temperature and pressure"**
Gay Lussacs law thus only apply to gases
Given the volume of one gas reactant, the other gaseous reactants can be deduced thus:
Examples
1. Calculate the volume of Oxygen required to completely react with 50cm3 of Hydrogen.
Chemical equation: $2H_2$ (g) + O_2 (g) -> $2H_2O$(l)
Volume ratios 2 : 1 : 0
Reacting volumes 50cm3 : **25cm3**
50cm3 of Oxygen is used

2. Calculate the volume of air required to completely reacts with 50cm3 of Hydrogen.(assume Oxygen is 21% by volume of air)
Chemical equation: $2H_2$ (g) + O_2 (g) -> $2H_2O$(l)
Volume ratios 2 : 1 : 0
Reacting volumes 50cm3 : **25cm3**
50cm3 of Oxygen is used
21% = 25cm3
100% = 100 x 25 =
21
3.If 5cm3 of a hydrocarbon C_xH_y burn in 15cm3 of Oxygen to form 10cm3 of Carbon(IV)oxide and 10cm3 of water vapour/steam, obtain the equation for the reaction and hence find the value of x and y in C_xH_y.
Chemical equation: C_xH_y (g) + O_2 (g) -> H_2O(g) + CO_2(g)
Volumes 5cm3 : 15cm3 : 10cm3 : 10cm3
Volume ratios 5cm3 : 15cm3 : 10cm3 : 10cm3
(divide by lowest volume) **5** **5** **5** **5**
Reacting volume ratios 1volume 3 volume 2 volume 2 volume
Balanced chemical equation: **C_xH_y (g) + $3O_2$ (g) -> $2H_2O$(g) + $2CO_2$(g)**
If "4H" are in $2H_2O$(g) the y=4

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If "2C" are in $2CO_2$ (g) the x=2
Thus(i) chemical formula of hydrocarbon = **C_2H_4**
(ii) chemical name of hydrocarbon = **Ethene**

4.100cm3 of nitrogen (II)oxide NO combine with 50cm3 of Oxygen to form 100cm3 of a single gaseous compound of nitrogen. All volumes measured at the same temperature and pressure. Obtain the equation for the reaction and name the gaseous product.

Chemical equation:	NO (g)	+	O_2 (g)	->	NO_x		
Volumes	100cm3	:	50cm3		: 100		
Volume ratios	100cm3	:	50cm3	:	100cm3	(divide	
by lowest volume)	**50**		**50**		**50**		
Reacting volume ratios	2volume		1 volume		2 volume		

Balanced chemical equation: 2 **NO (g) + O_2 (g) -> 2NO_x(g)**

Thus(i) chemical formula of the nitrogen compound = **2 NO_2**
(ii) chemical name of compound = **Nitrogen(IV)oxide**

5.When 15cm3 of a gaseous hydrocarbon was burnt in 100cm3 of Oxygen ,the resulting gaseous mixture occupied70cm3 at room temperature and pressure. When the gaseous mixture was passed through, potassium hydroxide its volume decreased to 25cm3.

(a)What volume of Oxygen was used during the reaction.(1mk)

Volume of Oxygen used =100-25 =75cm3√
(P was completely burnt)

(b)Determine the molecular formula of the hydrocarbon(2mk)

C_xH_y + O_2 -> xCO_2 + yH_2O
15cm3 : 75cm3
15 15
1 : 3√

=> 1 atom of C react with 6 (3x2)atoms of Oxygen
Thus x = 1 and y = 2 => P has molecula formula CH_4√

(g) Ionic equations

An ionic equation is a chemical statement showing the movement of ions (cations and anions) from reactants to products.

Solids, **gases** and **liquids** do not ionize/dissociate into free ions. **Only** ionic compounds in **aqueous/solution** or **molten** state ionize/dissociate into free cations and anions (**ions**)

An ionic equation is usually derived from a stoichiometric equation by using the following guidelines

Guidelines for writing ionic equations

1.Write the balanced stoichiometric equation
2.Indicate the state symbols of the reactants and products

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3.**Split** into cations and anions all the reactants and products that exist in **aqueous** state.
4.**Cancel out** any cation and anion that appear on **both** the product and reactant side.
5. Rewrite the chemical equation. It is an ionic equation.

Practice

(**a**)Precipitation of an insoluble salt

All insoluble salts are prepared in the laboratory from double decomposition /precipitation. This involves mixing **two soluble** salts to form **one soluble** and **one insoluble** salt

1. When silver nitrate(V) solution is added to sodium chloride solution,sodium nitrate(V) solution and a white precipitate of silver chloride are formed.

Balanced stoichiometric equation

$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl\ (\mathbf{s}) + NaNO_3\ (aq)$

Split reactants product existing in aqueous state as cation/anion

$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow AgCl(\mathbf{s}) + Na^+(aq) + NO_3^-(aq)$

Cancel out ions appearing on reactant and product side

$Ag^+(aq) + \cancel{NO_3^-(aq)} + \cancel{Na^+(aq)} + Cl^-(aq) \rightarrow AgCl(\mathbf{s}) + \cancel{Na^+(aq)} + \cancel{NO_3^-(aq)}$

Rewrite the equation

$\mathbf{Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)}$ **(**ionic equation**)**

2. When barium nitrate(V) solution is added to copper(II)sulphate(VI) solution, copper(II) nitrate(V) solution and a white precipitate of barium sulphate(VI) are formed.

Balanced stoichiometric equation

$Ba(NO_3)_2(aq) + CuSO_4(aq) \rightarrow BaSO_4\ (\mathbf{s}) + Cu(NO_3)_2\ (aq)$

Split reactants product existing in aqueous state as cation/anion

$Ba^{2+}(aq) + \mathbf{2}NO_3^-(aq) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4\ (\mathbf{s}) + \mathbf{2}NO_3^-(aq) + Cu^{2+}(aq)$

Cancel out ions appearing on reactant and product side

$Ba^{2+}(aq) + \cancel{\mathbf{2}NO_3^-(aq)} + \cancel{Cu^{2+}(aq)} + SO_4^{2-}(aq) \rightarrow BaSO_4(\mathbf{s}) + \cancel{\mathbf{2}NO_3^-(aq)} + Cu^{2+}(aq)$

Rewrite the equation

$\mathbf{Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)}$ **(**ionic equation**)**

3.A yellow precipitate of Potassium Iodide is formed from the reaction of Lead(II)nitrate(v) and potassium iodide.

Balanced stoichiometric equation

$Pb(NO_3)_2(aq) + 2KI\ (aq) \rightarrow PbI_2\ (\mathbf{s}) + 2KNO_3\ (aq)$

Split reactants product existing in aqueous state as cation/anion

$Pb^{2+}(aq) + \mathbf{2}NO_3^-(aq) + 2K^+(aq) + 2I^-(aq) \rightarrow PbI_2\ (\mathbf{s}) + \mathbf{2}NO_3^-(aq) + 2K^+(aq)$

Cancel out ions appearing on reactant and product side

$Pb^{2+}(aq) + \cancel{\mathbf{2}NO_3^-(aq)} + \cancel{2K^+(aq)} + 2I^-(aq) \rightarrow PbI_2\ (\mathbf{s}) + \cancel{\mathbf{2}NO_3^-(aq)} + \cancel{2K^+(aq)}$

Rewrite the equation

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$Pb^{2+}(aq) + 2I^-(aq)$ -> PbI_2 (s) (ionic equation)

(**b**)Neutralization

Neutralization is the reaction of an acid with a **soluble** base/alkali or **insoluble** base.

(i)Reaction of alkalis with acids

1.Reaction of nitric(V)acid with potassium hydroxide

Balanced stoichiometric equation

$HNO_3(aq) + KOH(aq) \rightarrow H_2O$ (**l**) + KNO_3 (aq)

Split reactants product existing in aqueous state as cation/anion

$H^+(aq) + NO_3^-(aq) + K^+(aq) + OH^-(aq) \rightarrow H_2O$ (**l**) + $NO_3^-(aq) + K^+(aq)$

Cancel out ions appearing on reactant and product side

$H^+(aq) + \cancel{NO_3^-(aq)} + \cancel{K^+(aq)} + OH^-(aq) \rightarrow H_2O$ (**l**) + $\cancel{NO_3^-(aq)} + \cancel{K^+(aq)}$

Rewrite the equation

$H^+(aq) + OH^-(aq)$ -> H_2O (l) (ionic equation)

2.Reaction of sulphuric(VI)acid with ammonia solution

Balanced stoichiometric equation

$H_2SO_4(aq) + 2NH_4OH(aq) \rightarrow H_2O$ (**l**) + $(NH_4)_2SO_4$ (aq)

Split reactants product existing in aqueous state as cation/anion

$2H^+(aq) + SO_4^{2-}(aq) + 2NH_4^+(aq) + 2OH^-(aq) \rightarrow 2H_2O$ (**l**) $+ SO_4^{2-}(aq) + 2NH_4^+(aq)$

Cancel out ions appearing on reactant and product side

$2H^+(aq) + \cancel{SO_4^{2-}(aq)} + \cancel{2NH_4^+(aq)} + 2OH^-(aq) \rightarrow 2H_2O$ (**l**) $+ \cancel{SO_4^{2-}(aq)} + \cancel{2NH_4^+(aq)}$

Rewrite the equation

$2H^+(aq) + 2OH^-(aq)$ -> $2H_2O$ (l)

$H^+(aq) + OH^-(aq)$ -> H_2O (l) (ionic equation)

3.Reaction of hydrochloric acid with Zinc hydroxide

Balanced stoichiometric equation

2HCl(aq) + $Zn(OH)_2$ (**s**) -> $2H_2O$ (**l**) + $ZnCl_2$ (aq)

Split reactants product existing in aqueous state as cation/anion

$2H^+(aq) + 2Cl^-(aq) + Zn(OH)_2$ (**s**) -> $2H_2O$ (**l**) + $2Cl^-(aq) + Zn^{2+}(aq)$

Cancel out ions appearing on reactant and product side

$2H^+(aq) + \cancel{2Cl^-(aq)} + Zn(OH)_2$ (**s**) -> $2H_2O$ (**l**) + $\cancel{2Cl^-(aq)} + Zn^{2+}(aq)$

Rewrite the equation

$2H^+(aq) + Zn(OH)_2$ (**s**) -> $2H_2O$ (**l**) + $Zn^{2+}(aq)$ **(ionic equation)**

(h)Molar solutions

A molar solution is one whose concentration is known. The SI unit of concentration is **Molarity** denoted **M**.

Molarity may be defined as the number of moles of solute present in **one** cubic **decimeter** of solution.

One cubic decimeter is equal to **one litre** and also equal to **1000cm3**.

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The higher the molarity the higher the concentration and the higher/more solute has been dissolved in the solvent to make one cubic decimeter/ litre/1000cm3 **solution.**

Examples

2M sodium hydroxide means 2 moles of sodium hydroxide solute is dissolved in enough water to make one cubic decimeter/ litre/1000cm3 uniform **solution** mixture of sodium hydroxide and water.

0.02M sodium hydroxide means 0.02 moles of sodium hydroxide solute is dissolved in enough water to make one cubic decimeter/ litre/1000cm3 uniform **solution** mixture of sodium hydroxide and water.

"2M" is **more concentrated** than"0.02M".

Preparation of molar solution

Procedure

Weigh accurately 4.0 g of sodium hydroxide pellets into a 250cm3 volumetric flask.

Using a wash bottle add about 200cm3 of distilled water.

Stopper the flask.

Shake vigorously for three minutes.

Remove the stopper for a second then continue to shake for about another two minutes until **all** the solid has dissolved.

Add more water slowly upto **exactly** the 250 cm3 **mark**.

Sample questions

1.Calculate the number of moles of sodium hydroxide pellets present in:

(i) 4.0 g.

Molar mass of NaOH = (23 + 16 + 1) = 40g

$$\text{Moles} = \frac{\text{Mass}}{\text{Molar mass}} \Rightarrow \frac{4.0}{40} = \mathbf{0.1} \text{ / } \mathbf{1.0 \times 10^{-1}} \text{ moles}$$

(ii) 250 cm3 solution in the volumetric flask.

Moles in 250 cm3 = **0.1 / 1.0 x 10 $^{-1}$** moles

(iii) one decimeter of solution

Method 1

Moles in decimeters = **Molarity** = $\frac{\text{Moles}}{\text{Volume of solution}} \times 1000\text{cm3/1dm3}$

$\Rightarrow 1.0 \times 10^{-1}$ moles $\times \frac{1000\text{cm3}}{250\text{cm3}} =$

= **0.4 M / 0.4 molesdm^{-3}**

Method 2

250cm3 solution contain 1.0×10^{-1} moles

1000cm3 solution = Molarity contain $\frac{1000}{250\text{ cm3}} \times 1.0 \times 10^{-1}$ moles

= **0.4 M / 0.4 molesdm^{-3}**

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Theoretical sample practice

1. Calculate the molarity of a solution containing:

(i) 4.0 g sodium hydroxide dissolved in 500cm3 solution

Molar mass of NaOH = (23 + 16 + 1) = 40g

Moles = Mass $\Rightarrow$ 4.0 = **0.1** / **1.0 x 10^{-1}** moles
Molar mass 40

Method 1

Moles in decimeters = **Molarity** = Moles x 1000cm3/1dm3
Volume of solution

$\Rightarrow$ 1.0 x 10^{-1} moles x 1000cm3
500cm3

= **0.2 M / 0.2 molesdm^{-3}**

Method 2

500 cm3 solution contain 1.0 x 10^{-1} moles

1000cm3 solution = Molarity contain 1000 x 1.0 x 10^{-1} moles
500 cm3

= **0.2 M / 0.2 molesdm^{-3}**

(ii) 5.3 g anhydrous sodium carbonate dissolved in 50cm3 solution

Molar mass of Na_2CO_3 = (23 x 2 + 12 + 16 x 3) = 106 g

Moles = Mass $\Rightarrow$ 5.3 = **0.05 / 5. 0 x 10^{-2}** moles
Molar mass 106

Method 1

Moles in decimeters = **Molarity** = Moles x 1000cm3/1dm3
Volume of solution

$\Rightarrow$ 1.0 moles x 1000cm3 =
50cm3

=**1.0 M**

Method 2

50 cm3 solution contain 5.0 x 10^{-2} moles

1000cm3 solution = Molarity contain 1000 x 5.0 x 10^{-2} moles
50 cm3

= **1.0M / 1.0 molesdm^{-3}**

(iii) 5.3 g hydrated sodium carbonate decahydrate dissolved in 50cm3 solution

Molar mass of $Na_2CO_3.10H_2O$ = (23 x 2 + 12 + 16 x 3 + 20 x 1 + 10 x 16) =286g

Moles = Mass $\Rightarrow$ 5.3 = **0.0185 / 1.85 x 10^{-2}** moles
Molar mass 286

Method 1

Moles in decimeters = **Molarity** = Moles x 1000cm3/1dm3
Volume of solution

$\Rightarrow$ 1.85 x 10^{-2} moles x 1000cm3 =

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50cm3

= **0.37 M/0.37 molesdm^{-3}**

Method 2

50 cm3 solution contain 1.85 x 10 $^{-2}$ moles

1000cm3 solution = Molarity contain 1000 x 1.85 x 10 $^{-2}$ moles

50 cm3

= **3.7 x 10^{-1} M / 3.7 x 10^{-1} molesdm^{-3}**

(iv) 7.1 g of anhydrous sodium sulphate(VI)was dissolved in 20.0 cm3 solution. Calculate the molarity of the solution.

Method 1

20.0cm3 solution ->7.1 g

1000cm3 solution -> 1000 x 71 = 3550 g **dm^{-3}**

20

Molar mass Na_2SO_4 = 142 g

Moles **dm^{-3}** = Molarity = Mass 3550 = 2.5 M/ moles**dm^{-3}**

Molar mass 142

Method 2

Molar mass Na_2SO_4 = 142 g

Moles = Mass => 7.1 = **0.05 / 5.0 x 10 $^{-2}$** moles

Molar mass 142

Method 2(a)

Moles in decimeters = **Molarity** = Moles x 1000cm3/1dm3

Volume of solution

=> 5.0 x 10 -2 moles x 1000cm3

20cm3

= **2.5 M/2.5 molesdm^{-3}**

Method 2(b)

20 cm3 solution contain 5.0 x 10 $^{-2}$ moles

1000cm3 solution = Molarity contain 1000 x 5.0 x 10 $^{-2}$ moles

20 cm3

= **2.5 M/2.5 molesdm^{-3}**

(iv) The density of sulphuric(VI) is 1.84gcm^{-3} Calculate the molarity of the acid.

Method 1

1.0cm3 solution ->1.84 g

1000cm3 solution -> 1000 x 1.84 = 1840 g **dm^{-3}**

1

Molar mass H_2SO_4 = 98 g

Moles **dm^{-3}** = Molarity = Mass = 1840

Molar mass 98

= **18.7755 M**/ moles**dm^{-3}**

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Method 2

Molar mass H_2SO_4 = 98 g

Moles = $\frac{\text{Mass}}{\text{Molar mass}}$ => $\frac{1.84}{98}$ = **$0.0188 / 1.88 \times 10^{-2}$** moles

Method 2(**a**)

Moles in decimeters = **Molarity** = Moles x $\frac{1000cm3/1dm3}{\text{Volume of solution}}$

=> 1.88×10^{-2} moles x $\frac{1000cm3}{1.0cm3}$

= **18.8M/18.8 molesdm^{-3}**

Method 2(**b**)

20 cm3 solution contain 1.88×10^{-2} moles

1000cm3 solution = Molarity contain $\frac{1000}{1.0\ cm3} \times 1.88 \times 10^{-2}$ moles

= **18.8M/18.8 molesdm^{-3}**

2. Calculate the mass of :

(i) 25 cm3 of 0.2M sodium hydroxide solution(Na =23.0.O =16.0, H=1.0)

Molar mass NaOH = **40g**

Moles in 25 cm3 = $\frac{\text{Molarity x volume}}{1000}$ => $\frac{0.2 \times 25}{1000}$ = **$0.005/5.0 \times 10^{-3}$**moles

Mass of NaOH =Moles x molar mass = 5.0×10^{-3} x 40 = **0.2 g**

(ii) 20 cm3 of 0.625 M sulphuric(VI)acid (S =32.0.O =16.0, H=1.0)

Molar mass H_2SO_4 = **98g**

Moles in 20 cm3 = $\frac{\text{Molarity x volume}}{1000}$ => $\frac{0.625 \times 20}{1000}$ = **$0.0125/1.25.0 \times 10^{-3}$**moles

Mass of H_2SO_4 =Moles x molar mass => 5.0×10^{-3} x 40 = **0.2 g**

(iii) 1.0 cm3 of 2.5 M Nitric(V)acid (N =14.0.O =16.0, H=1.0)

Molar mass HNO_3 = **63 g**

Moles in 1 cm3 = $\frac{\text{Molarity x volume}}{1000}$ => $\frac{2.5 \times 1}{1000}$ = **$0.0025 / 2.5. \times 10^{-3}$**moles

Mass of HNO_3 =Moles x molar mass => 2.5×10^{-3} x 40 = **0.1 g**

3. Calculate the volume required to dissolve :

(a)(i) 0.25moles of sodium hydroxide solution to form a 0.8M solution

Volume (in cm3) = $\frac{\text{moles x 1000}}{\text{Molarity}}$ => $\frac{0.25 \times 1000}{0.8}$ = **312.5cm3**

(ii) 100cm3 was added to the sodium hydroxide solution above. Calculate the concentration of the solution.

$C_1 \times V_1 = C_2 \times V_2$ where:

C_1 = molarity/concentration before diluting/adding water

C_2 = molarity/concentration after diluting/adding water

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V_1 = volume before diluting/adding water
V_2 = volume after diluting/adding water
=> 0.8M x 312.5cm3 = C_2 x (312.5 + 100)
C_2 = 0.8M x 312.5cm3 = **0.6061M**
412.5

(b)(ii) 0.01M solution containing 0.01moles of sodium hydroxide solution .
Volume (in cm3) = moles x 1000 => 0.01 x 1000 = **1000 cm3**
Molarity 0.01

(ii) Determine the quantity of water which must be added to the sodium hydroxide solution above to form a 0.008M solution.
$C_1 \times V_1 = C_2 \times V_2$ where:
C_1 = molarity/concentration before diluting/adding water
C_2 = molarity/concentration after diluting/adding water
V_1 = volume before diluting/adding water
V_2 = volume after diluting/adding water
=> 0.01M x 1000 cm3 = 0.008 x V_2
V_2 = 0.01M x 1000cm3 = **1250cm3**
0.008
Volume added = 1250 - 1000 = **250cm3**

(c)Volumetric analysis/Titration
Volumetric analysis/Titration is the process of determining unknown concentration of one reactant from a known concentration and volume of another.
Reactions take place in simple mole ratio of reactants and products.
Knowing the concentration/ volume of one reactant, the other can be determined from the relationship:

$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2}$ where:

M_1 = Molarity of 1st reactant
M_2 = Molarity of 2nd reactant
V_1 = Volume of 1st reactant
V_1 = Volume of 2nd reactant
n_1 = number of moles of 1st reactant from stoichiometric equation
n_2 = number of moles of 2nd reactant from stoichiometric equation

Examples
1.Calculate the molarity of MCO_3 if 5.0cm3 of MCO_3 react with 25.0cm3 of 0.5M hydrochloric acid.(C=12.0 ,O =16.0)
Stoichiometric equation: $MCO_3(s) + 2HCl(aq) \rightarrow MCl_2(aq) + CO_2(g) + H_2O(l)$
Method 1
$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2}$ -> $M_1 \times 5.0cm3 = 0.5M \times 25.0cm3$
1 2

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$\Rightarrow M_1 = \frac{0.5 \times 25.0 \times 1}{5.0 \times 2} =$ **1.25M / 1.25 moledm^{-3}**

Method 2

Moles of HCl used = molarity x volume / 1000

$\Rightarrow \frac{0.5 \times 25.0}{1000} =$ **0.0125 /1.25 x 10^{-2}moles**

Mole ratio MCO_3 : HCl = 1:2

Moles $MCO_3 = \frac{0.0125 /1.25 \times 10^{-2}\text{moles}}{2} =$ **0.00625 / 6.25 x 10^{-3} moles**

Molarity MCO_3 = moles x 1000 / Volume $\Rightarrow \frac{0.00625 / 6.25 \times 10^{-3} \times 1000}{5}$

= **1.25M / 1.25 moledm^{-3}**

2. 2.0cm3 of 0.5M hydrochloric acid react with 0.1M of M_2CO_3. Calculate the volume of 0.1M M_2CO_3 used.

Stoichiometric equation: M_2CO_3 (aq) + 2HCl(aq) -> 2MCl (aq) + CO_2(g) + H_2O(l)

Method 1

$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2}$ -> $\frac{0.5 \times 2.0\text{cm3}}{2} = \frac{0.1\text{M} \times V_2\text{ cm3}}{1}$

$\Rightarrow V_2 = \frac{0.5 \times 2.0 \times 1}{0.1 \times 2} =$ **1.25M / 1.25 moledm^{-3}**

Method 2

Moles of HCl used = molarity x volume / 1000

$\Rightarrow \frac{0.5 \times 2.0}{1000} =$ **0.0125 /1.25 x 10^{-2}moles**

Mole ratio M_2CO_3 : HCl = 1:2

Moles $M_2CO_3 = \frac{0.0125 /1.25 \times 10^{-2}\text{moles}}{2} =$ **0.00625 / 6.25 x 10^{-3} moles**

Molarity M_2CO_3 = moles x 1000 / Volume $\Rightarrow \frac{0.00625 / 6.25 \times 10^{-3} \times 1000}{5}$

= **1.25M / 1.25 moledm^{-3}**

3. 5.0cm3 of 0.1M sodium iodide react with 0.1M of Lead(II)nitrate(V). Calculate(i) the volume of Lead(II)nitrate(V) used.

(ii)the mass of Lead(II)Iodide formed

(Pb=207.0, I =127.0)

Stoichiometric equation: 2NaI(aq) + $Pb(NO_3)_2$(aq) -> $2NaNO_3$(aq) + PbI_2(**s**)

(i)Volume of Lead(II)nitrate(V) used

Method 1

$\frac{M_1V_1}{} = \frac{M_2V_2}{}$ -> $\frac{5 \times 0.1\text{cm3}}{} = \frac{0.1\text{M} \times V_2\text{ cm3}}{}$

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$\frac{n_1}{n_2} = \frac{2}{1}$

$\Rightarrow V_2 = \frac{0.1 \times 5.0 \times 1}{0.1 \times 2} =$ **1.25M / 1.25 moledm^{-3}**

Method 2

Moles of HCl used = molarity x volume

$\Rightarrow \frac{0.1 \times 5.0}{1000} =$ **0.0125 /1.25 x 10^{-2}moles**

Mole ratio M_2CO_3 : HCl = 1:2

Moles M_2CO_3 = $\frac{0.0125 /1.25 \times 10^{-2}\text{moles}}{2} =$ **0.00625 / 6.25 x 10^{-3} moles**

Molarity M_2CO_3 = $\frac{\text{moles} \times 1000}{\text{Volume}} \Rightarrow \frac{0.00625 / 6.25 \times 10^{-3} \times 1000}{5}$

= **1.25M / 1.25 moledm^{-3}**

4. 0.388g of a monobasic organic acid B required 46.5 cm3 of 0.095M sodium hydroxide for complete neutralization. Name and draw the structural formula of B

Moles of NaOH used = molarity x volume

$\Rightarrow \frac{0.095 \times 46.5}{1000} =$ **0.0044175 /4.4175 x 10^{-3}moles**

Mole ratio B : NaOH = 1:1

Moles B = **0.0044175 /4.4175 x 10^{-3}moles**

Molar mass B = $\frac{\text{mass}}{\text{moles}} \Rightarrow \frac{0.388}{0.0044175 /4.4175 \times 10^{-3}\text{moles}}$

= **87.8324 gmole^{-1}**

X-COOH = 87.8324 where X is an alkyl group

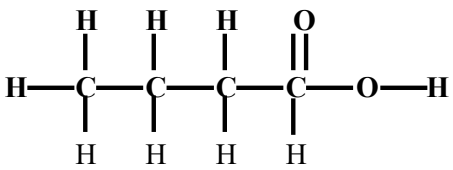
X =87.8324- 42 = 42.8324=43

By elimination: CH_3 = 15 CH_3CH_2 = 29 **CH_3CH_2 CH_2 = 43**

Molecula formula : CH_3CH_2 CH_2COOH

Molecule name : Butan-1-oic acid

Molecular structure



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5. 10.5 g of an impure sample containing ammonium sulphate (VI) fertilizer was warmed with 250cm3 of o.8M sodium hydroxide solution.The excess of the alkali was neutralized by 85cm3 of 0.5M hydrochloric acid. Calculate the % of impurities in the ammonium sulphate (VI)fertilizer. (N=14.0,S=32.0,O=16.0, H=1.0)

Equation for neutralization

$NaOH(aq) + HCl(aq) \rightarrow NaOH(aq) + H_2O(l)$

Mole ratio NaOH(aq):HCl(aq)= 1:1

Moles of HCl = Molarity x volume $\Rightarrow \frac{0.5 \times 85}{1000} =$ **0.0425 moles**

Excess moles of NaOH(aq)= **0.0425 moles**

Equation for reaction with ammonium salt

$2NaOH(aq) + (NH_4)_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2NH_3(g) + 2H_2O(l)$

Mole ratio NaOH(aq): $(NH_4)_2SO_4(aq)$= 2:1

Total moles of NaOH = Molarity x volume $\Rightarrow \frac{0.8 \times 250}{1000} =$ **0.2 moles**

Moles of NaOH that reacted with $(NH_4)_2SO_4$ = 0.2 - 0.0425 = **0.1575moles**

Moles $(NH_4)_2SO_4$ = ½ x 0.1575moles = **0. 07875moles**

Molar mass $(NH_4)_2SO_4$= **132 gmole^{-1}**

Mass of in impure sample = moles x molar mass =>0. 07875 x 132 = **10.395 g**

Mass of impurities = 10.5 -10.395 = **0.105 g**

% impurities $= \frac{0.105 \times 100}{10.5} =$ **1.0 %**

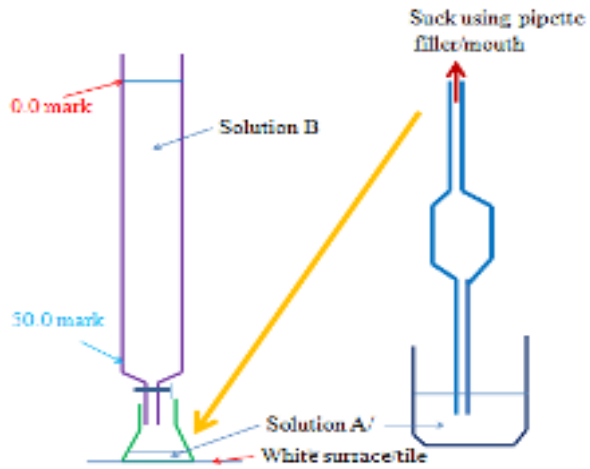
Practically volumetric analysis involves **titration**.

Titration generally involves filling a burette with known/unknown concentration of a solution then adding the solution to unknown/known concentration of another solution in a conical flask until there is complete reaction. If the solutions used are both colourless, an **indicator** is added to the conical flask. When the reaction is over, a **slight/little** excess of burette contents **change** the colour of the indicator. This is called the **end point**.

Set up of titration apparatus

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Set up of Titration apparatus

The titration process involve involves determination of **titre.** The titre is the volume of burette contents/reading **before** and **after** the end point. Burette contents/reading **before** titration is usually called the **Initial** burette reading. Burette contents/reading **after** titration is usually called the **Final** burette reading. The titre value is thus a sum of the **Final** less **Initial** burette readings.

To reduce errors, titration process should be repeated at least once more.

The results of titration are recorded in a **titration table** as below

Sample titration table

Titration number	1	2	3
Final burette reading (cm3)	20.0	20.0	20.0
Initial burette reading (cm3)	0.0	0.0	0.0
Volume of solution used(cm3)	20.0	20.0	20.0

As **evidence** of a titration **actually** done examining body requires the candidate to record their burette readings before and after the titration.

For KCSE candidates burette readings **must** be recorded **in** a titration table in the **format provided** by the Kenya National Examination Council.

As **evidence** of all titration **actually** done Kenya National Examination Council require the candidate to record their burette readings before and after the titration to complete the titration table **in the format provided**.

Calculate the average volume of solution used

$$\frac{24.0 + 24.0 + 24.0}{3} = \textbf{24.0 cm3}$$

As evidence of understanding the degree of accuracy of burettes , all readings must be recorded to **a** decimal point.

As evidence of accuracy in carrying the out the titration , candidates value should be **within 0.2** of the **school value** .

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The school value is the **teachers** readings presented to the examining body/council based on the concentrations of the solutions s/he presented to her/his candidates. Bonus mark is awarded for averaged reading **within 0.1** school value as Final answer.

Calculations involved after the titration require candidates **thorough** practical and theoretical **practice mastery** on the:

(i)relationship among the mole, molar mass, mole ratios, concentration, molarity.

(ii) mathematical application of 1st principles.

Very useful information which candidates forget appears usually in the beginning of the question paper as:

"**You are provided with…**"

All calculation must be to the **4th decimal point** unless they divide fully to a lesser decimal point.

Candidates are expected to use a non programmable scientific calculator.

(a) Sample Titration Practice 1 **(Simple Titration)**

You are provided with:

0.1M sodium hydroxide solution A

Hydrochloric acid solution B

You are required to determine the concentration of solution B in moles per litre.

Procedure

Fill the burette with solution B. Pipette 25.0cm3 of solution A into a conical flask. Titrate solution A with solution B using phenolphthalein indicator to complete the titration table 1

Sample results Titration table 1

Titration number	1	2	3
Final burette reading (cm3)	**20.0**	**20.0**	**20.0**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of solution B used(cm3)	**20.0**	**20.0**	**20.0**

Sample worked questions

1. Calculate the average volume of solution B used

Average titre = $\frac{\text{Titre 1 + Titre 2 +Titre 3}}{3}$ => $\frac{(20.0 + 20.0 + 20.0)}{3}$ = **20.0cm3**

2. How many moles of:

(i)solution A were present in 25cm3 solution.

Moles of solution A = $\frac{\text{Molarity x volume}}{1000}$ = $\frac{0.1 \times 25}{1000}$ = **2.5×10^{-3}** moles

(ii)solution B were present in the average volume.

Chemical equation: $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

Mole ratio 1:1 => Moles of A = Moles of B = **2.5×10^{-3}** moles

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(iii) solution B in moles per litre.

Moles of B per litre = $\frac{\text{moles} \times 1000}{\text{Volume}} = \frac{2.5 \times 10^{-3} \times 1000}{20} =$ **0.1M**

(**b**)Sample Titration Practice 2 **(Redox Titration)**

You are provided with:

Acidified Potassium manganate(VII) solution A

0.1M of an iron (II)salt solution B

8.5g of ammonium iron(II)sulphate(VI) crystals$(NH_4)_2\ SO_4FeSO_4.xH_2O$

solid C

You are required to

(i)standardize acidified potassium manganate(VII)

(ii)determine the value of x in the formula $(NH_4)_2\ SO_4FeSO_4.xH_2O$.

Procedure 1

Fill the burette with solution A. Pipette 25.0cm3 of solution B into a conical flask.

Titrate solution A with solution B until a pink colour just appears.

Record your results to complete table 1.

Table 1:**Sample results**

Titration number	1	2	3
Final burette reading (cm3)	**20.0**	**20.0**	**20.0**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of solution A used(cm3)	**20.0**	**20.0**	**20.0**

Sample worked questions

1. Calculate the average volume of solution A used

Average titre = $\frac{\text{Titre 1 + Titre 2 +Titre 3}}{3} \Rightarrow \frac{(20.0 + 20.0 + 20.0)}{3} =$ **20.0cm3**

2. How many moles of:

(i)solution B were present in 25cm3 solution.

Moles of solution A = $\frac{\text{Molarity} \times \text{volume}}{1000} = \frac{0.1 \times 25}{1000} =$ **2.5×10^{-3}** moles

(ii)solution A were present in the average volume. Assume one mole of B react with five moles of B

Mole ratio A : B = 1:5

$\Rightarrow$ Moles of A = Moles of B = $\frac{2.5 \times 10^{-3}\ \text{moles}}{5} =$ **5.0×10^{-4} moles**

(iii) solution B in moles per litre.

Moles of B per litre = $\frac{\text{moles} \times 1000}{\text{Volume}} = \frac{2.5 \times 10^{-3} \times 1000}{20}$

= **0.025** M /moles per litre /moles l^{-1}

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Procedure 2

Place all the solid C into the 250cm3 volumetric flask carefully. Add about 200cm3 of distilled water. Shake to dissolve. Make up to the 250cm3 of solution by adding more distilled water. Label this solution C. Pipette 25cm3 of solution C into a conical flask, Titrate solution C with solution A until a permanent pink colour just appears. Complete table 2.

Table 2:**Sample results**

Titration number	1	2	3
Final burette reading (cm3)	**20.0**	**20.0**	**20.0**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of solution A used(cm3)	**20.0**	**20.0**	**20.0**

Sample worked questions

1. Calculate the average volume of solution A used

Average titre = $\frac{\text{Titre 1 + Titre 2 +Titre 3}}{3}$ => $\frac{(20.0 + 20.0 + 20.0)}{3}$ = **20.0cm3**

2. How many moles of:

(i)solution A were present inin the average titre.

Moles of solution A = $\frac{\text{Molarity} \times \text{volume}}{1000}$ = $\frac{0.025 \times 20}{1000}$ = **5.0×10^{-4}** moles

(ii)solution C in 25cm3 solution given the equation for the reaction:

$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)$

Mole ratio $MnO_4^-(aq)$: $5Fe^{2+}(aq)$ = 1:5 =>

Moles of $5Fe^{2+}(aq)$ = $\frac{\text{Moles of } MnO_4^-(aq)}{5}$ = $\frac{5.0 \times 10^{-4} \text{ moles}}{5}$ = **1.0×10^{-4} moles**

(iii) solution B in 250cm3.

Moles of B per litre = $\frac{\text{moles} \times 250}{\text{Volume}}$ = $\frac{1.0 \times 10^{-4} \times 250}{25}$ = **1.0×10^{-3} moles**

3. Calculate the molar mass of solid C and hence the value of x in the chemical formula $(NH_4)_2SO_4FeSO_4.xH_2O$.

(N=14.0, S=32.0, Fe=56.0, H=1.0 O=16.0)

Molar mass = $\frac{\text{mass perlitre}}{\text{Moles per litre}}$ = $\frac{8.5}{1.0 \times 10^{-3} \text{ moles}}$ = **8500 g**

$(NH_4)_2SO_4FeSO_4.xH_2O$ = 8500

284 + 18x =8500

8500 - 284 = $\frac{8216}{18}$ = $\frac{18x}{18}$ = **454.4444**

x = **454** (whole number)

(c)Sample Titration Practice 3 **(Back titration)**

You are provided with:

(i)an impure calcium carbonate labeled M

(ii)Hydrochloric acid labeled solution N

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(iii)solution L containing 20g per litre sodium hydroxide.

You are required to determine the concentration of N in moles per litre and the % of calcium carbonate in mixture M.

Procedure 1

Pipette 25.0cm3 of solution L into a conical flask. Add 2-3 drops of phenolphthalein indicator. Titrate with dilute hydrochloric acid solution N and record your results in table 1(4mark)

Sample Table 1

	1	2	3
Final burette reading (cm3)	**6.5**	**6.5**	**6.5**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of N used (cm3)	**6.5**	**6.5**	**6.5**

Sample questions

(a) Calculate the average volume of solution N used

$$\frac{6.5 + 6.5 + 6.5}{3} = \mathbf{6.5}\ \text{cm3}$$

(b) How many moles of sodium hydroxide are contained in 25cm3of solution L

Molar mass NaOH =**40**g

$$\text{Molarity of L} = \frac{\text{mass per litre}}{\text{Molar mass NaOH}} \Rightarrow \frac{20}{40} = \mathbf{0.5}\text{M}$$

$$\text{Moles NaOH in 25cm3} = \frac{\text{molarity} \times \text{volume}}{1000} \Rightarrow \frac{0.5\text{M} \times 25\text{cm3}}{1000} = \mathbf{0.0125}\ \text{moles}$$

(c)Calculate:

(i)the number of moles of hydrochloric acidthat react with sodium hydroxide in (b)above.

Mole ratio NaOH : HCl from stoichiometric equation= **1:1**

Moles HCl =Moles NaOH => **0.0125** moles

(ii)the molarity of hydrochloric acid solution N.

$$\text{Molarity} = \frac{\text{moles} \times 1000}{6.5} \Rightarrow \frac{0.0125\ \text{moles} \times 1000}{6.5} = \mathbf{1.9231}\text{M/moledm}^{-3}$$

Procedure 2

Place the 4.0 g of M provided into a conical flask and add 25.0cm3 of the dilute hydrochloric acid to it using a clean pipette. Swirl the contents of the flask vigorously until effervescence stop.Using a 100ml measuring cylinder add 175cm3 distilled waterto make up the solution up to 200cm3.Label this solution K.Using a clean pipettetransfer 25.0cm3 of the solution into a clean conical flask and titrate with solution L from the burette using 2-3 drops of methyl orange indicator.Record your observations in table 2.

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Sample Table 2

	1	2	3
Final burette reading (cm3)	**24.5**	**24.5**	**24.5**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of N used (cm3)	**24.5**	**24.5**	**24.5**

Sample calculations

(a)Calculate the average volume of solution L used(1mk)

$\frac{24.5 + 24.5 + 24.5}{3}$ = **24.5**cm3

(b)How many moles of sodium hydroxide are present in the average volume of solution L used?

Moles = molarity x average burette volume => $0.5 \times \frac{24.5}{1000}$

$= \frac{1000}{}$ **0.01225 /1.225 x 10^{-2}** moles

(c) How many moles of hydrochloric acid are present in the original 200cm3 of solution K?

Mole ratio NaOH: HCl = 1:1 => moles of HCl = **0.01225 /1.225 x 10^{-2}** moles

Moles in 200cm3 = $200\text{cm}3 \times \frac{0.01225 /1.225 \times 10^{-2}\text{moles}}{25\text{cm}3\text{(volume pipetted)}}$

=**0.49 /4.9 x 10^{-1}**moles

(d)How many moles of hydrochloric acid were contained in original 25 cm3 solution N used

Original moles = Original molarity x pipetted volume =>

$1.9231\text{M/moledm}^{-3} \times \frac{25}{1000\text{cm}3}$ = **0.04807/4.807 x 10^{-2}** moles

(e)How many moles of hydrochloric acid were used to react with calcium carbonate present?

Moles that reacted = original moles –moles in average titre =>

$0.04807/4.807 \times 10^{-2}\text{moles} - 0.01225 /1.225 \times 10^{-2}\text{moles}$

= **0.03582/3.582 x 10^{-2}** moles

(f)Write the equation for the reaction between calcium carbonate and hydrochloric acid.

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

(g)Calculate the number of moles of calcium carbonate that reacted with hydrochloric acid.

From the equation $CaCO_3(s):2HCl(aq) = 1:2$

=> Moles $CaCO_3(s) = {}^1/_2$moles HCl

$= {}^1/_2 \times 0.03582/3.582 \times 10^{-2}$ moles

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= **0.01791 /1.791 x 10^{-2}**moles

(h)Calculate the mass of calcium carbonate in 4.0g of mixture M (Ca=40.0,O = 16.0,C=12.0)

Molar mass $CaCO_3$ = 100g

Mass $CaCO_3$ = moles x molar mass => 0.01791 /1.791 x 10^{-2}moles x 100g

= **1.791**g

(i)Determine the % of calcium carbonate present in the mixture

% $CaCO_3$ = mass of pure x 100% => 1.791g x 100% = **44.775**%

Mass of impure 4.0

(d)Sample titration practice 4 (**Multiple titration**)

You are provided with:

(i)sodium L containing 5.0g per litre of a dibasic organic acid $H_2X.2H_2O$.

(ii)solution M which is acidified potassium manganate(VII)

(iii)solution N a mixture of sodium ethanedioate and ethanedioic acid

(iv)0.1M sodium hydroxide solution P

(v)1.0M sulphuric(VI)

You are required to:

(i)standardize solution M using solution L

(ii)use standardized solution M and solution P to determine the % of sodium ethanedioate in the mixture.

Procedure 1

Fill the burette with solution M. Pipette 25.0cm3 of solution L into a conical flask. Heat this solution to about 70°C(**but not to boil**).Titrate the hot solution L with solution M until a permanent pink colour just appears .Shake thoroughly during the titration. Repeat this procedure to complete table 1.

Sample Table 1

	1	2	3
Final burette reading (cm3)	**24.0**	**24.0**	**24.0**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of N used (cm3)	**24.0**	**24.0**	**24.0**

Sample calculations

(a)Calculate the average volume of solution L used (1mk)

24.0 + 24.0 + 24.0 = **24.0**cm3

3

(b)Given that the concentration of the dibasic acid is 0.05molesdm^{-3}.determine the value of x in the formula $H_2X.2H_2O$ (H=1.0,O=16.0)

Molar mass $H_2X.2H_2O$= mass per litre => 5.0g/litre = **100**g

Moles/litre 0.05molesdm^{-3}

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$H_2X.2H_2O$ =100

X = 100 – ((2 x1) + 2 x (2 x1) + (2 x 16) => 100 – 34 = **66**

(c) Calculate the number of moles of the dibasic acid $H_2X.2H_2O$.

Moles = molarity x pipette volume => 0.5 x 25 = **0.0125/1.25 x10 $^{-2}$** moles
1000 1000

(d)Given the mole ratio manganate(VII)(MnO_4^-): acid H_2X is 2:5, calculate the number of moles of manganate(VII) (MnO_4^-) in the average titre.

Moles H_2X = $^2/_5$ moles of MnO_4^-

=> $^2/_5$ x 0.0125/1.25 x10 $^{-2}$ moles

= **0.005/5.0 x 10 $^{-3}$**moles

(e)Calculate the concentration of the manganate(VII)(MnO_4^-) in moles per litre.

Moles per litre/molarity = moles x 1000
average burette volume

=>0.005/5.0 x 10 $^{-3}$moles x 1000 = **0.2083** molesl^{-1}/M
24.0

Procedure 2

With solution M still in the burette ,pipette 25.0cm3 of solution N into a conical flask. Heat the conical flask containing solution N to about 70°C.Titrate while hot with solution M.Repeat the experiment to complete table 2.

Sample Table 2

	1	2	3
Final burette reading (cm3)	**12.5**	**12.5**	**12.5**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of N used (cm3)	**12.5**	**12.5**	**12.5**

Sample calculations

(a)Calculate the average volume of solution L used (1mk)

12.5 + 12.5 + 12.5 =**12.5**cm3
3

(b)Calculations:

(i)How many moles of manganate(VII)ions are contained in the average volume of solution M used?

Moles = molarity of solution M x average burette volume
1000

=> 0.2083 molesl^{-1}/ M x 12.5 = **0.0026 / 2.5 x 10^{-3}** moles
1000

(ii)The reaction between manganate(VII)ions and ethanedioate ions that reacted with is as in the equation:

$2MnO_4^-$ (aq) + $5C_2O_4^{2-}$ (aq) + $16H^+$ (aq) -> $2Mn^{2+}$(aq) + $10CO_2$(g) + $8H_2O$(l)

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Calculate the number of moles of ethanedioate ions that reacted with manganate (VII) ions in the average volume of solution M.

From the stoichiometric equation,mole ratio MnO_4^- (aq): $C_2O_4^{2-}$ (aq) = 2:5

=> moles $C_2O_4^{2-}$ = $^5/_2$ moles MnO_4^- => $^5/_2$ x 0.0026 / 2.5 x 10^{-3} moles

= **0.0065 /6.5 x10^{-3}** moles

(iii)Calculate the number of moles of ethanedioate ions contained in 250cm3 solution N.

25cm3 pipette volume -> 0.0065 /6.5 x10^{-3} moles

250cm3 ->

$$0.0065/6.5 \times 10^{-3} \text{ moles} \times \frac{250}{25} = \mathbf{0.065 / 6.5 \times 10^{-2}} \text{ moles}$$

Procedure 3

Remove solution M from the burette and rinse it with distilled water. Fill the burette with sodium hydroxide solution P. Pipette 25cm3 of solution N into a conical flask and add 2-3 drops of phenolphthalein indicator. Titrate this solution N with solution P from the burette. Repeat the procedure to complete table 3.

Sample Table 2

	1	2	3
Final burette reading (cm3)	**24.9**	**24.9**	**24.9**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of N used (cm3)	**24.9**	**24.9**	**24.9**

Sample calculations

(a)Calculate the average volume of solution L used (1mk)

$$\frac{24.9 + 24.9 + 24.9}{3} = \mathbf{24.9\ cm3}$$

(b)Calculations:

(i)How many moles of sodium hydroxide solution P were contained in the average volume?

Moles = molarity of solution P x average burette volume / 1000

=> 0.1 $moles l^{-1}$ x $\frac{24.9}{1000}$ = **0.00249 / 2.49 x 10^{-3}** moles

(ii)Given that NaOH solution P reacted with the ethanedioate ions from the acid only and the equation for the reaction is:

$2NaOH(aq) + H_2C_2O_4(aq) \rightarrow Na_2C_2O_4(g) + 2H_2O(l)$

Calculate the number of moles of ethanedioic acid that were used in the reaction

From the stoichiometric equation,mole ratio NaOH(aq): $H_2C_2O_4$ (aq) = 2:1

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=> moles $H_2C_2O_4$ = $^1/_2$ moles NaOH => $^1/_2$ x 0.00249 / 2.49 x 10^{-3} moles

= **0.001245/1.245 x10^{-3}** moles.

(iii)How many moles of ethanedioic acid were contained in 250cm3 of solution N?

25cm3 pipette volume -> 0.001245/1.245 x10^{-3} moles

250cm3 ->

0.001245/1.245 x10^{-3} moles x $\frac{250}{25}$ = **0.01245/1.245 x10^{-2}** moles

(iii)Determine the % by mass of sodium ethanedioate in the micture

(H= 1.0,O=16.0,C=12.0 and total mass of mixture =2.0 g in 250cm3 solution)

Molar mass $H_2C_2O_4$ = **90.0**g

Mass of $H_2C_2O_4$ in 250cm3 = moles in 250cm3 x molar mass $H_2C_2O_4$

=>0.01245/1.245 x10^{-2} moles x 90.0

= **1.1205**g

% by mass of sodium ethanedioate

$$=\frac{(\text{Mass of mixture} - \text{mass of } H_2C_2O_4) \times 100\%}{\text{Mass of mixture}}$$

$$\Rightarrow \frac{2.0 - 1.1205\ g}{2.0} = \mathbf{43.975}\%$$

Note

(i) L is 0.05M Oxalic acid

(ii) M is 0.01M KMnO4

(iii) N is 0.03M oxalic acid(without sodium oxalate)

Practice example 5.(Determining equation for a reaction)

You are provided with

-0.1M hydrochloric acid solution A

-0.5M sodium hydroxide solution B

You are to determine the equation for thereaction between solution A and B

Procedure

Fill the burette with solution A.Using a pipette and pipette filler transfer 25.0cm3 of solution B into a conical flask.Add 2-3 drops of phenolphthalein indicator.Run solution A into solution B until a permanent pink colour just appears.Record your results in Table 1.Repeat the experiment to obtain three concordant results to complete Table 1

Table 1(Sample results)

Titration	1	2	3
Final volume(cm3)	**12.5**	**25.0**	**37.5**
Initial volume(cm3)	**0.0**	**12.5**	**25.0**
Volume of solution A used(cm3)	**12.5**	**12.5**	**12.5**

Sample questions

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Calculate the average volume of solution A used.

$$\frac{12.5+12.5+12.5}{3} = \mathbf{12.5cm3}$$

Theoretical Practice examples

1. 1.0g of dibasic acid $HOOC(CH_2)_xCOOH$ was dissolved in 250cm3 solution. 25.0 cm3 of this solution reacted with 30.0cm3 of 0.06M sodium hydroxide solution. Calculate the value of x in $HOOC(CH_2)_xCOOH$.
(C=12.0,H=1.0,O=16.)

Chemical equation

$2NaOH(aq) + H_2X(aq) \rightarrow Na_2X(aq) + 2H_2O(aq)$

Mole ratio $NaOH(aq) : H_2X(aq) = 2:1$

Method 1

$Ma\ Va = na \Rightarrow \frac{Ma \times 25.0}{2} = \frac{0.06 \times 30.0 \times 1}{25.0 \times 2}$

$Mb\ Vb = nb$

Molarity of acid = **0.036M/Mole l^{-1}**

Mass of acid per lite = $\frac{1.0 \times 1000}{250}$ = **4.0 g/l**

0.036M/ Mole l^{-1} -> 4.0 g /l

1 mole= molar mass of $HOOC(CH_2)_xCOOH = \frac{4.0 \times 1}{0.036}$ = **111.1111** g

Molar mass $(CH_2)_x$ = 111.1111 – (HOOCCOOH = 90.0) **= 21.1111**

$(CH_2)_x = 14x = 21.1111$ = **1.5 = 1 (whole number)**

$\frac{}{14}$

Method 2

Moles of sodium hydroxide = $\frac{\text{Molarity} \times \text{volume}}{1000} = \frac{0.06 \times 30}{1000}$ = **1.8×10^{-3}moles**

Moles of Hydrochloric acid = $\frac{1}{2} \times 1.8 \times 10^{-3}$moles = 9.0×10^{-4}moles

Molarity of Hydrochloric acid = $\frac{\text{moles} \times 1000}{\text{Volume}} = \frac{9.0 \times 10^{-4}\text{moles} \times 1000}{25}$

Molarity of acid = **0.036M/Mole l^{-1}**

Mass of acid per lite = $\frac{1.0 \times 1000}{250}$ = **4.0 g/l**

0.036M/ Mole l^{-1} -> 4.0 g /l

1 mole= molar mass of $HOOC(CH_2)_xCOOH = \frac{4.0 \times 1}{0.036}$ = **111.1111** g

Molar mass $(CH_2)_x$ = 111.1111 – (HOOCCOOH = 90.0) **= 21.1111**

$(CH_2)_x = 14x = 21.1111$ = **1.5 = 1 (whole number)**

$\frac{}{14}$

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2. 20.0cm3 of 0.05 M acidified potassium manganate(VII)solution oxidized 25.0cm3 of $Fe^{2+}(aq)$ ions in 40.0g/l of impure Iron (II)sulphate(VI) to $Fe^{3+}(aq)$ ions. Calculate the percentage impurities in the Iron (II)sulphate(VI).
$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow 5Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(aq)$
Fe=56.0,S= 32.0, O=16.0).

Moles of MnO_4^- (aq) = Molarity x volume = $\frac{0.05 \times 20.0}{1000}$ = **0.001 Moles**
1000

Mole ratio MnO_4^- (aq): $5Fe^{2+}(aq)$= 1:5

Moles $5Fe^{2+}(aq)$ = 5 x0.001 = **0.005 Moles**

Moles of $5Fe^{2+}(aq)$ per litre/molarity = $\frac{\text{Moles} \times 1000}{\text{Volume}}$ = $\frac{0005 \times 1000}{25.0}$

= 0.2 M/ Moles/litre

Molar mass =$FeSO_4$=**152 g**

Mass of in the mixture = Moles x molar mass => 0.2 x 152 = **30.4** g

Mass of impurity = 40.0 – 30.4 =**9.6 g**

% impurity = $\frac{9.6 \text{ g}}{40.0} \times 100$ = **24.0 % impurity**

3.9.7 g of a mixture of Potassium hydroxide and Potassium chloride was dissolved to make one litre solution.20.0cm3 of this solution required 25.0cm3 of 0.12M hydrochloric acid for completed neutralization. Calculate the percentage by mass of Potassium chloride.(K=39.0,Cl= 35.5)

Chemical equation

$KOH(aq) + HCl(aq) \rightarrow KCl(aq) + H_2O(l)$

Moles of HCl = Molarity x volume => $\frac{0.12 \times 25.0}{1000}$ = **$0.003/3.0 \times 10^{-3}$** moles
1000

Mole ratio KOH(aq) : HCl(aq) -= 1:1

Moles KOH =**$0.003/3.0 \times 10^{-3}$** moles

Method 1

Molar mass KOH =**56.0**g

Mass KOH in 25cm3 =$0.003/3.0 \times 10^{-3}$ moles x56.0 = 0.168g

Mass KOH in 1000cm3/1 litre = $0.168 \times \frac{1000}{20}$= **8.4 g/l**

Mass of KCl = 9.7g - 8.4g = **1.3** g

% of KCl = $\frac{1.3 \times 100}{9.7}$ = **13.4021%**

Method 2

Moles KOH in 1000cm3 /1 litre = Moles in 20cm3 x $\frac{1000}{20}$ => $0.003 \times \frac{1000}{20}$

=0.15M/Moles /litre

Molar mass KOH =**56.0**g

Mass KOH in 1000/1 litre = 0.15M/Moles /litre x 56.0 = **8.4g/l**

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Mass of KCl = 9.7g - 8.4g = **1.3** g
% of KCl = 1.3 x 100 = **13.4021%**
9.7

4.**A certain carbonate, GCO3, reacts with dilute hydrochloric acid according to the equation given below:**

$GCO_{3(s)} + 2HCl_{(aq)} \rightarrow GCl_{2\ (aq)} + CO_2\ (g) + H_2O_{(l)}$

If 1 g of the carbonate reacts completely with 20 cm3 of 1 M hydrochloric acid ,calculate the relative atomic mass of G (C = 12.0 = 16.0)

Moles of HCl = Molarity x volume => 1 x20 = **0.02 moles**
1000 1000
Mole ratio HCl; GCO_3 = 2:1
Moles of GCO_3= 0.02 moles = **0.01moles**
2
Molar mass of GCO_3 = mass => 1 = **100** g
moles 0.01moles
G= GCO_3 - CO_3 =>100g – (12+ 16 x3 = 60) = **40(no units)**

5. 46.0g of a metal carbonate MCO_3 was dissolved 160cm3 of 0.1M excess hydrochloric acid and the resultant solution diluted to one litre.25.0cm3 of this solution required 20.0cm3 of 0.1M sodium hydroxide solution for complete neutralization. Calculate the atomic mass of 'M'

Equation
Chemical equation
$NaOH(aq) + HCl(aq) \rightarrow KCl(aq) + H_2O(l)$
Moles of NaOH = Molarity x volume => 0.1 x20 = **0.002 moles**
1000 1000
Mole ratio HCl; NaOH = 1:1
Excess moles of HCl = **0.002 moles**
25cm3 -> 0.002 moles
1000cm3 -> 1000 x 0.002 = **0.08moles**
25cm3
Original moles of HCl = Molarity x volume => 1M x 1litre = **1.0 moles**

Moles of HCl reacted with MCO_3 = **1.0 - 0.08 moles** = **0.92moles**
Chemical equation
$MCO_{3(s)} + 2HCl_{(aq)} \rightarrow MCl_{2\ (aq)} + CO_2\ (g) + H_2O_{(l)}$
Mole ratio $MCO_{3(s)}$: $HCl_{(aq)}$ =1:2

Moles of MCO_3 = 0.92moles => **0.46moles**
2
Molar mass of MCO_3= mass => 46g = **100 g**

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moles 0.46moles

M= MCO_3 - CO_3 =>100g – (12+ 16 x3 = 60) = **40**

6. 25.0cm3 of a mixture of Fe^{2+} and Fe^{3+} ions in an aqueous salt was acidified with sulphuric(VI)acid then titrated against potassium manganate(VI).The salt required 15cm3 ofe0.02M potassium manganate(VI) for complete reaction.

A second 25cm3 portion of the Fe^{2+} and Fe^{3+} ion salt was reduced by Zinc then titrated against the same concentration of potassium manganate(VI).19.0cm3 of potassium manganate(VI)solution was used for complete reaction. Calculate the concentration of Fe^{2+} and Fe^{3+} ion in the solution on moles per litre.

Mole ratio Fe^{2+} :$Mn0_4^-$ = 5:1

Moles $Mn0_4^-$ used = $\frac{0.02 \times 15}{1000}$ = **3.0×10^{-4}** moles

Moles **Fe^{2+}** = $\frac{3.0 \times 10^{-4} \text{ moles}}{5}$ = **6.0×10^{-5}** moles

Molarity of **Fe^{2+}** = $\frac{\mathbf{6.0 \times 10^{-4}} \text{ moles} \times 1000}{25}$ = **2.4×10^{-3} moles l^{-1}**

Since Zinc reduces **Fe^{3+}** to **Fe^{2+}** in the mixture:

Moles $Mn0_4^-$ that reacted with all **Fe^{2+}**= $\frac{0.02 \times 19}{1000}$ = **3.8×10^{-4}** moles

Moles of all **Fe^{2+}** = $\frac{3.8 \times 10^{-4} \text{ moles}}{5}$ = 7.6×10^{-5} moles

Moles of Fe^{3+} = $3.8 \times 10^{-4} - 6.0 \times 10^{-5}$ = **1.6×10^{-5}** moles

Molarity of **Fe^{3+}** = 1.6×10^{-5} moles x 1000 = **4.0×10^{-4} moles l^{-1}**

CHEMISTRY FORM FOUR NOTES

ACIDS, BASES AND SALTS

A.ACIDS AND BASES

At a school laboratory:

(i)An acid may be defined as a substance that turn litmus red.

(ii)A base may be defined as a substance that turn litmus blue.

Litmus is a lichen found mainly in West Africa. It changes its colour depending on whether the solution it is in, is basic/alkaline or acidic.It is thus able to identify/show whether

1. An acid is a substance that dissolves in water to form H^+/H_3O^+ as the only positive ion/cation. This is called the **Arrhenius definition** of an acid. From this definition, an acid dissociate/ionize in water releasing H^+ thus:

HCl(aq) -> **H^+** (aq) + Cl^- (aq)

HNO_3(aq) -> **H^+** (aq) + NO_3^- (aq)

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$CH_3COOH(aq) \rightarrow H^+(aq) + CH_3COO^-(aq)$
$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$
$H_2CO_3(aq) \rightarrow 2H^+(aq) + CO_3^{2-}(aq)$
$H_3PO_4(aq) \rightarrow 3H^+(aq) + PO_4^{3-}(aq)$

2.A base is a substance which dissolves in water to form OH^- as the only negatively charged ion/anion.
This is called Arrhenius definition of a base.
From this definition, a base dissociate/ionize in water releasing OH^- thus:

$KOH(aq) \rightarrow K^+(aq) + OH^-(aq)$
$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$
$NH_4OH(aq) \rightarrow NH_4^+(aq) + OH^-(aq)$
$Ca(OH)_2(aq) \rightarrow Ca^{2+}(aq) + 2OH^-(aq)$
$Mg(OH)_2(aq) \rightarrow Mg^{2+}(aq) + 2OH^-(aq)$

3. An acid is a proton donor.
A base is a proton acceptor.
This is called Bronsted-Lowry definition of acids and bases.
From this definition, an acid donates H^+ .
H^+ has no electrons and neutrons .It contains only a proton.

Examples

I. From the equation:

$HCl(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Cl^-(aq)$

(a)(i)For the forward reaction from left to right, H_2O gains a proton to form H_3O^+ and thus H_2O is a proton **acceptor** .It is a **Bronsted-Lowry base**
(ii) For the backward reaction from right to left, H_3O^+ donates a proton to form H_2O and thus H_3O^+ is an 'opposite' proton **donor**. It is a **Bronsted-Lowry conjugate acid**
(b)(i)For the forward reaction from left to right, HCl donates a proton to form Cl^- and thus HCl is a proton **donor** .
It is a **Bronsted-Lowry acid**
(ii) For the backward reaction from right to left, Cl^- gains a proton to form HCl and thus Cl^- is an 'opposite' proton **acceptor**.
It is a **Bronsted-Lowry conjugate base.**

Every base /acid from Bronsted-Lowry definition thus must have a conjugate product/reactant.

II. From the equation:

$HCl(aq) + NH_3(aq) \rightleftharpoons NH_4^+(aq) + Cl^-(aq)$

(a)(i)For the forward reaction from left to right, NH_3 gains a proton to form NH_4^+ and thus NH_3 is a proton **acceptor** .
It is a **Bronsted-Lowry base**
(ii) For the backward reaction from right to left, NH_4^+ donates a proton to form NH_3 and thus NH_4^+ is an 'opposite' proton **donor**.

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It is a **Bronsted-Lowry conjugate acid**

(b)(i)For the forward reaction from left to right, HCl donates a proton to form Cl^- and thus HCl is a proton **donor** .
It is a **Bronsted-Lowry acid**
(ii) For the backward reaction from right to left, Cl^- gains a proton to form HCl and thus Cl^- is an 'opposite' proton **acceptor**.
It is a **Bronsted-Lowry conjugate base**.

4. Acids and bases show acidic and alkaline properties/characteristics only in **water** but not in other solvents e.g.

(a)Hydrogen chloride gas dissolves in water to form hydrochloric acid
Hydrochloric acid dissociates/ionizes in water to free $H^+(aq)/H_3O^+(aq)$ ions. The free $H_3O^+(aq)$ / $H^+(aq)$ ions are responsible for:

(i)turning blue litmus paper/solution red.
(ii)show pH value 1/2/3/4/5/6
(iii)are good electrolytes/conductors of electricity/undergo electrolysis.
(iv)react with metals to produce /evolve hydrogen gas and a salt. i.e.
Ionically:
-For a monovalent metal: $2M(s) + 2H^+(aq) \rightarrow 2M^+(aq) + H_2(g)$
-For a divalent metal: $M(s) + 2H^+(aq) \rightarrow M^{2+}(aq) + H_2(g)$
-For a trivalent metal: $2M(s) + 6H^+(aq) \rightarrow 2M^{3+}(aq) + 3H_2(g)$
Examples:
-For a monovalent metal: $2Na(s) + 2H^+(aq) \rightarrow 2Na^+(aq) + H_2(g)$
-For a divalent metal: $Ca(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$
-For a trivalent metal: $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

(v)react with metal carbonates and hhydrogen carbonates to produce /evolve carbon(IV)oxide gas ,water and a salt. i.e.
Ionically:
-For a monovalent metal: $M_2CO_3(s)+ 2H^+(aq) \rightarrow 2M^+(aq) + H_2O(l)+ CO_2(g)$
$MHCO_3(s)+ H^+(aq) \rightarrow M^+(aq) + H_2O(l)+ CO_2(g)$
-For a divalent metal: $MCO_3(s)+ 2H^+(aq) \rightarrow M^{2+}(aq) + H_2O(l)+ CO_2(g)$
$M(HCO_3)_2(\mathbf{aq})+2H^+(aq) \rightarrow M^{2+}(aq)+2H_2O(l)+2CO_2(g)$
Examples:
-For a monovalent metal: $K_2CO_3(s)+ 2H^+(aq) \rightarrow 2K^+(aq) + H_2O(l)+ CO_2(g)$
$NH_4HCO_3(s)+ H^+(aq) \rightarrow NH_4^+(aq) + H_2O(l)+ CO_2(g)$
-For a divalent metal: $ZnCO_3(s)+ 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2O(l)+ CO_2(g)$
$Mg(HCO_3)_2(\mathbf{aq})+2H^+(aq) \rightarrow Mg^{2+}(aq)+2H_2O(l)+2CO_2(g)$

(vi)neutralize metal oxides/hydroxides to salt and water only. i.e.
Ionically:
-For a monovalent metal: $M_2O(s) + 2H^+(aq) \rightarrow 2M^+(aq) + H_2O(l)$
$MOH(\mathbf{aq}) + H^+(aq) \rightarrow M^+(aq) + H_2O(l)$

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-For a divalent metal: $MO(s) + 2H^+(aq) \rightarrow M^{2+}(aq) + H_2O(l)$

$M(OH)_2(s) + 2H^+(aq) \rightarrow M^{2+}(aq) + 2H_2O(l)$

-For a trivalent metal: $M_2O_3(s) + 6H^+(aq) \rightarrow 2M^{3+}(aq) + 3H_2O(l)$

$M(OH)_3(s) + 3H^+(aq) \rightarrow M^{3+}(aq) + 3H_2O(l)$

Examples:

-For a monovalent metal: $K_2O(s) + 2H^+(aq) \rightarrow 2K^+(aq) + H_2O(l)$

$NH_4OH(aq) + H^+(aq) \rightarrow NH_4^+(aq) + H_2O(l)$

-For a divalent metal: $ZnO(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2O(l)$

$Pb(OH)_2(s) + 2H^+(aq) \rightarrow Pb^{2+}(aq) + 2H_2O(l)$

(b)Hydrogen chloride gas dissolves in methylbenzene /benzene but does not dissociate /ionize into free ions.

It exists in molecular state showing none of the above properties.

(c)Ammonia gas dissolves in water to form aqueous ammonia which dissociate/ionize to free NH_4^+ (aq) and OH^-(aq) ions.

This dissociation/ionization makes aqueous ammonia to:

(i)turn litmus paper/solution blue.

(ii)have pH 8/9/10/11

(iii)be a good electrical conductor

(iv)react with acids to form ammonium salt and water only.

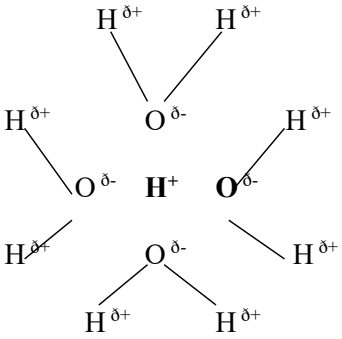
$NH_4OH(aq) + HCl(aq) \rightarrow NH_4Cl(aq) + H_2O(l)$

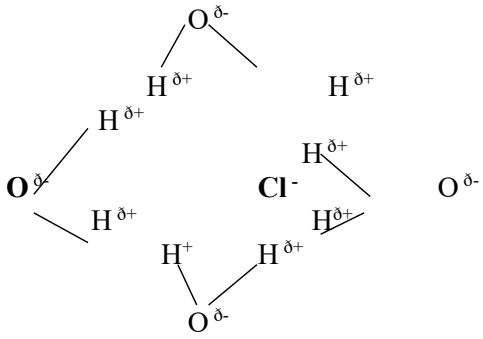
(d)Ammonia gas dissolves in methylbenzene/benzene /kerosene but does not dissociate into free ions therefore existing as molecules

6. Solvents are either **polar** or **non-polar**.

A polar solvent is one which dissolves ionic compounds and other polar solvents.

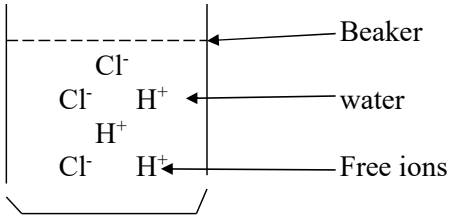
Water is polar solvent that dissolves ionic and polar substance by surrounding the free ions as below:



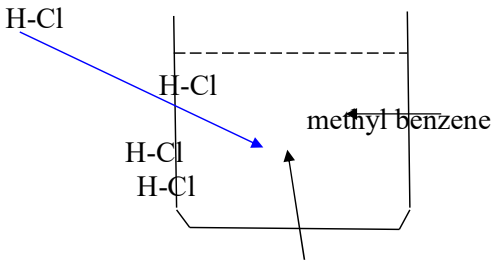


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Note:Water is polar .It is made up of :
Oxygen atom is partially negative and two hydrogen atoms which are partially positive.
They surround the free H^+ and Cl^- ions.
A non polar solvent is one which dissolved non-polar substances and covalent compounds.
If a polar ionic compound is dissolved in non-polar solvent ,it does not ionize/dissociate into free ions as below:



Covalent bond

7. Some acids and bases are **strong** while others are **weak**.
(a)A strong acid/base is one which is fully/wholly/completely dissociated / ionized into many free H^+ /OH^- ions i.e.
I. Strong acids exists more as free $\mathbf{H^+}$ ions than molecules. e.g.

$HCl(aq) \rightleftharpoons H^+(aq) + Cl^-(aq)$
(molecules) (cation) (anion)

$HNO_3(aq) \rightleftharpoons H^+(aq) + NO_3^-(aq)$
(molecules) (cation) (anion)

$H_2SO_4(aq) \rightleftharpoons 2H^+(aq) + SO_4^{2-}(aq)$
(molecules) (cation) (anion)

II. Strong bases/alkalis exists more as free $\mathbf{OH^-}$ ions than molecules. e.g.

$KOH(aq) \rightleftharpoons K^+(aq) + OH^-(aq)$
(molecules) (cation) (anion)

$NaOH(aq) \rightleftharpoons Na^+(aq) + OH^-(aq)$
(molecules) (cation) (anion)

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(b) A weak base/acid is one which is partially /partly dissociated /ionized in water into free OH^- (aq) and H^+(aq) ions.

I. Weak acids exists more as molecules than as free **H^+** ions. e.g.

$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$
(molecules) (cation) (anion)

$H_3PO_4(aq) \rightleftharpoons 3H^+(aq) + PO_4^{3-}(aq)$
(molecules) (cation) (anion)

$H_2CO_3(aq) \rightleftharpoons 2H^+(aq) + CO_3^{2-}(aq)$
(molecules) (cation) (anion)

II. Weak bases/alkalis exists more as molecules than free **OH^-** ions. e.g.

$NH_4OH(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$
(molecules) (cation) (anion)

$Ca(OH)_2(aq) \rightarrow Ca^{2+}(aq) + 2OH^-(aq)$
(molecules) (cation) (anion)

$Mg(OH)_2(aq) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$
(molecules) (cation) (anion)

8. The concentration of an acid/base/alkali is based on the number of moles of acid/bases dissolved in a decimeter(litre)of the solution.
An acid/base/alkali with more acid/base/alkali in a decimeter(litre) of solution is said to be **concentrated** while that with less is said to be **dilute**.

9. (a) (i)strong acids have pH 1/2/3 while weak acids have high pH 4/5/6.
(ii)a neutral solution have pH 7
(iii)strong alkalis/bases have pH 12/13/14 while weak bases/alkalis have pH 11/10 /9 / 8.

(b) pH is a measure of H^+(aq) concentration in a solution.
The higher the H^+(aq)ions concentration ;
-the higher the acidity
-the lower the pH
-the lower the concentration of OH^-(aq)
-the lower the alkalinity
At pH 7 , a solution has **equal** concentration of H^+(aq) and OH^-(aq).
Beyond pH 7,the concentration of the OH^-(aq) increases as the H^+(aq) ions decreases.

10.(a) When acids /bases dissolve in water, the ions present in the solution conduct electricity.
The more the dissociation the higher the yield of ions and the greater the electrical conductivity of the solution.
A compound that conducts electricity in an electrolyte and thus a compound showing high electrical conductivity is a strong electrolyte while a compound showing low electrical conductivity is a weak electrolyte.

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(b) Practically, a bright light on a bulb ,a high voltage reading from a voltmeter high ammeter reading from an ammeter, a big deflection on a galvanometer is an indicator of strong electrolyte(acid/base) and the opposite for weak electrolytes(acids/base)

11. Some compounds exhibit/show both properties of acids and bases/alkalis.

A substance that reacts with both acids and bases is said to be **amphotellic.**

The examples below show the amphotellic properties of:

(a) **Zinc (II)oxide(ZnO) and Zinc hydroxide($Zn(OH)_2$)**

(i) When ½ spatula full of Zinc(II)oxide is placed in a boiling tube containing 10cm3 of either 2M nitric(V)acid or 2M sodium hydroxide hydroxide solution, it dissolves on both the acid and the alkali/base to form a colourless solution. i.e.

(i) when reacting with nitric(V)acid, the oxide shows **basic** properties by reacting with an acid to form a **simple salt** and **water** only.

Basic oxide + Acid -> salt + water

Examples:

Chemical equation

$ZnO(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2O(l)$

$ZnO(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$

$ZnO(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2O(l)$

Ionic equation

$ZnO(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2O(l)$

(ii) when reacting with sodium hydroxide, the oxide shows **acidic** properties by reacting with a base to form a **complex** salt.

Basic oxide + Base/alkali + Water -> Complex salt

Examples:

Chemical equation

1.When Zinc oxide is reacted with sodium hydroxide the complex salt is sodium tetrahydroxozincate(II) complex salt.

$ZnO(s) + 2NaOH(aq) + H_2O(l) \rightarrow Na_2Zn(OH)_4(aq)$

2.When Zinc oxide is reacted with potassium hydroxide the complex salt is potassium tetrahydroxozincate(II) complex salt.

$ZnO(s) + 2KOH(aq) + H_2O(l) \rightarrow K_2Zn(OH)_4(aq)$

Ionic equation

$ZnO(s) + 2OH^-(aq) + H_2O(l) \rightarrow 2[Zn(OH)_4]^{2-}(aq)$

(ii) When Zinc(II)hydroxide is placed in a boiling tube containing 10cm3 of either 2M nitric(V)acid or 2M sodium hydroxide hydroxide solution, it dissolves on both the acid and the alkali/base to form a colourless solution. i.e.

(i) when reacting with nitric(V)acid, the hydroxide shows **basic** properties. It reacts with an acid to form a **simple salt** and **water** only.

Basic hydroxide + Acid -> salt + water

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Examples:
Chemical equation

$Zn(OH)_2(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2H_2O(l)$
$Zn(OH)_2(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + 2H_2O(l)$
$Zn(OH)_2(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + 2H_2O(l)$

Ionic equation

$Zn(OH)_2(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + 2H_2O(l)$

(ii) when reacting with sodium hydroxide, the hydroxide shows **acidic** properties by reacting with a base to form a **complex** salt.

Basic hydroxide + Base/alkali -> Complex salt

Examples:
Chemical equation
1.When Zinc hydroxide is reacted with sodium hydroxide the complex salt is sodium tetrahydroxozincate(II) complex salt.

$Zn(OH)_2(s) + 2NaOH(aq) \rightarrow Na_2Zn(OH)_4(aq)$

2.When Zinc hydroxide is reacted with potassium hydroxide the complex salt is potassium tetrahydroxozincate(II) complex salt.

$Zn(OH)_2(s) + 2KOH(aq) \rightarrow K_2Zn(OH)_4(aq)$

Ionic equation

$Zn(OH)_2(s) + 2OH^-(aq) \rightarrow 2[Zn(OH)_4]^{2-}(aq)$

(b) Lead (II)oxide(PbO) and Lead(II) hydroxide ($Pb(OH)_2$)

(i) When ½ spatula full of Lead(II)oxide is placed in a boiling tube containing 10cm3 of either 2M nitric(V)acid or 2M sodium hydroxide hydroxide solution, it dissolves on both the acid and the alkali/base to form a colourless solution. i.e.

(i) when reacting with nitric(V)acid, the oxide shows **basic** properties by reacting with an acid to form a **simple salt** and **water** only. All other Lead salts are insoluble.

Chemical equation

$PbO(s) + 2HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + H_2O(l)$

Ionic equation

$PbO(s) + 2H^+(aq) \rightarrow Pb^{2+}(aq) + H_2O(l)$

(ii) when reacting with sodium hydroxide, the oxide shows **acidic** properties by reacting with a base to form a **complex** salt.

Chemical equation
1.When Lead(II) oxide is reacted with sodium hydroxide the complex salt is sodium tetrahydroxoplumbate(II) complex salt.

$PbO(s) + 2NaOH(aq) + H_2O(l) \rightarrow Na_2Pb(OH)_4(aq)$

2.When Lead(II) oxide is reacted with potassium hydroxide the complex salt is potassium tetrahydroxoplumbate(II) complex salt.

$PbO(s) + 2KOH(aq) + H_2O(l) \rightarrow K_2Pb(OH)_4(aq)$

Ionic equation

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$PbO(s) + 2OH^-(aq) + H_2O(l) \rightarrow 2[Pb(OH)_4]^{2-}(aq)$

(ii) When Lead(II)hydroxide is placed in a boiling tube containing 10cm3 of either 2M nitric(V)acid or 2M sodium hydroxide hydroxide solution, it dissolves on both the acid and the alkali/base to form a colourless solution. i.e.

(i) when reacting with nitric(V)acid, the hydroxide shows **basic** properties. It reacts with the acid to form a **simple salt** and **water** only.

Chemical equation

$Pb(OH)_2(s) + 2HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + 2H_2O(l)$

Ionic equation

$Pb(OH)_2(s) + 2H^+(aq) \rightarrow Pb^{2+}(aq) + 2H_2O(l)$

(ii) when reacting with sodium hydroxide, the hydroxide shows **acidic** properties. It reacts with a base to form a **complex** salt.

Chemical equation

1.When Lead(II) hydroxide is reacted with sodium hydroxide the complex salt is sodium tetrahydroxoplumbate(II) complex salt.

$Pb(OH)_2(s) + 2NaOH(aq) \rightarrow Na_2Pb(OH)_4(aq)$

2.When Lead(II) hydroxide is reacted with potassium hydroxide the complex salt is potassium tetrahydroxoplumbate(II) complex salt.

$Pb(OH)_2(s) + 2KOH(aq) \rightarrow K_2Pb(OH)_4(aq)$

Ionic equation

$Pb(OH)_2(s) + 2OH^-(aq) \rightarrow 2[Pb(OH)_4]^{2-}(aq)$

(c)Aluminium(III)oxide(Al_2O_3) and Aluminium(III)hydroxide($Al(OH)_3$)

(i) When ½ spatula full of Aluminium(III)oxide is placed in a boiling tube containing 10cm3 of either 2M nitric(V)acid or 2M sodium hydroxide hydroxide solution, it dissolves on both the acid and the alkali/base to form a colourless solution. i.e.

(i) when reacting with nitric(V)acid, the oxide shows **basic** properties by reacting with an acid to form a **simple salt** and **water** only.

Chemical equation

$Al_2O_3(s) + 6HNO_3(aq) \rightarrow Al(NO_3)_3(aq) + 3H_2O(l)$

$Al_2O_3(s) + 6HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$

$Al_2O_3(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2O(l)$

Ionic equation

$Al_2O_3(s) + 3H^+(aq) \rightarrow Al^{3+}(aq) + 3H_2O(l)$

(ii) when reacting with sodium hydroxide, the oxide shows **acidic** properties by reacting with a base to form a **complex** salt.

Chemical equation

1.When Aluminium(III) oxide is reacted with sodium hydroxide the complex salt is sodium tetrahydroxoaluminate(III) complex salt.

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$Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \rightarrow 2NaAl(OH)_4(aq)$

2.When Aluminium(III) oxide is reacted with potassium hydroxide the complex salt is potassium tetrahydroxoaluminate(II) complex salt.

$Al_2O_3(s) + 2KOH(aq) + 3H_2O(l) \rightarrow 2NaAl(OH)_4(aq)$

Ionic equation

$Al_2O_3(s) + 2OH^-(aq) + 3H_2O(l) \rightarrow 2[Al(OH)_4]^-(aq)$

(ii) When Aluminium(III)hydroxide is placed in a boiling tube containing 10cm3 of either 2M nitric(V)acid or 2M sodium hydroxide hydroxide solution, it dissolves on both the acid and the alkali/base to form a colourless solution. i.e.

(i) when reacting with nitric(V)acid, the hydroxide shows **basic** properties. It reacts with the acid to form a **simple salt** and **water** only.

Chemical equation

$Al(OH)_3(s) + 3HNO_3(aq) \rightarrow Al(NO_3)_3(aq) + 3H_2O(l)$

$Al(OH)_3(s) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$

$2Al(OH)_3(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2O(l)$

Ionic equation

$Al(OH)_3(s) + 3H^+(aq) \rightarrow Al^{3+}(aq) + 3H_2O(l)$

(ii) when reacting with sodium hydroxide, the hydroxide shows **acidic** properties. It reacts with a base to form a **complex** salt.

Chemical equation

1.When aluminium(III) hydroxide is reacted with sodium hydroxide the complex salt is sodium tetrahydroxoaluminate(III) complex salt.

$Al(OH)_3(s) + NaOH(aq) \rightarrow NaAl(OH)_4(aq)$

2.When aluminium(III) hydroxide is reacted with potassium hydroxide the complex salt is potassium tetrahydroxoaluminate(III) complex salt.

$Al(OH)_3(s) + KOH(aq) \rightarrow KAl(OH)_4(aq)$

Ionic equation

$Al(OH)_3(s) + OH^-(aq) \rightarrow [Al(OH)_4]^-(aq)$

Summary of amphotellic oxides/hydroxides

Oxide	**Hydroxide**	**Formula of simple salt from nitric (V)acid**	**Formula of complex salt from sodium hydroxide**
ZnO	$Zn(OH)_2$	$Zn(NO_3)_2$	$Na_2Zn(OH)_4$ $[Zn(OH)_4]^{2-}(aq)$ Sodium tetrahydroxozincate(II)
PbO	$Pb(OH)_2$	$Pb(NO_3)_2$	$Na_2Pb(OH)_4$ $[Pb(OH)_4]^{2-}(aq)$ Sodium tetrahydroxoplumbate(II)

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Al_2O_3	$Al(OH)_3$	$Al(NO_3)_3$	$NaAl(OH)_4$ $[Al(OH)_4]^-(aq)$ Sodium tetrahydroxoaluminate(II)

12.(a) A salt is an ionic compound formed when the cation from a base combine with the anion derived from an acid.
A salt is therefore formed when the hydrogen ions in an acid are replaced wholly/fully or partially/partly ,directly or indirectly by a metal or ammonium radical. (b) The number of ionizable/replaceable hydrogen in an acid is called basicity of an acid.
Some acids are therefore:

(i)**mono**basic acids generally denoted **H**X e.g.
HCl, **H**NO_3,HCOO**H**,CH3COO**H**.
(ii)**di**basic acids ; generally denoted $\mathbf{H_2}$X e.g.
$\mathbf{H_2}SO_4$, $\mathbf{H_2}SO_3$, $\mathbf{H_2}CO_3$,**H**OOCOO**H**.
(iii)**tri**basic acids ; generally denoted $\mathbf{H_3}$X e.g.
$\mathbf{H_3}PO_4$.

(c) Some salts are **normal** salts while other are **acid** salts.
(i)A normal salt is formed when all the ionizable /replaceable hydrogen in an acid is replaced by a metal or metallic /ammonium radical.
(ii)An acid salt is formed when part/portion the ionizable /replaceable hydrogen in an acid is replaced by a metal or metallic /ammonium radical.

Table showing normal and acid salts derived from common acids

Acid name	**Chemical formula**	**Basicity**	**Normal salt**	**Acid salt**
Hydrochloric acid	HCl	Monobasic	Chloride(Cl^-)	None
Nitric(V)acid	HNO_3	Monobasic	Nitrate(V)(NO_3^-)	None
Nitric(III)acid	HNO_2	Monobasic	Nitrate(III)(NO_2^-)	None
Sulphuric(VI) acid	H_2SO_4	Dibasic	Sulphate(VI) (SO_4^{2-})	Hydrogen sulphate(VI) (HSO_4^-)
Sulphuric(IV) acid	H_2SO_3	Dibasic	Sulphate(IV) (SO_3^{2-})	Hydrogen sulphate(IV)

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				(HSO_3^-)
Carbonic(IV)acid	H_2CO_3	Dibasic	Carbonate(IV)(CO_3^{2-})	Hydrogen carbonate(IV) (HCO_3^-)
Phosphoric(V) acid	H_3PO_4	Tribasic	Phosphate(V)(PO_4^{3-})	Dihydrogen phosphate(V) ($H_2PO_4^{2-}$) Hydrogen diphosphate(V) ($HP_2O_4^{2-}$)

The table below show shows some examples of salts.

Base/alkali	Cation	Acid	Anion	Salt	Chemical name of salts
NaOH	Na^+	HCl	Cl^-	NaCl	Sodium(I)chloride
$Mg(OH)_2$	Mg^{2+}	H_2SO_4	SO_4^{2-}	$MgSO_4$ $Mg(HSO_4)_2$	Magnesium sulphate(VI) Magnesium hydrogen sulphate(VI)
$Pb(OH)_2$	Pb^{2+}	HNO_3	NO_3^-	$Pb(NO_3)_2$	Lead(II)nitrate(V)
$Ba(OH)_2$	Ba^{2+}	HNO_3	NO_3^-	$Ba(NO_3)_2$	Barium(II)nitrate(V)
$Ca(OH)_2$	Ba^{2+}	H_2SO_4	SO_4^{2-}	$MgSO_4$	Calcium sulphate(VI)
NH_4OH	NH_4^+	H_3PO_4	PO_4^{3-}	$(NH_4)_3PO_4$ $(NH_4)_2HPO_4$ $NH_4H_2PO_4$	Ammonium phosphate(V) Diammonium phosphate(V) Ammonium diphosphate(V)
KOH	K^+	H_3PO_4	PO_4^{3-}	K_3PO_4	Potassium phosphate(V)
$Al(OH)_3$	Al^{3+}	H_2SO_4	SO_4^{2-}	$Al_2(SO_4)_2$	Aluminium(III)sulphate(VI)
$Fe(OH)_2$	Fe^{2+}	H_2SO_4	SO_4^{2-}	$FeSO_4$	Iron(II)sulphate(VI)

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$Fe(OH)_3$	Fe^{3+}	H_2SO_4	SO_4^{2-}	$Fe_2(SO_4)_2$	Iron(III)sulphate(VI)

(d) Some salts undergo **hygroscopy**, **deliquescence** and **efflorescence**.
(i) Hygroscopic salts /compounds are those that absorb water from the atmosphere but do not form a solution.
Some salts which are hygroscopic include anhydrous copper(II)sulphate(VI), anhydrous cobalt(II)chloride, potassium nitrate(V) common table salt.
(ii)Deliquescent salts /compounds are those that absorb water from the atmosphere and form a solution.
Some salts which are deliquescent include: Sodium nitrate(V),Calcium chloride, Sodium hydroxide, Iron(II)chloride, Magnesium chloride.
(iii)Efflorescent salts/compounds are those that lose their water of crystallization to the atmosphere.
Some salts which effloresces include: sodium carbonate decahydrate, Iron(II)sulphate(VI)heptahydrate, sodium sulphate (VI)decahydrate.
(e)Some salts contain water of crystallization.They are hydrated.Others do not contain water of crystallization. They are anhydrous.

Table showing some hydrated salts.

Name of hydrated salt	Chemical formula
Copper(II)sulphate(VI)pentahydrate	$CuSO_4.5H_2O$
Aluminium(III)sulphate(VI)hexahydrate	$Al_2(SO_4)_3.6H_2O$
Zinc(II)sulphate(VI)heptahydrate	$ZnSO_4.7H_2O$
Iron(II)sulphate(VI)heptahydrate	$FeSO_4.7H_2O$
Calcium(II)sulphate(VI)heptahydrate	$CaSO_4.7H_2O$
Magnesium(II)sulphate(VI)heptahydrate	$MgSO_4.7H_2O$
Sodium sulphate(VI)decahydrate	$Na_2SO_4.10H_2O$
Sodium carbonate(IV)decahydrate	$Na_2CO_3.10H_2O$
Potassium carbonate(IV)decahydrate	$K_2CO_3.10H_2O$
Potassium sulphate(VI)decahydrate	$K_2SO_4.10H_2O$

(f)Some salts exist as a simple salt while some as complex salts. Below are some complex salts.

Table of some complex salts

Name of complex salt	Chemical formula	Colour of the complex salt
Tetraamminecopper(II)sulphate(VI)	$Cu(NH_3)_4SO_4H_2O$	Royal/deep blue solution
Tetraamminezinc(II)nitrate(V)	$Zn(NH_3)_4(NO_3)_2$	Colourless solution
Tetraamminecopper(II) nitrate(V)	$Cu(NH_3)_4(NO_3)_2$	Royal/deep blue solution

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Tetraamminezinc(II)sulphate(VI)	$Zn(NH_3)_4SO_4$	Colourless solution

(g)Some salts exist as two salts in one. They are called **double salts**.

Table of some double salts

Name of double salts	Chemical formula
Trona(sodium sesquicarbonate)	$Na_2CO_3\ NaHCO_3.2H_2O$
Ammonium iron(II)sulphate(VI)	$FeSO_4(NH_4)_2SO_4.2H_2O$
Ammonium aluminium(III)sulphate(VI)	$Al_2(SO_4)_3(NH_4)_2SO_4.H_2O$

(h)Some salts dissolve in water to form a **solution**. They are said to be soluble. Others do not dissolve in water. They form a **suspension/precipitate** in water.

Table of solubility of salts

Soluble salts	Insoluble salts
All nitrate(V)salts	
All sulphate(VI)/SO_4^{2-} salts **except** →	Barium(II) sulphate(VI)/$BaSO_4$ Calcium(II) sulphate(VI)/$CaSO_4$ Lead(II) sulphate(VI)/$PbSO_4$
All sulphate(IV)/SO_3^{2-} salts **except** →	Barium(II) sulphate(IV)/$BaSO_3$ Calcium(II) sulphate(IV)/$CaSO_3$ Lead(II) sulphate(IV)/$PbSO_3$
All chlorides/Cl^- **except** →	Silver chloride/AgCl Lead(II)chloride/$PbCl_2$(dissolves in hot water)
All phosphate(V)/PO_4^{3-}	
All sodium,potassium and ammonium salts	
All hydrogen carbonates/HCO_3^-	
All hydrogen sulphate(VI)/ HSO_4^-	
Sodium carbonate/Na_2CO_3, potassium carbonate/ K_2CO_3, ammonium carbonate $(NH_4)_2CO_3$ ←	**except** All carbonates
All alkalis(KOH,NaOH, NH_4OH) ←	**except** All bases

13 Salts can be prepared in a school laboratory by a method that uses its solubility in water.

(b)Soluble salts may be prepared by using any of the following methods:

(i)Direct displacement/reaction of a metal with an acid.

By reacting a metal higher in the reactivity series than hydrogen with a dilute acid,a salt is formed and hydrogen gas is evolved.

Excess of the metal must be used to ensure all the acid has reacted.

When effervescence/bubbling /fizzing has stopped ,excess metal is filtered.

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The filtrate is heated to concentrate then allowed to crystallize.
Washing with distilled water then drying between filter papers produces a sample crystal of the salt. i.e.

$$M(s) + H_2X \rightarrow MX(aq) + H_2(g)$$

Examples

$$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$$
$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$
$$Pb(s) + 2HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + H_2(g)$$
$$Ca(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2(g)$$
$$Mg(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$$
$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$
$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

(ii)Reaction of an insoluble base with an acid

By adding an insoluble base (oxide/hydroxide)to a dilute acid until **no more dissolves**, in the acid,a salt and water are formed. Excess of the base is filtered off. The filtrate is heated to concentrate ,allowed to crystallize then washed with distilled water before drying between filter papers e.g.

$$PbO(s) + 2HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + H_2O(l)$$
$$Pb(OH)_2(s) + 2HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + 2H_2O(l)$$
$$CaO(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l)$$
$$MgO(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2O(l)$$
$$MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$$
$$ZnO(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$$
$$Zn(OH)_2(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2H_2O(l)$$
$$CuO(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2O(l)$$
$$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$
$$Ag_2O(s) + 2HNO_3(aq) \rightarrow 2AgNO_3(aq) + H_2O(l)$$
$$Na_2O(s) + 2HNO_3(aq) \rightarrow 2NaNO_3(aq) + H_2O(l)$$

(iii)reaction of insoluble /soluble carbonate /hydrogen carbonate with an acid.

By adding an excess of a soluble /insoluble carbonate or hydrogen carbonate to adilute acid, effervescence /fizzing/bubbling out of carbon(IV)oxide gas shows the reaction is taking place. When effervescence /fizzing/bubbling out of the gas is over, excess of the insoluble carbonate is filtered off. The filtrate is heated to concentrate ,allowed to crystallize then washed with distilled water before drying between filter paper papers e.g.

$$PbCO_3(s) + 2HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + H_2O(l) + CO_2(g)$$
$$ZnCO_3(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2O(l) + CO_2(g)$$
$$CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l) + CO_2(g)$$
$$MgCO_3(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2O(l) + CO_2(g)$$
$$CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l) + CO_2(g)$$
$$Ag_2CO_3(s) + 2HNO_3(aq) \rightarrow 2AgNO_3(aq) + H_2O(l) + CO_2(g)$$

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Na_2CO_3 (s) + $2HNO_3$(aq) -> $2NaNO_3$(aq) + H_2O (l) + CO_2(g)
K_2CO_3 (s) + 2HCl(aq) -> 2KCl(aq) + H_2O (l) + CO_2(g)
$NaHCO_3$ (s) + HNO_3(aq) -> $NaNO_3$(aq) + H_2O (l) + CO_2(g)
$KHCO_3$ (s) + HCl(aq) -> KCl(aq) + H_2O (l) + CO_2(g)

(iv)neutralization/reaction of soluble base/alkali with dilute acid

By adding an acid to a burette into a known volume of an alkali with 2-3 drops of an indicator, the colour of the indicator changes when the acid has completely reacted with an alkali at the **end point.** The procedure is then repeated without the indicator .The solution mixture is then heated to concentrate , allowed to crystallize ,washed with distilled water before drying with filter papers. e.g.

NaOH (aq) + HNO_3(aq) -> $NaNO_3$(aq) + H_2O (l)
KOH (aq) + HNO_3(aq) -> KNO_3(aq) + H_2O (l)
KOH (aq) + HCl(aq) -> KCl(aq) + H_2O (l)
2KOH (aq) + H_2SO_4(aq) -> K_2SO_4(aq) + $2H_2O$ (l)
2 NH_4OH (aq) + H_2SO_4(aq) -> $(NH_4)_2SO_4$(aq) + $2H_2O$ (l)
NH_4OH (aq) + HNO_3(aq) -> NH_4NO_3(aq) + H_2O (l)

(iv)direct synthesis/combination.

When a metal **burn** in a gas jar containing a non metal , the two directly combine to form a salt. e.g.

2Na(s) + Cl_2(g) -> 2NaCl(s)
2K(s) + Cl_2(g) -> 2KCl(s)
Mg(s) + Cl_2(g) -> $Mg Cl_2$ (s)
Ca(s) + Cl_2(g) -> $Ca Cl_2$ (s)

Some salts once formed undergo **sublimation** and **hydrolysis**. Care should be taken to avoid water/moisture into the reaction flask during their preparation.Such salts include aluminium(III)chloride($AlCl_3$) and iron (III)chloride($FeCl_3$)

1. Heated aluminium foil reacts with chlorine to form aluminium(III)chloride that sublimes away from the source of heating then deposited as solid again

2Al(s) + $3Cl_2$(g) -> $2AlCl_3$ (**s/g**)

Once formed aluminium(III)chloride hydrolyses/reacts with water vapour / moisture present to form aluminium hydroxide solution and highly acidic fumes of hydrogen chloride gas.

$AlCl_3$(s)+ $3H_2O$(g) -> $Al(OH)_3$ (**aq**) + 3HCl(g)

2. Heated iron filings reacts with chlorine to form iron(**III**)chloride that sublimes away from the source of heating then deposited as solid again

2Fe(s) + $3Cl_2$(g) -> $2FeCl_3$ (**s/g**)

Once formed , aluminium(III)chloride hydrolyses/reacts with water vapour / moisture present to form aluminium hydroxide solution and highly acidic fumes of hydrogen chloride gas.

$FeCl_3$(s)+ $3H_2O$(g) -> $Fe(OH)_3$ (**aq**) + 3HCl(g)

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(b)**Insoluble** salts can be prepared by reacting **two** suitable **soluble** salts to form **one soluble** and **one insoluble**. This is called **double decomposition** or **precipitation**. The mixture is filtered and the **residue** is washed with distilled water then dried.

CuSO_4(aq) + Na_2**CO_3** (aq) -> **$CuCO_3$ (s)** + Na_2 SO_4(aq)
BaCl_2(aq) + K_2**SO_4** (aq) -> **$BaSO_4$ (s)** + 2KCl (aq)
Pb$(NO_3)_2$(aq) + K_2**SO_4** (aq) -> **$PbSO_4$ (s)** + 2KNO_3 (aq)
2**Ag**NO_3(aq) + Mg**Cl_2** (aq) -> 2**AgCl(s)** + $Mg(NO_3)_2$ (aq)
Pb$(NO_3)_2$(aq) + $(NH_4)_2$**SO_4** (aq) -> **$PbSO_4$ (s)** + 2NH_4NO_3(aq)
BaCl_2(aq) + K_2**SO_3** (aq) -> **$BaSO_3$ (s)** + 2KCl (aq)

14. Salts may lose their water of crystallization , decompose ,melt or sublime on heating on a Bunsen burner flame.

The following shows the behavior of some salts on heating gently /or strongly in a laboratory school burner:

(a)effect of heat on chlorides

All chlorides have very high melting and boiling points and therefore are not affected by laboratory heating except ammonium chloride. Ammonium chloride **sublimes** on **gentle** heating. It **dissociate** into the constituent **ammonia** and **hydrogen chloride** gases on strong heating.

$NH_4Cl(s) \rightleftharpoons NH_4Cl(g) \rightleftharpoons NH_3(g) + HCl(g)$
(sublimation) (dissociation)

(b)effect of heat on nitrate(V)

(i) Potassium nitrate(V)/KNO_3 and sodium nitrate(V)/$NaNO_3$ decompose on heating to form Potassium nitrate(III)/KNO_2 and sodium nitrate(III)/$NaNO_2$ and producing Oxygen gas in each case.

2**KNO_3** (s) -> 2KNO_2(s) + O_2(g)
2**$NaNO_3$** (s) -> 2$NaNO_2$(s) + O_2(g)

(ii)Heavy metal nitrates(V) salts decompose on heating to form the oxide and a mixture of brown acidic nitrogen(IV)oxide and oxygen gases. e.g.

2$Ca(NO_3)_2$ (s) -> 2CaO(s) + 4NO_2(g) + O_2(g)
2$Mg(NO_3)_2$(s) -> 2MgO(s) + 4NO_2(g) + O_2(g)
2$Zn(NO_3)_2$(s) -> 2ZnO(s) + 4NO_2(g) + O_2(g)
2$Pb(NO_3)_2$(s) -> 2PbO(s) + 4NO_2(g) + O_2(g)
2$Cu(NO_3)_2$(s) -> 2CuO(s) + 4NO_2(g) + O_2(g)
2$Fe(NO_3)_2$(s) -> 2FeO(s) + 4NO_2(g) + O_2(g)

(iii)Silver(I)nitrate(V) and mercury(II) nitrate(V) are lowest in the reactivity series. They decompose on heating to form the **metal**(silver and mercury)and the Nitrogen(IV)oxide and oxygen gas. i.e.

2$AgNO_3$(s) -> 2Ag (s) + 2NO_2(g) + O_2(g)
2$Hg(NO_3)_2$ (s) -> 2Hg (s) + 4NO_2(g) + O_2(g)

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(iv)Ammonium nitrate(V) and Ammonium nitrate(III) decompose on heating to Nitrogen(I)oxide(relights/rekindles glowing splint) and nitrogen gas respectively.Water is also formed.i.e.

$NH_4NO_3(s) \rightarrow N_2O(g) + H_2O(l)$

$NH_4NO_2(s) \rightarrow N_2(g) + H_2O(l)$

(c) effect of heat on nitrate(V)

Only Iron(II)sulphate(VI), Iron(III)sulphate(VI) and copper(II)sulphate(VI) decompose on heating. They form the **oxide**, and produce highly acidic fumes of acidic **sulphur(IV)oxide** gas.

$2FeSO_4(s) \rightarrow Fe_2O_3(s) + SO_3(g) + SO_2(g)$

$Fe_2(SO_4)_3(s) \rightarrow Fe_2O_3(s) + SO_3(g)$

$CuSO_4(s) \rightarrow CuO(s) + SO_3(g)$

(d) effect of heat on carbonates(IV) and hydrogen carbonate(IV).

(i)Sodium carbonate(IV)and potassium carbonate(IV)**do not decompose** on heating.

(ii)Heavy metal nitrate(IV)salts decompose on heating to form the **oxide** and produce **carbon(IV)oxide** gas. Carbon (IV)oxide gas forms a white precipitate when bubbled in lime water. The white precipitate dissolves if the gas is in excess.

e.g. $CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

$PbCO_3(s) \rightarrow PbO(s) + CO_2(g)$

$FeCO_3(s) \rightarrow FeO(s) + CO_2(g)$

$ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$

(iii)Sodium hydrogen carbonate(IV) and Potassium hydrogen carbonate(IV)decompose on heating to give the corresponding carbonate (IV) and form water and carbon(IV)oxide gas. i.e.

$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(l)$

$2KHCO_3(s) \rightarrow K_2CO_3(s) + CO_2(g) + H_2O(l)$

(iii) Calcium hydrogen carbonate (IV) and Magnesium hydrogen carbonate(IV) decompose on heating to give the corresponding carbonate (IV) and form water and carbon(IV)oxide gas. i. e.

$Ca(HCO_3)_2(aq) \rightarrow CaCO_3(s) + CO_2(g) + H_2O(l)$

$Mg(HCO_3)_2(aq) \rightarrow MgCO_3(s) + CO_2(g) + H_2O(l)$

15. Salts contain cation(positively charged ion) and anions(negatively charged ion).When dissolved in polar solvents/water.

The cation and anion in a salt is determined/known usually by precipitation of the salt using a **precipitating reagent**.

The colour of the precipitate is a basis of qualitative analysis of a compound.

16.Qualitative analysis is the process of identifying an unknown compound /salt by identifying the unique qualities of the salt/compound.

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It involves some of the following processes.

(a)Reaction of cation with sodium/potassium hydroxide solution.

Both sodium/potassium hydroxide solutions are precipitating reagents.

The alkalis produce **unique** colour of a precipitate/suspension when a few/three drops is added and then excess alkali is added to **unknown** salt/compound solution.

NB: Potassium hydroxide is not commonly used because it is more expensive than sodium hydroxide.

The table below shows the observations, inferences / deductions and explanations from the following test tube experiments:

Procedure

Put about 2cm3 of $MgCl_2$, $CaCl_2$, $AlCl_3$, NaCl, KCl, $FeSO_4$, $Fe_2(SO_4)_3$, $CuSO_4$, $ZnSO_4NH_4NO_3$, $Pb(NO_3)_2$, $Ba(NO_3)_2$ each into separate test tubes. Add three drops of 2M sodium hydroxide solution then excess ($^2/_3$ the length of a standard test tube).

Observation	**Inference**	**Explanation**
No white precipitate	Na^+ and K^+	Both Na^+ and K^+ ions react with OH^- from 2M sodium hydroxide solution to form soluble colourless solutions $Na^+(aq) + OH^-(aq) \rightarrow NaOH(aq)$ $K^+(aq) + OH^-(aq) \rightarrow KOH(aq)$
No white precipitate then pungent smell of ammonia /urine	NH_4^+ ions	NH_4^+ ions react with 2M sodium hydroxide solution to produce pungent smelling ammonia gas $NH_4^+(aq) + OH^-(aq) \rightarrow NH_3(g) + H_2O(l)$
White precipitate insoluble in excess	Ba^{2+}, Ca^{2+}, Mg^{2+} ions	Ba^{2+}, Ca^{2+} and Mg^{2+} ions react with OH^- from 2M sodium hydroxide solution to form insoluble white precipitate of their hydroxides. $Ba^{2+}(aq) + 2OH^-(aq) \rightarrow Ba(OH)_2(\mathbf{s})$ $Ca^{2+}(aq) + 2OH^-(aq) \rightarrow Ca(OH)_2(\mathbf{s})$ $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(\mathbf{s})$
White precipitate soluble in excess	Zn^{2+}, Pb^{2+}, Al^{3+} ions	Pb^{2+}, Zn^{2+} and Al^{3+} ions react with OH^- from 2M sodium hydroxide solution to form insoluble white precipitate of their hydroxides.

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		$Zn^{2+}(aq) + 2OH^-(aq) \rightarrow Zn(OH)_2(s)$ $Pb^{2+}(aq) + 2OH^-(aq) \rightarrow Pb(OH)_2(s)$ $Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$ The hydroxides formed react with more OH^- ions to form complex salts/ions. $Zn(OH)_2(s) + 2OH(aq) \rightarrow [Zn(OH)_4]^{2-}(aq)$ $Pb(OH)_2(s) + 2OH(aq) \rightarrow [Pb(OH)_4]^{2-}(aq)$ $Al(OH)_3(s) + OH(aq) \rightarrow [Al(OH)_4]^-(aq)$
Blue precipitate insoluble in excess	Cu^{2+}	Cu^{2+} ions react with OH^- from 2M sodium hydroxide solution to form insoluble blue precipitate of copper(II) hydroxide. $Cu^{2+}(aq) + 2OH^-(aq) \rightarrow \mathbf{Cu(OH)_2(s)}$
Green precipitate insoluble in excess On adding 3cm3 of hydrogen peroxide, **brown**/yellow solution formed	Fe^{2+} Fe^{2+} oxidized to Fe^{3+}	Fe^{2+} ions react with OH^- from 2M sodium hydroxide solution to form insoluble green precipitate of Iron(II) hydroxide. $Fe^{2+}(aq) + 2OH^-(aq) \rightarrow \mathbf{Fe(OH)_2(s)}$ Hydrogen peroxide is an oxidizing agent that oxidizes green Fe^{2+} oxidized to brown Fe^{3+} $\mathbf{Fe(OH)_2(s)} + 2H^+ \rightarrow \mathbf{Fe(OH)_3(aq)}$
Brown precipitate insoluble in excess	Fe^{3+}	Fe^{3+} ions react with OH^- from 2M sodium hydroxide solution to form insoluble **brown** precipitate of **Iron(II) hydroxide**. $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow \mathbf{Fe(OH)_3(s)}$

(b)Reaction of cation with aqueous ammonia

Aqueous ammonia precipitating reagent that can be used to identify the cations present in a salt.

Like NaOH/KOH the OH^- ion in NH_4OH react with the cation to form a characteristic hydroxide .

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Below are the observations ,inferences and explanations of the reactions of aqueous ammonia with salts from the following test tube reactions.

Procedure

Put about 2cm3 of $MgCl_2$, $CaCl_2$, $AlCl_3$, NaCl, KCl, $FeSO_4$, $Fe_2(SO_4)_3$, $CuSO_4$, $ZnSO_4NH_4NO_3$, $Pb(NO_3)_2$, $Ba(NO_3)_2$ each into separate test tubes.

Add three drops of 2M aqueous ammonia then excess ($^2/_3$ the length of a standard test tube).

Observation	**Inference**	**Explanation**
No white precipitate	Na^+ and K^+	NH_4^+,Na^+ and K^+ ions react with OH^- from 2M aqueous ammonia to form soluble colourless solutions $NH_4^+(aq) + OH^-(aq) \rightarrow NH_4^+OH(aq)$ $Na^+(aq) + OH^-(aq) \rightarrow NaOH(aq)$ $K^+(aq) + OH^-(aq) \rightarrow KOH(aq)$
White precipitate insoluble in excess	Ba^{2+} ,Ca^{2+}, Mg^{2+} ,Pb^{2+}, Al^{3+}, ions	Ba^{2+} ,Ca^{2+},Mg^{2+} ,Pb^{2+} and Al^{3+}, ions react with OH^- from 2M aqueous ammonia to form insoluble white precipitate of their hydroxides. $Pb^{2+}(aq) + 2OH^-(aq) \rightarrow Pb(OH)_2(\mathbf{s})$ $Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(\mathbf{s})$ $Ba^{2+}(aq) + 2OH^-(aq) \rightarrow Ba(OH)_2(\mathbf{s})$ $Ca^{2+}(aq) + 2OH^-(aq) \rightarrow Ca(OH)_2(\mathbf{s})$ $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(\mathbf{s})$
White precipitate soluble in excess	Zn^{2+} ions	Zn^{2+} ions react with OH^- from 2M aqueous ammonia to form insoluble white precipitate of Zinc hydroxide. $Zn^{2+}(aq) + 2OH^-(aq) \rightarrow Zn(OH)_2(\mathbf{s})$ The Zinc hydroxides formed react $NH_3(aq)$ to form a complex salts/ions. $Zn(OH)_2(\mathbf{s}) + 4NH_3(aq) \rightarrow [Zn(NH_3)_4]^{2+}(\mathbf{aq}) + 2OH^-(aq)$
Blue precipitate that dissolves in	Cu^{2+}	Cu^{2+} ions react with OH^- from 2M aqueous ammonia to form blue precipitate of

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excess ammonia solution to form a **deep/royal blue solution**		copper(II) hydroxide. $Cu^{2+}(aq) + 2OH^-(aq) \rightarrow \mathbf{Cu(OH)_2(s)}$ The copper(II) hydroxide formed react $NH_3(aq)$ to form a complex salts/ions. $\mathbf{Cu(OH)_2(s)} + 4NH_3(aq) \rightarrow \mathbf{[Cu(NH_3)_4]^{2+}(aq)} + 2OH^-(aq)$
Green precipitate insoluble in excess. On adding 3cm3 of hydrogen peroxide, **brown**/yellow solution formed	Fe^{2+} Fe^{2+} oxidized to Fe^{3+}	Fe^{2+} ions react with OH^- from 2M aqueous ammonia to form insoluble green precipitate of Iron(II) hydroxide. $Fe^{2+}(aq) + 2OH^-(aq) \rightarrow \mathbf{Fe(OH)_2(s)}$ Hydrogen peroxide is an oxidizing agent that oxidizes green Fe^{2+} oxidized to brown Fe^{3+} $\mathbf{Fe(OH)_2(s)} + 2H^+ \rightarrow \mathbf{Fe(OH)_3(aq)}$
Brown precipitate insoluble in excess	Fe^{3+}	Fe^{3+} ions react with OH^- from 2M aqueous ammonia to form insoluble **brown** precipitate of **Iron(II) hydroxide**. $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow \mathbf{Fe(OH)_3(s)}$

Note

(i) Only $\mathbf{Zn^{2+}}$ ions/salts form a **white precipitate** that **dissolve** in excess of both 2M sodium hydroxide and 2M aqueous ammonia.

(ii) $\mathbf{Pb^{2+}}$ and $\mathbf{Al^{3+}}$ ions/salts form a **white precipitate** that **dissolve** in excess of 2M sodium hydroxide but not in 2M aqueous ammonia.

(iii) $\mathbf{Cu^{2+}}$ ions/salts form a **blue precipitate** that **dissolve** to form a **deep/royal blue** solution in excess of 2M aqueous ammonia but only **blue** insoluble **precipitate** in 2M sodium hydroxide

(c)Reaction of cation with Chloride (Cl^-)ions

All chlorides are soluble in water except **Silver chloride** and **Lead (II)chloride** (That dissolve in hot water).When a soluble chloride like NaCl, KCl, NH_4Cl is added to about 2cm3 of a salt containing Ag^+ or Pb^{2+}ions a white precipitate of AgCl or $PbCl_2$ is formed. The following test tube reactions illustrate the above.

Experiment

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Put about 2cm3 of silver nitrate(V) andLead(II)nitrate(V)solution into separate test tubes. Add five drops of NaCl /KCl / NH_4Cl/HCl. Heat to boil.

Observation	**Inference**	**Explanation**
(i)White precipitate does not dissolve on heating	Ag^+ ions	Ag^+ ions reacts with Cl^- ions from a soluble chloride salt to form a white precipitate of AgCl
(ii)White precipitate dissolve on heating	Pb^{2+} ions	Pb^{2+} ions reacts with Cl^- ions from a soluble chloride salt to form a white precipitate of $PbCl_2$. $PbCl_2$ dissolves on heating.

Note

Both Pb^{2+} and Al^{3+} ions forms an insoluble white precipitate in excess aqueous ammonia. A white precipitate on adding Cl^- ions/salts shows **Pb^{2+}**.

No white precipitate on adding Cl^- ions/salts shows **Al^{3+}**.

Adding a chloride/ Cl^- ions/salts can thus be used to separate the identity of **Al^{3+} and Pb^{2+}**.

(d)Reaction of cation with sulphate(VI)/SO_4^{2-} and sulphate(IV)/SO_3^{2-} ions

All sulphate(VI) and sulphate(IV)/SO_3^{2-} ions/salts are soluble/dissolve in water **except** Calcium sulphate(VI)/$CaSO_4$**,** Calcium sulphate(IV)/$CaSO_3$**,** Barium sulphate(VI)/$BaSO_4$**,** Barium sulphate(IV)/$BaSO_3$**,** Lead(II) sulphate(VI)/$PbSO_4$ and Lead(II) sulphate(IV)/$PbSO_3$.When a soluble sulphate(VI)/SO_4^{2-} salt like Na_2SO_4, H_2SO_4, $(NH_4)_2SO_4$ or Na_2SO_3 is added to a salt containing Ca^{2+}, Pb^{2+}, Ba^{2+} ions, a white precipitate is formed.

The following test tube experiments illustrate the above.

Procedure

Place about 2cm3 of $Ca(NO_3)_2$, $Ba(NO_3)_2$, $BaCl_2$ and $Pb(NO_3)_2$, in separate boiling tubes. Add six drops of sulphuric(VI)acid /sodium sulphate(VI)/ammonium sulphate(VI)solution. Repeat with six drops of sodium sulphate(IV).

Observation	Inference	Explanation
White precipitate	Ca^{2+}, Ba^{2+}, Pb^{2+} ions	$CaSO_3$ and $CaSO_4$ do not form a thick precipitate as they are sparingly soluble. $Ca^{2+}(aq) + SO_3^{2-}(aq) \rightarrow CaSO_3(s)$ $Ca^{2+}(aq) + SO_4^{2-}(aq) \rightarrow CaSO_4(s)$ $Ba^{2+}(aq) + SO_3^{2-}(aq) \rightarrow BaSO_3(s)$ $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$ $Pb^{2+}(aq) + SO_3^{2-}(aq) \rightarrow PbSO_3(s)$ $Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$

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(e)Reaction of cation with carbonate(IV)/CO_3^{2-} ions

All carbonate salts are insoluble except sodium/potassium carbonate(IV) and ammonium carbonate(IV).

They dissociate /ionize to release CO_3^{2-} ions. CO_3^{2-} ions produce a white precipitate when the soluble carbonate salts is added to any metallic cation.

Procedure

Place about 2cm3 of $Ca(NO_3)_2$, $Ba(NO_3)_2$, $MgCl_2$,$Pb(NO_3)_2$ and$ZnSO_4$ in separate boiling tubes.

Add six drops of Potassium /sodium carbonate(IV)/ ammonium carbonate (IV)solution.

Observation	Inference	Explanation
Green precipitate	Cu^{2+} ,Fe^{2+},ions CO_3^{2-}(aq)	Copper(II)carbonate(IV) and Iron(II) carbonate (IV) are precipitated as insoluble **green** precipitates. $Cu^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CuCO_3(s)$ $Fe^{2+}(aq) + CO_3^{2-}(aq) \rightarrow FeCO_3(s)$ When sodium carbonate(IV)is added to $CuCO_3(s)$ the $CO_3^{2-}(aq)$ ions are first **hydrolysed** to produce $CO_2(g)$ and OH^- (aq)ions. $CO_3^{2-}(aq) + H_2O(l) \rightarrow CO_2$ **(g)** $+ 2OH^-$ (aq) The $OH^-(aq)$ ions further react to form basic copper(II) carbonate(IV). Basic copper(II) carbonate(IV) is the **only green** salt of copper. $Cu^{2+}(aq) + CO_3^{2-}(aq) + 2OH^-(aq) \rightarrow$ **$CuCO_3.Cu(OH)_2$** (s)
White precipitate	CO_3^{2-}	White ppt of the carbonate(IV)salt is precipitated $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$ $Mg^{2+}(aq) + CO_3^{2-}(aq) \rightarrow MgCO_3(s)$ $Pb^{2+}(aq) + CO_3^{2-}(aq) \rightarrow PbCO_3(s)$ $Zn^{2+}(aq) + CO_3^{2-}(aq) \rightarrow ZnCO_3(s)$

Note

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(i)Iron(III)carbonate(IV) does not exist.
(ii)Copper(II)Carbonate(IV) exist only as the basic $CuCO_3.Cu(OH)_2$
(iii)Both $BaCO_3$ and $BaSO_3$ are insoluble white precipitate. If hydrochloric acid is added to the white precipitate;

I. $BaCO_3$ produces CO_2 gas. When bubbled/directed into lime water solution,a white precipitate is formed.
II. I. $BaSO_3$ produces SO_2 gas. When bubbled/directed into orange acidified potassium dichromate(VI) solution, it turns to green/decolorizes acidified potassium manganate(VII).

(f) Reaction of cation with sulphide / S^{2-} ions

All sulphides are insoluble **black** solids/precipitates **except** sodium sulphide/ Na_2S/ potassium sulphide/K_2S.When a few/3drops of the soluble sulphide is added to a metal cation/salt, a black precipitate is formed.

Procedure

Place about 2cm3 of $Cu(NO_3)_2$, $FeSO_4$, $MgCl_2$,$Pb(NO_3)_2$ and $ZnSO_4$ in separate boiling tubes.
Add six drops of Potassium /sodium sulphide solution.

Observation	Inference	Explanation
Black ppt	S^{2-} ions	CuS, FeS,MgS,PbS, ZnS are black insoluble precipitates $Cu^{2+}(aq) + S^{2-}(aq) \rightarrow$ **CuS**(s) $Pb^{2+}(aq) + S^{2-}(aq) \rightarrow$ **PbS**(s) $Fe^{2+}(aq) + S^{2-}(aq) \rightarrow$ **FeS**(s) $Zn^{2+}(aq) + S^{2-}(aq) \rightarrow$ **ZnS**(s)

Sample qualitative analysis guide

You are provided with solid Y(aluminium (III)sulphate(VI)hexahydrate).Carry out the following tests and record your observations and inferences in the space provided.

1(a) Appearance

Observations (1mark)	inference
White crystalline solid	Coloured ions Cu^{2+} , Fe^{2+} ,Fe^{3+} absent

(b)Place about a half spatula full of the solid into a clean dry boiling tube. Heat gently then strongly.

Observations (1mark)	inference
Colourless droplets formed on the cooler part of the test tube	Hydrated compound/compound containing water of crystallization

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Solid remains a white residue

(c)Place all the remaining portion of the solid in a test tube .Add about 10cm3 of distilled water. Shake thoroughly. Divide the mixture into five portions.

Observation	Inference (1mark)
Solid dissolves to form a colourless solution	Polar soluble compound Cu^{2+} , Fe^{2+} ,Fe^{3+} absent

(i)To the first portion, add three drops of sodium hydroxide then add excess of the alkali.

Observation	Inference (1mark)
White ppt, soluble in excess	Zn^{2+} , Pb^{2+} , Al^{3+}

(ii)To the second portion, add three drops of aqueous ammonia then add excess of the alkali.

Observation	Inference (1mark)
White ppt, insoluble in excess	Pb^{2+} , Al^{3+}

(iii)To the third portion, add three drops of sodium sulphate(VI)solution.

Observation	Inference (1mark)
No white ppt	Al^{3+}

(iv)I.To the fourth portion, add three drops of Lead(II)nitrate(IV)solution. **Preserve**

Observation	Inference (1mark)
White ppt	CO_3^{2-}, SO_4^{2-}, SO_3^{2-}, Cl^-,

II.To the portion in (iv) I above , add five drops of dilute hydrochloric acid.

Observation	Inference (1mark)
White ppt persist/remains	SO_4^{2-}, Cl^-,

III.To the portion in (iv) II above, heat to boil.

Observation	Inference (1mark)
White ppt persist/remains	SO_4^{2-},

Note that:

(i)From test above, it can be deduced that solid Y is hydrated aluminium(III)sulphate(VI) solid

(ii)Any ion inferred from an observation **below** must be derived from previous correct observation and inferences **above**. e.g.

Al^{3+} in c(iii) must be correctly inferred in either/or in c(ii) or c(i)**above**

SO_4^{2-} in c(iv)III must be correctly inferred in either/or in c(iv)II or c(iv)I **above**

(iii)Contradiction in observations and inferences should be avoided.e.g.

"White ppt soluble in excess" to infer presence of Al^{3+} ,$\mathbf{Ba^{2+}}$,Pb^{3+}

(iv)Symbols of elements/ions should be correctly capitalized. e.g.

"SO_4^{-2}" is wrong, "sO_4^{2-}" is wrong, "cu^{2+}" is wrong.

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Sample solutions of salt were labeled as I,II, III and IV. The actual solutions, not in that order are lead nitrate, zinc sulphate potassium chloride and calcium chloride.
a)When aqueous sodium carbonate was added to each sample separately, a white precipitate was formed in I, III and IV only. Identify solution II.
b)When excess sodium hydroxide was added to each sample separately, a white precipitate was formed in solutions III and I only.
Identify solution I

17.When solids/salts /solutes are added to a solvent ,some dissolve to form a solution.

Solute + Solvent -> Solvent

If a solution has **a lot** of solute dissolved in a solvent ,it is said to be **concentrated**.
If a solution has **little** solute dissolved in a solvent ,it is said to be **dilute**.

There is a limit to how much solute can dissolve in a given /specified amount of solvent/water at a given /specified temperature.
The maximum mass of salt/solid/solute that dissolve in 100g of solvent/water at a specified temperature is called **solubility of a salt.**
When **no more solute can dissolve in a given amount of solvent at a specified temperature**, a saturated solution is formed.
For some salts, on heating, more of the salt/solid/solute dissolve in the saturated solution to form a super saturated solution.
The solubility of a salt is thus calculated from the formula

$$\textbf{Solubility} = \frac{\textbf{Mass of solute/salt/solid}}{\textbf{Mass/volume of water/solvent}} \times 100$$

Practice examples

(a)Calculate the solubility of potassium nitrate(V) if 5.0 g of the salt is dissolved in 50.0cm3 of water.

$$\text{Solubility} = \frac{\text{Mass of solute/salt/solid}}{\text{Mass/volume of water/solvent}} \times 100 \Rightarrow \frac{5.0}{50.0} \times 100 = \textbf{10.0 g /100g } H_2O$$

(b)Calculate the solubility of potassium chlorate(V) if 50.0 g of the salt is dissolved in 250.0cm3 of water.

$$\text{Solubility} = \frac{\text{Mass of solute/salt/solid}}{\text{Mass/volume of water/solvent}} \times 100 \Rightarrow \frac{50.0}{250.0} \times 100 = \textbf{20.0 g /100g } H_2O$$

(c)If the solubility of potassium chlorate(V) is 5g/100g H_2O at 80°C,how much can dissolve in 5cm3 of water at 80°C .

$$\text{Mass of solute/salt/solid} = \frac{\text{Solubility}}{100} \times \text{Mass/volume of water/solvent}$$

$$\Rightarrow \frac{5}{100} \times 5 = \textbf{0.25g} \text{ of } KClO_3 \text{ dissolve}$$

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(d)If the solubility of potassium chlorate(V) is 72g/100g H_2O at 20°C,how much can saturate 25g of water at 20°C .

Mass of solute/salt/solid = Solubility x Mass/volume of water/solvent
100

=> 72 x 25 = **18.0g** of $KClO_3$ dissolve/saturate
100

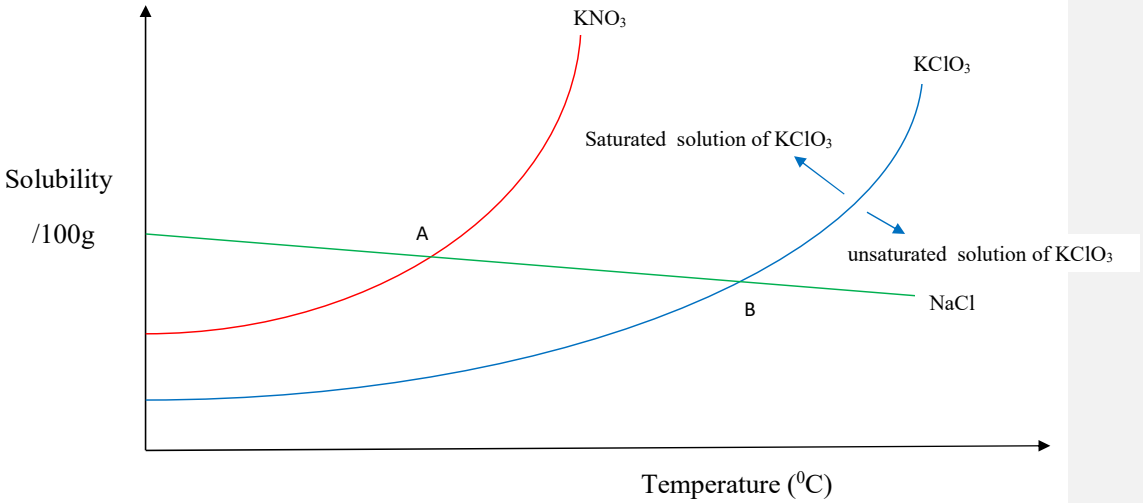
(e) 22g of potassium nitrate(V) was dissolved in 40.0g of water at 10°C. Calculate the solubility of potassium nitrate(V) at 10°C.

Solubility = Mass of solute/salt/solid x 100 =>(22 x 100) = **55.0 g /100g H_2O**
Mass/volume of water/solvent 40.0.

(f)What volume of water should be added to 22.0g of water at 10°C if the solubility of KNO_3 at 10°C is 5.0g/100g H_2O?

Solubility is mass/100g H_2O => 22.0g + x = 100cm3/100g H_2O
X= 100 – 22 = **78** cm3 of H_2O

18. A graph of solubility against temperature is called solubility curve.
It shows the influence of temperature on solubility of different substances/solids/salts.
Some substances dissolve more with increase in temperature while for others dissolve less with increase in temperature



Note:
(i)solubility of KNO_3 and $KClO_3$ increase with increase in temperature.
(ii)solubility of KNO_3 is always higher than that of $KClO_3$ at any specified temperature.
(iii)solubility of NaCl decrease with increase in temperature.

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(iv)NaCl has the highest solubility at low temperature while $KClO_3$ has the lowest solubility at low temperature.
(v)At point A both NaCl and KNO_3 are equally soluble.
(vi)At point B both NaCl and $KClO_3$ are equally soluble.
(vii) An area above the solubility curve of the salt shows a **saturated** /supersaturated solution.
(viii) An area below the solubility curve of the salt shows an **unsaturated** solution.
19.(a) For salts whose solubility **increases** with increase in temperature, crystals form when the salt solution at **higher** temperatures is **cooled** to a lower temperature.

(c) For salts whose solubility **decreases** with increase in temperature, crystals form when the salt solution at **lower** temperatures is **heated** to a higher temperature.

The examples below shows determination of the mass of crystals deposited with changes in temperature.

1.The solubility of $KClO_3$ at 100°C is 60g/100g water .What mass of $KClO_3$ will be deposited at:

(i)75 °C if the solubility is now 39g/100g water.

At 100°C = 60.0g
Less at 75°C = - 39.0g
Mass of crystallized out **21.0g**

(i)35 °C if the solubility is now 28 g/100g water.

At 100°C = 60.0g
Less at 35°C = - 28.0.0g
Mass of crystallized out **32.0g**

2. KNO_3 has a solubility of 42 g/100g water at 20°C.The salt was heated and added 38g more of the solute which dissolved at100°C. Calculate the solubility of KNO_3 at 100°C.

Solubility of KNO_3 at 100°C = solubility at 20°C + mass of KNO_3 added
=> 42g + 38g = **80g KNO_3 /100g H_2O**

3. A salt solution has a mass of 65g containing 5g of solute. The solubility of this salt is 25g per 100g water at 20°C. 60g of the salt are added to the solution at 20°C.Calculate the mass of the solute that remain undissolved.

Mass of solvent at 20°C = mass of solution – mass of solute
=> 65 - 5 = **60g**

Solubility before adding salt = $\frac{\text{mass of solute}}{\text{Volume of solvent}} \times 100$
=> $\frac{5}{60} \times 100$ = **8.3333g/100g** water

Mass of solute to equalize with solubility = 25 – 8.3333g = **16.6667g**
Mass of solute undissolved = 60.0 - 16.6667g = **43.3333** g

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4. Study the table below

Salt	Solubility in gram at	
	50°C	20°C
KNO_3	90	30
$KClO_3$	20	6

(i)What happens when the two salts are dissolved in water then cooled from 50°C to 20°C.

(90 – 30) = 60.0 g of KNO_3 crystals precipitate

(20 – 6) = 14.0 g of $KClO_3$ crystals precipitate

(ii)State the assumption made in (i) above.

Solubility of one salt has no effect on the solubility of the other.

5. 10.0 g of hydrated potassium carbonate (IV) $K_2CO_3.xH_2O$ on heating leave 7.93 of the hydrate.

(a)Calculate the mass of anhydrous salt obtained.

Hydrated on heating leave anhydrous = **7.93** g

(b)Calculate the mass of water of crystallization in the hydrated salt

Mass of water of crystallization = hydrated – anhydrous

=> 10.0 - 7.93 = **2.07 g**

(c)How many moles of anhydrous salt are there in 10of hydrate? (K= 39.0,C=12.0.O= 16.0)

Molar mass K_2CO_3= 138

Moles K_2CO_3 = mass of K_2CO_3 => 7.93 = **0.0515 moles**

Molar mass K_2CO_3 138

(d)How many moles of water are present in the hydrate for every one mole of K_2CO_3 ? (H=1.0.O= 16.0)

Molar mass H_2O = 18

Moles H_2O = mass of H_2O => 2.07 = **0.115 moles**

Molar mass H_2O 18

Mole ratio H_2O : K_2CO_3 = 0.115 moles 2 = **2**

0.0515 moles 1

(e)What is the formula of the hydrated salt?

K_2CO_3 .2 H_2O

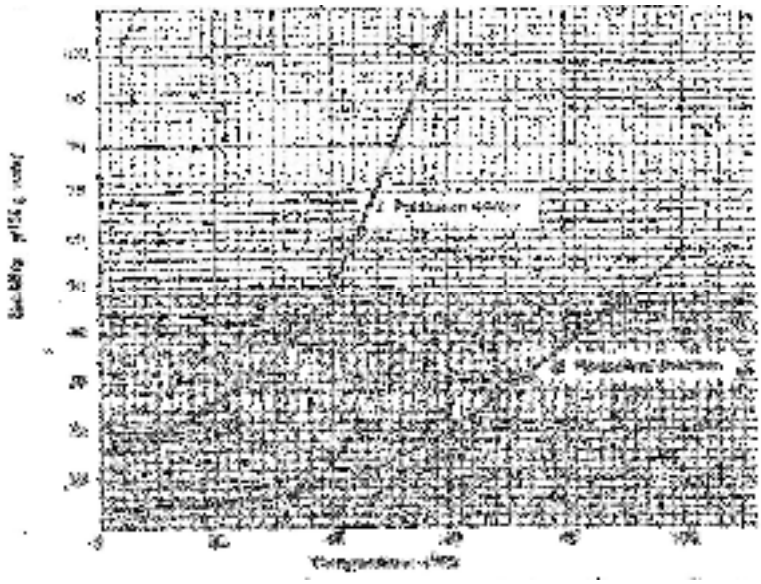
6. The table below shows the solubility of Potassium nitrate(V) at different temperatures.

Temperature(°C)	5.0	10.0	15.0	30.0	40.0	50.0	60.0
mass KNO_3/ 100g water	15.0	20.0	25.0	50.0	65.0	90.0	120.0

(a)Plot a graph of mass of in 100g water(y-axis) against temperature in °C

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(b)From the graph show and determine

(i)the mass of KNO_3 dissolved at:

I. 20°C

From a correctly plotted graph = **32g**

II. 35°C

From a correctly plotted graph = **57g**

III. 55°C

From a correctly plotted graph = **104g**

(ii)the temperature at which the following mass of KNO_3 dissolved:

I. 22g

From a correctly plotted graph =**13.0°C**

II. 30g

From a correctly plotted graph =**17.5°C**

III.100g

From a correctly plotted graph =**54.5°C**

(c)Explain the shape of your graph.

Solubility of KNO_3 increase with increase in temperature/More KNO_3 dissolve as temperature rises.

(d)Show on the graph the supersaturated and unsaturated solutions.

Above the solubility curve write; "supersaturated"

Below the solubility curve write; "unsaturated"

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(e)From your graph, calculate the amount of crystals obtained when a saturated solution of KNO_3 containing 180g of the salt is cooled from 80°C to:

I. 20°C

Solubility before heating = 180 g
Less Solubility after heating(from the graph) = 32 g
Mass of KNO_3crystals = **148 g**

II. 35°C

Solubility before heating = 180 g
Less Solubility after heating(from the graph) = 58 g
Mass of KNO_3crystals = **122 g**

III. 55°C

Solubility before heating = 180 g
Less Solubility after heating(from the graph) = 102 g
Mass of KNO_3crystals = **78 g**

7. The table below shows the solubility of salts A and B at various temperatures.

Temperature(°C)	0.0	10.0	20.0	30.0	40.0	50.0	60.0	70.0	80.0
Solubility of A	28.0	31.0	34.0	37.0	40.0	43.0	45.0	48.0	51.0
Solubility of B	13.0	21.0	32.0	46.0	64.0	85.0	110.0	138.0	169.0

(a)On the same axis plot a graph of solubility (y-axis) against temperature for each salt.

(b)At what temperature are the two salts equally soluble.

The point of intersection of the two curves = 24°C

(c)What happens when a mixture of 100g of salt B with 100g if water is heated to 80°C

From the graph, the solubility of B at 80°C is 169g /100g water. All the 100g crystals of B dissolve.

(d)What happens when the mixture in (c) above is then cooled from 50°C to 20°C.

Method I.

Total mass before cooling at 50°C = 100.0 g
(From graph) Solubility/mass after cooling at 20°C = 32.0 g
Mass of crystals deposited **68.0 g**

Method II.

Mass of soluble salt crystals at 50°C added = 100 g
(From graph)Solubility/mass before cooling at 50°C = 85.0 g

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Mass of crystals that cannot dissolve at 50°C **15.0** g
(From graph) Solubility/mass before cooling at 50°C = 85.0 g
(From graph) Solubility/mass after cooling at 20°C = 32.0 g
Mass of crystals deposited after cooling **53.0 g**
Total mass of crystals deposited = 15.0 + 53.0 = **68.0 g**

(e)A mixture of 40g of A and 60g of B is added to 10g of water and heated to 70°C.The solution is then allowed to cool to 10°C.Describe clearly what happens.

I.For salt A

Solubility of A before heating = mass of A x 100
Volume of water added

=> 40 x 100 = **400g/100g Water**
10

(Theoretical)Solubility of A before heating = 400 g
Less (From graph) Solubility of A after heating at 70°C = 48g
Mass of crystals that can not dissolve at70°C = **352 g**
(From graph) Solubility of A after heating at 70°C = 48g
Less (From graph) Solubility of A after cooling to 10°C = 31g
Mass of crystals that crystallize out on cooling to10°C = **17 g**
Mass of crystals that can not dissolve at70°C = **352 g**
Add Mass of crystals that crystallize out on cooling to10°C = **17 g**
Total mass of A that does not dissolve/crystallize/precipitate = 369 g

I.For salt B

Solubility of B before heating = mass of B x 100
Volume of water added

=> 60 x 100 = **600g/100g Water**
10

(Theoretical)Solubility of B before heating = 600 g
Less (From graph) Solubility of B after heating at 70°C = 138g
Mass of crystals that cannot dissolve at70°C = **462 g**
(From graph) Solubility of B after heating at 70°C = 138g
Less (From graph) Solubility of B after cooling to 10°C = 21g
Mass of crystals that crystallize out on cooling to10°C = **117 g**
Mass of crystals that cannot dissolve at70°C = **462 g**
Add Mass of crystals that crystallize out on cooling to10°C = **117 g**
Total mass of A that does not dissolve/crystallize/precipitate = 579 g

(f)State the assumption made in (e)above

Solubility of one salt has no effect on the solubility of the other

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8. When 5.0 g of potassium chlorate (V) was put in 10cm3 of water and heated, the solid dissolves. When the solution was cooled , the temperature at which crystals reappear was noted. Another 10cm3 of water was added and the mixture heated to dissolve then cooled for the crystals to reappear .The table below shows the the results obtained

Total volume of water added(cm3)	10.0	20.0	30.0	40.0	50.0
Mass of $KClO_3$	5.0	5.0	5.0	5.0	5.0
Temperature at which crystals appear	80.0	65.0	55.0	45.0	30.0
Solubility of $KclO_3$	**50.0**	**25.0**	**16.6667**	**12.5**	**10.0**

(a)Complete the table to show the solubility of $KclO_3$ at different temperatures.

(b)Plot a graph of mass of $KClO_3$ per 100g water against temperature at which crystals form.

(c)From the graph, show and determine ;

(i)the solubility of $KClO_3$ at

I. 50°C

From a well plotted graph = 14.5 g $KClO_3$/100g water

II. 35°C

From a well plotted graph = 9.0 g $KclO_3$/100g water

(ii)the temperature at which the solubility is:

I.10g/100g water

From a well plotted graph = 38.0 °C

II.45g/100g water

From a well plotted graph = 77.5 °C

(d)Explain the shape of the graph.

Solubility of $KClO_3$ increase with increase in temperature/more $KclO_3$dissolve as temperature rises.

(e)What happens when 100g per 100g water is cooled to 35.0 °C

Solubility before heating = 100.0

(From the graph) Solubility after cooling = 9.0

Mass of salt precipitated/crystallization = 91.0 g

9. 25.0cm3 of water dissolved various masses of ammonium chloride crystals at different temperatures as shown in the table below.

Mass of ammonium chloride(grams)	4.0	4.5	5.5	6.5	9.0
Temperature at which solid dissolved(°C)	30.0	50.0	70.0	90.0	120.0
Solubility of NH_4Cl	**16.0**	**18.0**	**22.0**	**26.0**	**36.0**

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(a)Complete the table
(b)Plot a solubility curve
(c)What happens when a saturated solution of ammonium chloride is cooled from 80°C to 40°C.

(From the graph)Solubility at 80°C = 24.0 g
Less (From the graph)Solubility at 40°C = 16.8 g
Mass of crystallized/precipitated = 7.2 g

20. Solubility and solubility curves are therefore **used**
(i) to know the effect of temperature on the solubility of a salt
(ii)to fractional crystallize two soluble salts by applying their differences in solubility at different temperatures.
(iii)determine the mass of crystal that is obtained from crystallization.

21.Natural fractional crystallization takes place in Kenya/East Africa at:
(i) Lake Magadi during extraction of soda ash(Sodium carbonate) from Trona(sodium sesquicarbonate)
(ii) Ngomeni near Malindi at the Indian Ocean Coastline during the extraction of common salt(sodium chloride).

22.**Extraction of soda ash from Lake Magadi in Kenya**
Rain water drains underground in the great rift valley and percolate underground where it is heated geothermically.
The hot water dissolves underground soluble sodium compounds and comes out on the surface as alkaline springs which are found around the edges of Lake Magadi in Kenya.
Temperatures around the lake are very high (30-40°C) during the day.
The solubility of trona decrease with increase in temperature therefore solid crystals of trona grows on top of the lake (upto or more than 30metres thick)

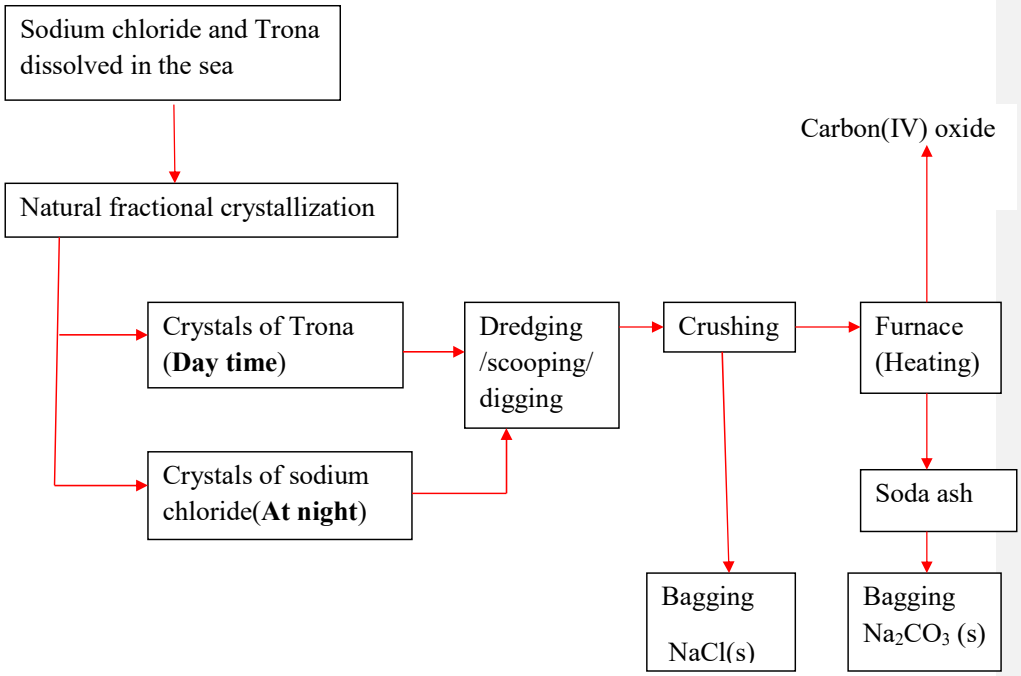
A bucket dredger mines the trona which is then crushed ,mixed with lake liquor and pumped to washery plant where it is further refined to a green granular product called CRS.
The CRS is then heated to chemically decompose trona to soda ash(Sodium carbonate)
Chemical equation
$2Na_2CO_3.NaHCO_3.2H_2O(s) \rightarrow 3Na_2CO_3(s) + CO_2(g) + 5H_2O(l)$
Soda ash(Sodium carbonate) is then stored .It is called Magadi Soda. Magadi Soda is used :
(i) make glass
(ii) for making soapless detergents
(iii) softening hard water.
(iv)

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Common salt is colledcted at night because its solubility decreases with decrease in temperature. It is used as salt lick/feed for animals.

Summary flow diagram showing the extraction of Soda ash from Trona



23.Extraction of common salt from Indian Ocean at Ngomeni in Kenya

Oceans are salty.They contain a variety of dissolved salts (about 77% being sodium chloride).

During high tide ,water is collected into shallow pods and allowed to crystallize as evaporation takes place.The pods are constructed in series to increase the rate of evaporation.

At the final pod ,the crystals are scapped together,piled in a heap and washed with brine (concentrated sodium chloride).

It contains **$MgCl_2$** and **$CaCl_2$** . $MgCl_2$ and $CaCl_2$are **hygroscopic**. They absorb water from the atmosphere and form a solution.

This makes table salt damp/wet on exposure to the atmosphere.

24.Some water form lather easily with soap while others do not.

Water which form lather easily with soap is said to be "**soft**"

Water which do not form lather easily with soap is said to be "**hard**"

Hardness of water is caused by the presence of Ca^{2+} and Mg^{2+} ions.

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Ca^{2+} and Mg^{2+} ions react with soap to form an **insoluble** grey /white suspension/precipitate called **Scum/ curd.** Ca^{2+} and Mg^{2+} ions in water come from the water sources passing through rocks containing soluble salts of Ca^{2+} and Mg^{2+} e.g. Limestone or gypsum

There are two types of water hardness:

(a)temporary hardness of water

(b)permanent hardness of water

(a)temporary hardness of water

Temporary hardness of water is caused by the presence of dissolved calcium hydrogen carbonate/**$Ca(HCO_3)_2$** and magnesium hydrogen carbonate/**$Mg(HCO_3)_2$** When rain water dissolve carbon(IV) oxide from the air it forms waek carbonic(IV) acid i.e.

$$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$$

When carbonic(IV) acid passes through limestone/dolomite rocks it reacts to form soluble salts i.e.

In limestone areas; $H_2CO_3(aq) + CaCO_3(s) \rightarrow Ca(HCO_3)_2(aq)$

In dolomite areas; $H_2CO_3(aq) + MgCO_3(s) \rightarrow Mg(HCO_3)_2(aq)$

(b)permanent hardness of water

Permanent hardness of water is caused by the presence of dissolved calcium sulphate(VI)/**$CaSO_4$** and magnesium sulphate(VI)/**$Mg\ SO_4$** Permanent hardness of water is caused by water dissolving **$CaSO_4$** and **$MgSO_4$** from ground rocks.

Hardness of water can be removed by the following methods:

(a) Removing temporary hardness of water

(i)Boiling/heating.

Boiling decomposes insoluble calcium hydrogen carbonate/**$Ca(HCO_3)_2$** and magnesium hydrogen carbonate/**$Mg(HCO_3)_2$** to insoluble **$CaCO_3$** and **$MgCO_3$** that precipitate away. i.e

Chemical equation

$$Ca(HCO_3)_2(\mathbf{aq}) \rightarrow CaCO_3(\mathbf{s}) + CO_2(g) + H_2O(l)$$

$$Mg(HCO_3)_2(\mathbf{aq}) \rightarrow MgCO_3(\mathbf{s}) + CO_2(g) + H_2O(l)$$

(ii)Adding sodium carbonate (IV) /Washing soda.

Since boiling is expensive on a large scale ,a calculated amount of sodium carbonate decahydrate /**$Na_2CO_3.10H_2O$** precipitates insoluble $Ca^{2+}(aq)$ and $Mg^{2+}(aq)$ ions as carbonates to remove **both** temporary and permanent hardness of water .This a double decomposition reaction where **two soluble** salts form an **insoluble** and **soluble** salt. i.e.

(i)with temporary hard water

Chemical equation

$$\mathbf{Na_2CO_3\ (aq) + Ca(HCO_3)_2\ (aq) \rightarrow NaHCO_3(aq) + CaCO_3\ (s)}$$

$$\mathbf{Na_2CO_3\ (aq) + Mg(HCO_3)_2\ (aq) \rightarrow NaHCO_3(aq) + MgCO_3\ (s)}$$

Ionic equation

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CO_3^{2-} (aq) + Ca^{2+} (aq) -> $CaCO_3$ (s)
CO_3^{2-} (aq) + Mg^{2+} (aq) -> $MgCO_3$ (s)

(ii)with permanent hard water

Chemical equation

Na_2CO_3 (aq) + $MgSO_4$ (aq) -> Na_2SO_4 (aq) + $MgCO_3$ (s)
Na_2CO_3 (aq) + $CaSO_4$ (aq) -> Na_2SO_4 (aq) + $MgCO_3$ (s)

Ionic equation

CO_3^{2-} (aq) + Ca^{2+} (aq) -> $CaCO_3$ (s)
CO_3^{2-} (aq) + Mg^{2+} (aq) -> $MgCO_3$ (s)

(iii)Adding calcium (II)hydroxide/Lime water

Lime water/calcium hydroxide removes only temporary hardness of water from by precipitating insoluble calcium carbonate(IV).

Chemical equation

$Ca(OH)_2$ (aq) + $Ca(HCO_3)_2$ (aq) -> $2H_2O$(l) + $2CaCO_3$ (s)

Excess of Lime water/calcium hydroxide should not be used because it dissolves again to form soluble calcium hydrogen carbonate(IV) causing the hardness again.

(iv)Adding aqueous ammonia

Aqueous ammonia removes temporary hardness of water by precipitating insoluble calcium carbonate(IV) and magnesium carbonate(IV)

Chemical equation

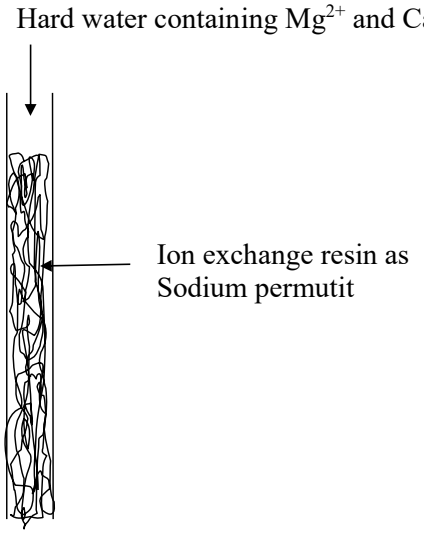
$2NH_3$ (aq) + $Ca(HCO_3)_2$ (aq) -> $(NH_4)_2CO_3$(aq) + $CaCO_3$ (s)
$2NH_3$ (aq) + $Mg(HCO_3)_2$ (aq) -> $(NH_4)_2CO_3$(aq) + $MgCO_3$ (s)

(v)Use of ion-exchange permutit

This method involves packing a chamber with a resin made of insoluble complex of sodium salt called **sodium permutit**.

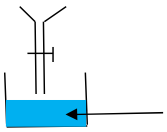
The sodium permutit releases sodium ions that are exchanged with Mg^{2+} and Ca^{2+} ions in hard water making the water to be soft. i.e.

Na_2**X**(aq) + Ca^{2+} (aq) -> Na^+ (aq) + Ca**X**(s)
Na_2**X**(aq) + Mg^{2+} (aq) -> Na^+ (aq) + Mg**X**(s)



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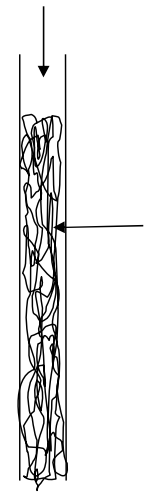


Na^+ ions replace Mg^{2+} and Ca^{2+} to make the water soft.

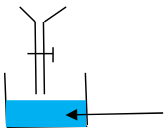
When all the Na+ ions in the resin is fully exchanged with Ca2+ and Ng2+ ions in the permutit column ,it is said to be **exhausted**.

Brine /concentrated sodium chloride solution is passed through the permutit column to regenerated /recharge the column again.

Hard water containing Mg^{2+} and Ca^{2+}



Ion exchange resin as Sodium permutit



Na^+ ions replace Mg^{2+} and Ca^{2+} to make the water soft.

(vi)Deionization /demineralization

This is an advanced ion exchange method of producing deionized water .Deionized water is extremely pure water made only of hydrogen and oxygen only without any dissolved substances.

Deionization involve using the resins that remove all the cations by using:

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(i)A cation exchanger which remove /absorb all the **cations** present in water and leave **only H^+** ions.

(ii)An anion exchanger which remove /absorb all the **anions** present in water and leave **only OH^-** ions.

The $H^+(aq)$ and $OH^-(aq)$ neutralize each other to form pure water.

Chemical equation

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

When exhausted the cation exchanger is regenerated by adding $H^+(aq)$ from sulphuric(VI)acid/hydrochloric acid.

When exhausted the anion exchanger is regenerated by adding $OH^-(aq)$ from sodium hydroxide.

Advantages of hard water

Hard water has the following advantages:

(i)$Ca^{2+}(aq)$ in hard water are useful in bone and teeth formation

(ii) is good for brewing beer

(iii)contains minerals that cause it to have better /sweet taste

(iv)animals like snails and coral polyps use calcium to make their shells and coral reefs respectively.

(v)processing mineral water

Disadvantages of hard water

Hardness of water:

(i)waste a lot of soap during washing before lather is formed.

(ii)causes stains/blemishes/marks on clothes/garments

(iii)causes fur on electric appliances like kettle ,boilers and pipes form decomposition of carbonates on heating .This reduces their efficiency hence more/higher cost of power/electricity.

Sample revision questions

In an experiment, soap solution was added to three separate samples of water. The table below shows the volumes of soap solution required to form lather with 1000cm3 of each sample of water before and after boiling.

	Sample I	Sample II	Sample III
Volume of soap before water is boiled (cm3)	27.0	3.0	10.0
Volume of soap after water is boiled(cm3)	27.0	3.0	3.0

a) Which water sample is likely to be soft? Explain. (2mks)

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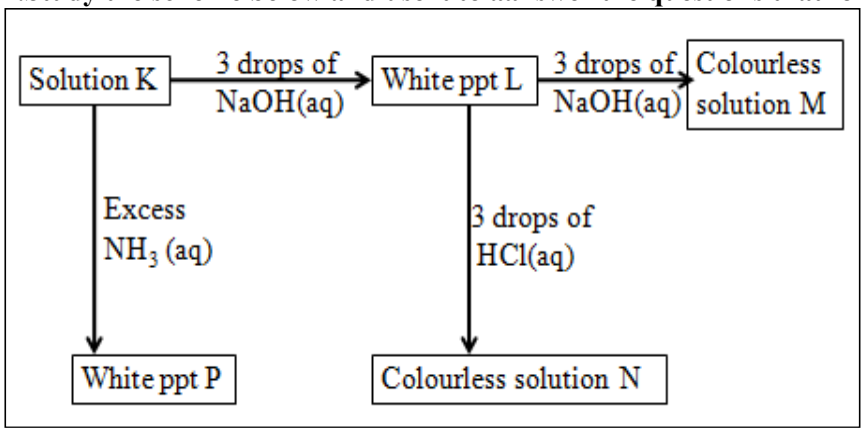
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Sample II: Uses little sample of soap .

c) Name the change in the volume of soap solution used in sample III (1mk)

On heating the sample water become soft bcause it is temporary hard.

2.Study the scheme below and use it to aanswer the questions that follow:



(a)Write the formula of:

(i)Cation in solution K

Al^{3+}

(ii)white ppt L

$Al(OH)_3$

(iii) colourless solution M

$[Al(OH)_4]^-$

(iv) colourless solution N

$AlCl_3$

(v)white ppt P

$Al(OH)_3$

(b)Write the ionic equation for the reaction for the formation of:

(i)white ppt L

$Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$

(v)white ppt P

$Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$

(c)What property is illustrated in the formation of colourless solution M and N.

Amphotellic

BASIC PRINCIPLES OF CHEMISTRY PRACTICALS

Introduction/Rationale

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Chemistry is a science.
Chemistry practical all over the world is emphasized to all candidates sitting for a Chemistry paper.
There are about seven main basic universal emphasis for all chemistry candidates sitting for a chemistry paper;

(i)Titration /volumetric analysis
(ii)Thermochemistry(energy changes)
(iii)Chemical kinetic(rates of reaction)
(iv)Qualitative analysis(organic/inorganic)
(v)Solubility and solubility curves
(vi)Flame test
(vii)Physical / general chemistry

(a) **Titration/volumetric analysis**

Titration is determining the end point of the burette contents that react with fixed (usually 25.0cm3 from a pipette) conical flask contents.
As **evidence** of a titration **actually** done examining body require the candidate to record their burette readings before and after the titration.
For KCSE candidates burette readings **must** be recorded in a titration table in the **format provided** by the Kenya National Examination Council.
As **evidence** of all titration **actually** done Kenya National Examination Council require the candidate to record their burette readings before and after the titration to complete the titration table **in the format provided**.

Sample Titration table format

Final burette reading (cm3)	24.0	24.0	24.0
Initial burette reading (cm3	0.0	0.0	0.0
Volume of solution used(cm3)	24.0	24.0	24.0

Calculate the average volume of solution used

$$\frac{24.0 + 24.0 + 24.0}{3} = \textbf{24.0 cm3}$$

As evidence of understanding the degree of accuracy of burettes ,all readings must be recorded to **a** decimal point.
As evidence of accuracy in carrying the out the titration ,candidates value should be **within 0.2** of the **school value** .
The school value is the **teachers** readings presented to the examining body/council based on the concentrations of the solutions s/he presented to her/his candidates.
Bonus mark is awarded for averaged reading **within 0.1** school value as Final answer.
Calculations involved after the titration require candidates **thorough practice mastery** on the:

(i)relationship among the mole, molar mass, mole ratios, concentration, molarity.

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(ii) mathematical application of 1st principles.
Very useful information which candidates forget appear usually in the beginning of the paper as:
"You are provided with…"
All calculation must be to the **4th decimal point** unless they divide fully to a lesser decimal point.
Never round off answers.
b)Thermochemistry/energy changes
Energy is the capacity to do work which is measured in Joules(**J**) or(k**J**) .
Chemical/physical changes take place with **absorption** (Endothermic) or **evolution/ production** (Exothermic)of heat.
Practically:
(i)endothermic changes show absorption of heat by a fall / drop in temperature and has a **+ΔH**
(ii)exothermic changes show evolution/ production of heat by a rise in temperature and has a **-ΔH**
(iii)temperature is measure using a **thermometer.**
(iv)a school thermometer is either coloured (alcohol) or colourless(mercury)
(v) For accuracy ,candidates in the same practical session should use the same type of thermometer.
(vi) fall / drop **(+ΔH)** in temperature is movement of thermometer **level downward.**
(vii) rise **(-ΔH)** in temperature is movement of thermometer **level upwards.**
Physical changes changes mainly involve melting/freezing/fussion and boiling /vapourization.
Chemical changes changes mainly involve displacement ,dissolving , neutralization
a).Energy changes in physical processes
Melting/freezing/fusion/solidification and **boiling/vaporization/evaporation** are the two physical processes.
Melting /freezing point of pure substances is fixed /constant.
The boiling point of pure substance depends on **external** atmospheric **pressure**.
Melting/fusion is the physical change of a **solid** to **liquid**. Freezing/fusion is the physical change of a **liquid** to **solid**.
Melting/freezing/fusion/solidification are therefore two **opposite** but **same** reversible physical processes. i.e

A (**s**) ========A(**l**)

Boiling/vaporization/evaporation is the physical change of a **liquid** to **gas/vapour**.
Condensation/liquidification is the physical change of **gas/vapour** to **liquid**.
Boiling/vaporization/evaporation and condensation/liquidification are therefore two **opposite** but **same** reversible physical processes. i.e

B (**l**) ========B(**g**)

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Practically

(i) Melting/liquidification/fusion involves **heating** a solid to **weaken** the strong bonds holding the solid particles together.

Solids are made up of very strong bonds holding the particles **very close** to each other (**Kinetic Theory of matter**

On heating these particles gain energy/heat from the surrounding heat source to form a liquid with **weaker** bonds holding the particles close together but with some degree of **freedom**.

Melting/fusion is an **endothermic** (+ΔH)process that require/absorb energy from the surrounding.

(ii)Freezing/fusion/solidification involves cooling a a liquid to reform /rejoin the very strong bonds to hold the particles **very close** to each other as solid and thus lose their degree of **freedom** (**Kinetic Theory of matter**).

Freezing /fusion / solidification is an **exothermic** (**-**ΔH)process that require particles holding the liquid together to lose energy to the surrounding.

(iii)Boiling/vaporization/evaporation involves **heating** a liquid to completely **break/free** the bonds holding the liquid particles together.

Gaseous particles have high degree of **freedom** (**Kinetic Theory of matter**).

Boiling /vaporization / evaporation is an **endothermic** (+ΔH) process that require/absorb energy from the surrounding.

(iv)Condensation/liquidification is **reverse** process of boiling /vaporization / evaporation.

It involves gaseous particles losing energy to the surrounding to form a liquid.It is an **exothermic**(+ΔH) process.

The quantity of energy required to **change** one mole of a solid **to** liquid or to **form** one mole of a solid **from** liquid at constant temperature is called **molar enthalpy/latent heat of fusion**. e.g.

H_2O(**s**) -> H_2O(**l**) ΔH = **+6.0kJ mole^{-1}** (endothermic process)

H_2O(**l**) -> H_2O(**s**) ΔH = **-6.0kJ mole^{-1}** (exothermic process)

The quantity of energy required to **change** one mole of a liquid **to** gas/vapour or to **form** one mole of a liquid **from** gas/vapour at constant temperature is called **molar enthalpy/latent heat of vapourization**. e.g.

H_2O(**l**) -> H_2O(**g**) ΔH = **+44.0kJ mole^{-1}** (endothermic process)

H_2O(**g**) -> H_2O(**l**) ΔH = **-44.0kJ mole^{-1}** (exothermic process)

- **To determine the boiling point of water**

Procedure:

Measure 20cm3 of tap water into a 50cm3 glass beaker. Determine and record its temperature.Heat the water on a strong Bunsen burner flame and record its temperature after every thirty seconds for four minute

Sample results

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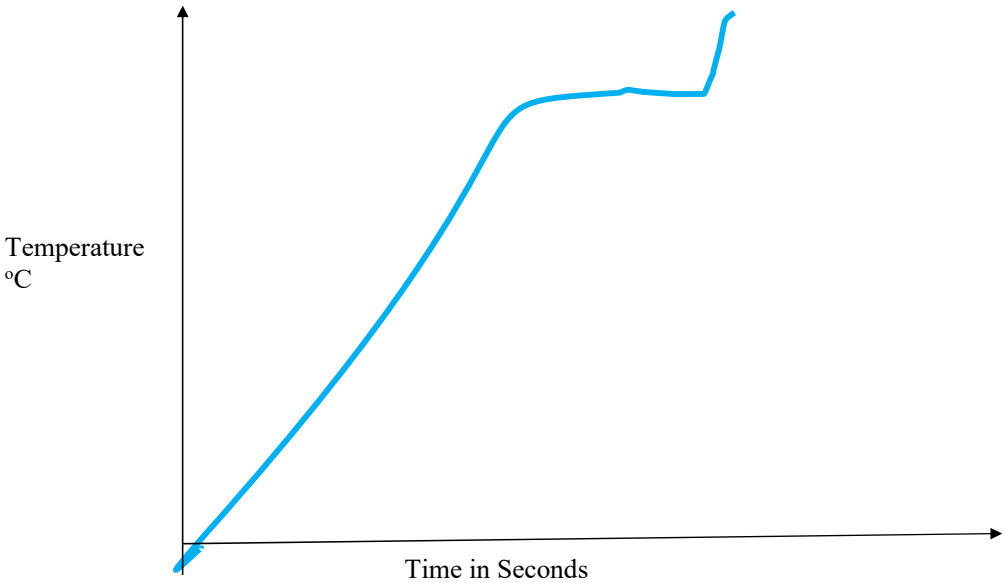
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Time (seconds)	0	30	60	90	120	150	180	210	240
Temperature(^{o}C)	25.0	45.0	85.0	95.0	96.0	96.0	96.0	97.0	98.0

Questions

1.Plot a graph of temperature against time(y-axis)

Sketch graph of temperature against time



2.From the graph show and determine the boiling point of water

Note:

Water boils at 100°C at sea level/one atmosphere pressure/101300Pa but boils at below 100°C at higher altitudes.

The sample results above are from Kiriari Girls High School-Embu County on the slopes of Mt Kenya in Kenya. Water here boils at 96°C.

3.Calculate the molar heat of vaporization of water.(H= 1.0,O= 16.O)

Working:

Mass of water = density x volume => (20 x 1) /1000 = 0.02kg

Quantity of heat produced

= mass of water x specific heat capacity of water x temperature change

=>0.02kg x 4.2 x (96 – 25) = -5.964kJ

Heat of vaporization of one mole H_2O

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= Quantity of heat
Molar mass of H_2O

=> -5.964kJ = -0.3313 kJ mole $^{-1}$
18

To determine the melting point of candle wax

Procedure

Weigh exactly 5.0 g of candle wax into a boiling tube. Heat it on a strongly Bunsen burner flame until it completely melts.

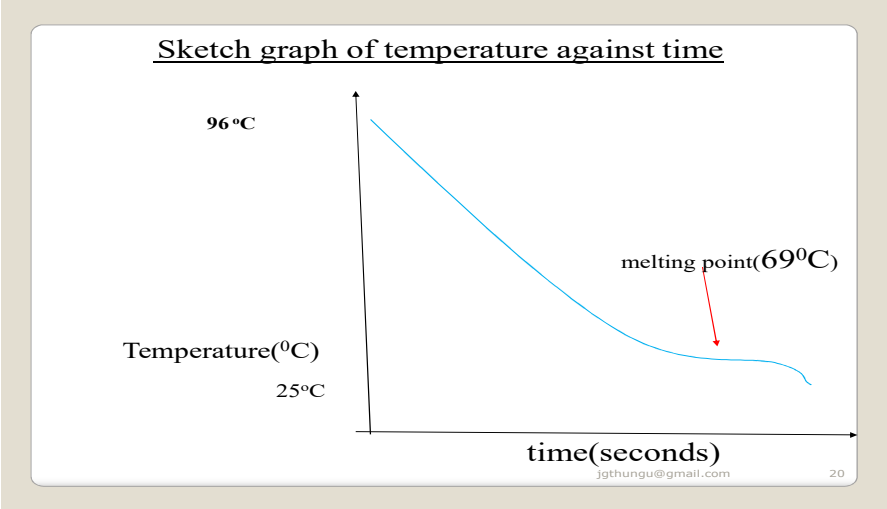
Insert a thermometer and remove the boiling tube from the flame. Stir continuously. Determine and record the temperature after every 30seconds for four minutes.

Sample results

Time (seconds)	0	30	60	90	120	150	180	210	240	240
Temperature (^{o}C)	93.0	85.0	78.0	70.0	69.0	69.0	69.0	67.0	65.0	65.0

Questions

1.Plot a graph of temperature against time(y-axis)



b)Energy changes in chemical processes

(i)Standard enthalpy/heat of displacement ΔH^{θ}_{d}

(ii)Standard enthalpy/heat of neutralization ΔH^{θ}_{n}

(iii)Standard enthalpy/heat of solution/dissolution ΔH^{θ}_{s}

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(i) Standard enthalpy/heat of displacement ΔH^{θ}_{d}

The molar standard enthalpy/heat of displacement may be defined as the energy/heat change when one mole of substance is displaced /removed from its solution at standard conditions

Some displacement reactions

(i)$Zn(s) + CuSO_4(aq) \rightarrow Cu(s) + ZnSO_4(aq)$

Ionically: $Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}(aq)$

(ii)$Fe(s) + CuSO_4(aq) \rightarrow Cu(s) + FeSO_4(aq)$

Ionically: $Fe(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Fe^{2+}(aq)$

(iii)$Pb(s) + CuSO_4(aq) \rightarrow Cu(s) + PbSO_4(\mathbf{s})$

This reaction stops after some time as insoluble $PbSO_4(\mathbf{s})$ coat/cover unreacted lead.

(iv)$Cl_2(g) + 2NaBr(aq) \rightarrow Br_2(aq) + 2NaCl(aq)$

Ionically:

$Cl_2(g) + 2Br^-(aq) \rightarrow Br_2(aq) + 2Cl^-(aq)$

To determine the molar standard enthalpy/heat of **displacement(ΔH^{θ}_{d})** of copper

Procedure

Place 20cm3 of 0.2M copper(II)sulphate(VI)solution into a 50cm3 plastic beaker/calorimeter.

Determine and record the temperature of the solution T_1.

Put all the Zinc powder provided into the plastic beaker. Stir the mixture using the thermometer.

Determine and record the highest temperature change to the nearest 0.5°C- T_2 .

Repeat the experiment to complete table 1 below

Sample results Table 1

Experiment	I	II
Final temperature of solution(T_2)	$30.0^{o}C$	$31.0^{o}C$
initial temperature of solution(T_1)	$25.0^{o}C$	$24.0^{o}C$
Change in temperature(ΔT)	5.0	6.0

Questions

1.(a) Calculate:

(i)average ΔT

Average ΔT = change in temperature in experiment I and II

$$\Rightarrow \frac{5.0 + 6.0}{2} = \mathbf{5.5^{o}C}$$

(ii)the number of moles of solution used

$$\text{Moles used} = \frac{\text{molarity} \times \text{volume of solution}}{1000} = \frac{0.2 \times 20}{1000} = \mathbf{0.004\ moles}$$

(iii) the enthalpy change ΔH for the reaction

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Heat produced ΔH = mass of solution(**m**) x specific heat capacity (**c**)x ΔT

$$\Rightarrow \frac{20 \times 4.2 \times 5.5}{1000} = \frac{462 \text{ Joules}}{1000} = \textbf{-0.462 kJ}$$

(iv)State two assumptions made in the above calculations.

Density of solution = density of water = $1 gcm^{-3}$

Specific heat capacity of solution=Specific heat capacity of water =$4.2 kJ^{-1}kg^{-1}K$

This is because the solution is assumed to be **infinite dilute.**

2. Calculate the enthalpy change for one mole of displacement of Cu^{2+} (aq) ions.

Molar heat of displacement ΔH_d = Heat produced ΔH

Number of moles of fuel

$$\Rightarrow \frac{0.462 \text{ kJ}}{0.004} = \textbf{-115.5 kJmole}^{-1}$$

3.Write an ionic equation for the reaction taking place.

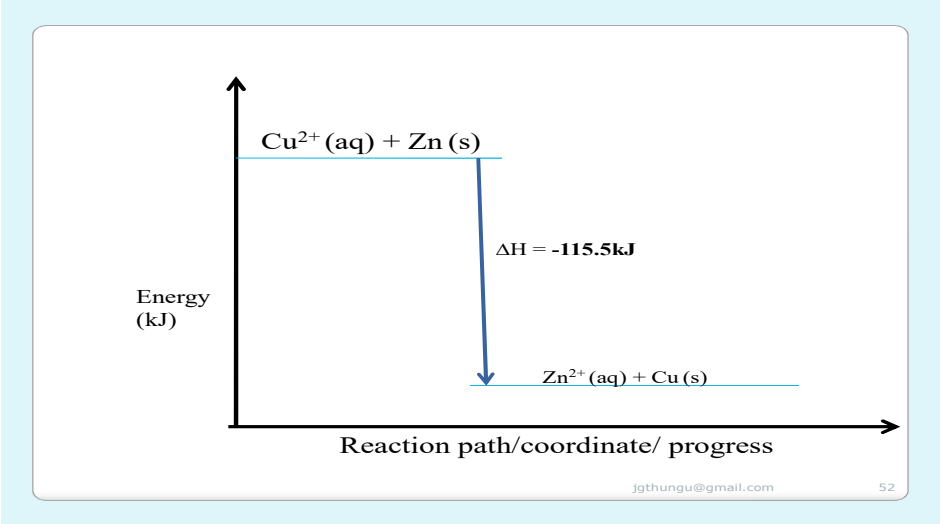
$Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}(aq)$

4.State the observation made during the reaction.

Blue colour of copper(II)sulphate(VI) fades/becomes less blue/colourless.

Brown solid deposits are formed at the bottom of reaction vessel/ beaker.

5.Illustrate the above reaction using an energy level diagram.



8. The enthalpy of displacement ΔH_d of copper(II)sulphate (VI) solution is $12k6kJmole^{-1}$.Calculate the molarity of the solution given that 40cm3 of this solution produces 2.204kJ of energy during a displacement reaction with excess iron filings.

Number of moles = Heat produced ΔH

Molar heat of displacement ΔH_d

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$$\Rightarrow \quad \frac{2.204\ \text{kJ}}{126\ \text{moles}} = \mathbf{0.0206 moles}$$

Molarity of the solution = $\frac{\text{moles} \times 1000}{\text{Volume of solution used}}$

$$= \frac{0.0206\text{moles} \times 1000}{40} = \mathbf{0.5167\ M}$$

Graphical determination of the molar enthalpy of displacement of copper

Procedure:

Place 20cm3 of 0.2M copper(II)sulphate (VI) solution into a calorimeter/50cm3 of plastic beaker wrapped in cotton wool/tissue paper.

Record its temperature at time T= 0.Stir the solution with the thermometer carefully and continue recording the temperature after every 30 seconds .

Place all the (1.5g) Zinc powder provided after 1 ½ minutes.

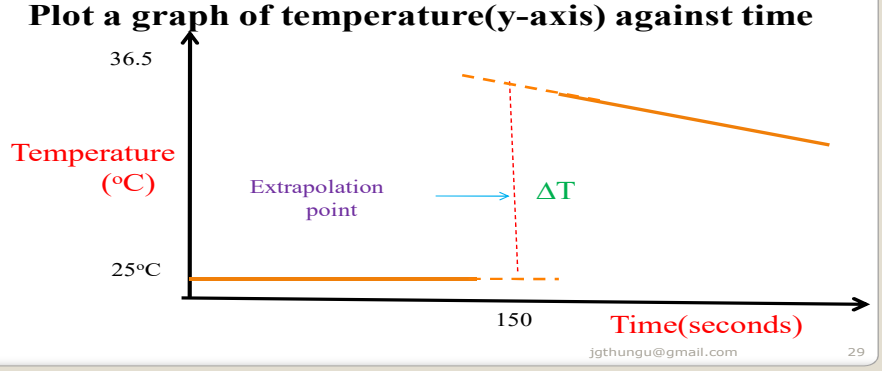
Stir the solution with the thermometer carefully and continue recording the temperature after every 30 seconds for five minutes.

Determine the highest temperature change to the nearest 0.5°C.

Sample results

Time °C	0.0	30.0	60.0	90.0	120.0	150.0	180.0	210.0	240.0	270.0
Temperature	25.0	25.0	25.0	25.0	25.0	xxx	36.0	35.5	35.0	34.5

Plot a graph of temperature(y-axis) against time



Questions

1.Show and determine the change in temperature ΔT

From a well constructed graph ΔT= T2 –T1 at **150** second by **extrapolation**

$\Delta T = 36.5 - 25.0 = 11.5^{o}C$

2.Calculate the number of moles of copper(II) sulphate(VI)used given the molar heat of displacement of Cu^{2+} (aq)ions is 125kJmole^{-1}

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Heat produced ΔH = mass of solution(**m**) x specific heat capacity (**c**)x ΔT

$$\Rightarrow \frac{20 \times 4.2 \times 11.5}{1000} = 966 \text{ Joules} = \textbf{-0.966 kJ}$$

Number of moles = Heat produced ΔH / Molar heat of displacement ΔH_d

$$\Rightarrow \frac{0.966 \text{ kJ}}{125 \text{ moles}} = \textbf{-0.007728moles}$$

$$\textbf{-7.728 x } 10^{-3}\textbf{moles}$$

3. What was the concentration of copper(II)sulphate(VI) in moles per litre.

$$\text{Molarity} = \frac{\text{moles} \times 1000}{\text{Volume used}}$$

$$\Rightarrow \frac{7.728 \times 10^{-3}\text{moles} \times 1000}{20} = \textbf{0.3864M}$$

4.The actual concentration of copper (II) Sulphate (VI) solution was 0.4M. Explain the differences between the two.

Practical value is **lower** than theoretical

. Heat/energy **loss** to the surrounding and that absorbed by the reaction vessel **decreases** ΔT hence **lowering** the practical number of moles and molarity against the theoretical value

(c)Standard enthalpy/heat of neutralization ΔH^{θ}_{n}

The molar standard enthalpy/heat of **neutralization** ΔH^{θ}_{n} is defined as the energy/heat change when one mole of a H^+ (H_3O^+)ions react completely with one mole of OH^- ions to form one mole of H_2O/water.

Neutralization is thus a reaction of an acid /H^+ (H_3O^+)ions with a base/alkali/ OH^- ions to form salt and water only.

Strong acids/bases/alkalis are completely/fully/wholly dissociated to **many** free ions(H^+ /H_3O^+ and OH^- ions).

(ii) for strong acid/base/alkali neutralization, no **energy** is used to dissociate /ionize since **molecule** is wholly/fully dissociated/ionized into free H^+ H_3O^+ and OH^- ions. The overall energy evolved is comparatively **higher / more** than weak acid-base/ alkali neutralizations.

For strong acid-base/alkali neutralization, the enthalpy of neutralization is **constant** at about **57.3**kJmole^{-1} irrespective of the acid-base used.

This is because ionically:

$$\textbf{OH}^-\textbf{(aq)} + \textbf{H}^+\textbf{(aq)} \rightarrow \textbf{H}_2\textbf{O(l)}$$

for all wholly/fully /completely dissociated acid/base/alkali

Weak acids/bases/alkalis are partially dissociated to **few** free ions(H^+ (H_3O^+ and OH^- ions) and exist **more** as molecules.

Neutralization is an exothermic(-ΔH) process.

The energy produced during neutralization depend on the amount of free ions (H^+ H_3O^+ and OH^-)ions existing in the acid/base/alkali reactant:

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(i)for weak acid-base/alkali neutralization,some of the **energy** is used to dissociate /ionize the **molecule** into free H^+ H_3O^+ and OH^- ions therefore the overall energy evolved is comparatively **lower/lesser/smaller** than strong acid / base/ alkali neutralizations.

Practically ΔH^{θ}_n can be determined as in the examples below:

To determine the molar enthalpy of neutralization ΔH_n of Hydrochloric acid

Procedure

Place 50cm3 of 2M hydrochloric acid into a calorimeter/200cm3 plastic beaker wrapped in cotton wool/tissue paper.

Record its temperature T_1.

Using a clean measuring cylinder, measure another 50cm3 of 2M sodium hydroxide.

Rinse the bulb of the thermometer in distilled water.

Determine the temperature of the sodium hydroxide T_2.

Average T_2 andT_1 to get the initial temperature of the mixture T_3.

Carefully add all the alkali into the calorimeter/200cm3 plastic beaker wrapped in cotton wool/tissue paper containing the acid.

Stir vigorously the mixture with the thermometer.

Determine the highest temperature change to the nearest 0.5°C T_4 as the final temperature of the mixture.

Repeat the experiment to complete table 1.

(ii)enthalpy change ΔH of neutralization.

ΔH = (**m**)mass of solution(acid+base) x (**c**)specific heat capacity of solution x Δ**T**(T_6) => (50 +50) x 4.2 x 13.5 = **5670J**oules = **5.67kJ**

(iii) the molar heat of neutralization the acid.

ΔH_n = Enthalpy change ΔH => 5.67kJ = **56.7kJ mole^{-1}**

Number of moles 0.1moles

(c)Write the ionic equation for the reaction that takes place

$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$

(d)The theoretical enthalpy change is 57.4kJ. Explain the difference with the results above.

The theoretical value is higher

Heat/energy loss to the surrounding/environment lowers Δ**T**/T_6 and thus ΔH_n

Heat/energy is absorbed by the reaction vessel/calorimeter/plastic cup lowers ΔT and hence ΔH_n

Sample results

Experiment	I	II
Temperature of acid T_1 (°C)	22.5	22.5
Temperature of base T_2 (°C)	22.0	23.0

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Final temperature of solution T_4(^oC)	35.5	36.0
Initial temperature of solution T_3(^oC)	22.25	22.75
Temperature change(T_5)	13.25	13.75

(a)Calculate T_6 the average temperature change T_6 =

13.25 +13.75 = **13.5 ^oC** 2

(b)Why should the apparatus be very clean?

Impurities present in the apparatus reacts with acid /base lowering the overall temperature change and hence ΔH^{θ}_n.

(c)Calculate the:

(i)number of moles of the acid used

number of moles = molarity x volume => 2 x 50 = **0.1moles**

1000 1000

(e)Compare the ΔH_n of the experiment above with similar experiment repeated with neutralization of a solution of:

(i) potassium hydroxide with nitric(V) acid

The results would be the same/similar.

Both are neutralization reactions of strong acids and bases/alkalis that are fully /wholly dissociated into many free H^+ / H_3O^+ and OH^- ions.

(ii) ammonia with ethanoic acid

The results would be lower/ΔH_n would be less.

Both are neutralization reactions of weak acids and bases/alkalis that are partially /partly dissociated into few free H^+ / H_3O^+ and OH^- ions. Some energy is used to ionize the molecule.

(f)Draw an energy level diagram to illustrate the energy changes

Theoretical examples

1.The molar enthalpy of neutralization was experimentary shown to be 51.5kJ per mole of 0.5M hydrochloric acid and 0.5M sodium hydroxide. If the volume of sodium hydroxide was 20cm3, what was the volume of hydrochloric acid used if the reaction produced a 5.0^{o}C rise in temperature?

Working:

Moles of sodium hydroxide = molarity x volume

1000

=> 0.5 M x 20cm3 = **0.01 moles**

1000

Enthalpy change ΔH = ΔH_n => 51.5 = **0.515kJ**

Moles sodium hydroxide 0.01 moles

Mass of base + acid = Enthalpy change ΔH in Joules

Specific heat capacity x ΔT

=> 0.515kJ x 1000 = **24.5238**g

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$$\frac{4.2 \times 5}{}$$

Mass/volume of HCl = Total volume – volume of NaOH

=>24.5238 - 20.0 = **4.5238 cm3**

Graphically ΔH_n can be determined as in the example below:

Procedure

Place 8 test tubes in a test tube rack .

Put 5cm3 of 2M sodium hydroxide solution into each test tube. Measure 25cm3 of 1M hydrochloric acid into 100cm3 plastic beaker.

Record its initial temperature at volume of base =0.

Put one **portion** of the base into the beaker containing the acid.

Stir carefully with the thermometer and record the highest temperature change to the nearest 0.5°C.

Repeat the procedure above with other portions of the base to complete table 1 below

Volume of acid(cm3)	25.0	25.0	25.0	25.0	25.0	25.0	25.0	25.0	25.0	25.0	25.0
Volume of alkali(cm3)	0	5.0	10.0	15.0	20.0	25.0	30.0	35.0	40.0	35.0	40.0
Final temperature(^{o}C)	**22.0**	**24.0**	**26.0**	**28.0**	**28.0**	**27.0**	**26.0**	**25.0**	**24.0**	**25.0**	**24.0**
Initial temperature(^{o}C)	**22.0**	**22.0**	**22.0**	**22.0**	**22.0**	**22.0**	**22.0**	**22.0**	**22.0**	**22.0**	**22.0**
Change in temperature	**0.0**	**2.0**	**4.0**	**6.0**	**6.0**	**5.0**	**4.0**	**3.0**	**2.0**	**3.0**	**2.0**

Complete the table to determine the change in temperature.

Plot a graph of volume of sodium hydroxide against temperature change.

From the graph show and determine :

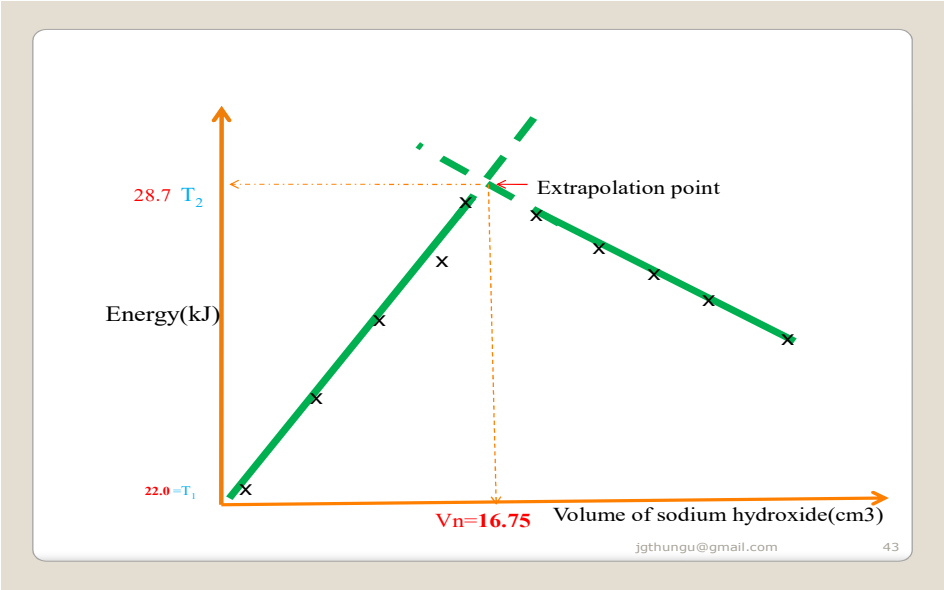
(i)the highest temperature change ΔT

ΔT =T_2-T_1 : highest temperature-T_2 (from extrapolating a correctly plotted graph) less lowest temperature at volume of base=0-T_1

=> 28.7 – 22.0 = **6.7 0 °C**

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(ii) the volume of sodium hydroxide used for complete neutralization
From correctly plotted graph = **16.75 cm3**
(iii) Calculate the number of moles of the alkali used
Moles NaOH = molarity x volume ()Vn=

$$\frac{1000}{1000}$$

$$\Rightarrow \frac{2 \times 16.75}{1000} = \mathbf{0.0335\ moles}$$

(iv)Calculate ΔH for the reaction.

ΔH = mass of solution **mixture** x c x **ΔT**

$$\Rightarrow (25.0 + 16.75) \times 4.2 \times 6.7$$

$$= \frac{\mathbf{1174.845\ J}}{1000} = \mathbf{1.174845\ kJ}$$

(iii) Calculate the molar enthalpy of the alkali:

ΔHn = Heat change => 1.174845 **kJ**
number of moles 0.0335 **moles**
= **35.0699kJ mole^{-1}**

(i) Standard enthalpy/heat of solution/dissolution ΔH^{θ}_{s}

The standard enthalpy of solution ΔH^{θ}_{s} is defined as the energy change when one mole of a substance is **dissolved** in excess distilled water to form an **infinite** dilute solution.

An infinite dilute solution is one which is **too** dilute to be diluted further.

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Practically the heat of solution is determined by dissolving a known mass /volume of a solute in known mass/volume of water/solvent and determining the temperature change.

To determine the heat of dissolution of ammonium nitrate(V)

Place 100cm3 of distilled water into a plastic cup/beaker/calorimeter

Put all the 5.0g of ammonium nitrate(v)/potassium nitrate(V)/ ammonium chloride into the water.

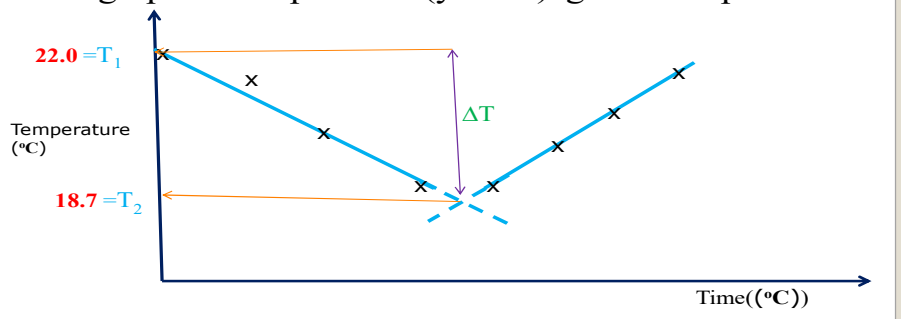
Stir the mixture using the thermometer and record the temperature change after every ½ minute to complete table1.

Continue stirring throughout the experiment.

Sample results:Table1

time(minutes)	0	½	1	1 ½	2	2 ½	3	3 ½
Temperature (°C)	22.0	21.0	20.0	19.0	19.0	19.5	20.0	20.5

Plot a graph of temperature (y-axis)against temperature



(a)From the graph show and determine:

(i)the highest temperature change ΔT

$\Delta T = T_2 - T_1$: highest temperature-T_2 (from extrapolating a correctly plotted graph) less lowest temperature at volume of base=0-T_1

=> 18.7 – 22.0 = **3.3 °C** **(not -3.3 °C)**

(b) Calculate the total energy change ΔH during the reaction

ΔH = mass of water x c x ΔT

=>ΔH=100 x4.2 x 3.3 °C = $\frac{+1386\ J}{1000}$ = **+ 1.386 kJ**

(c) Calculate the number of moles of ammonium nitrate (v) used

Moles = $\frac{\text{mass}}{\text{molar mass}}$ => $\frac{5.0}{80}$ = **0.0625 moles**

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(d)What is the molar heat of dissolution of ammonium nitrate(V)

ΔH = Heat change = + 1.386 kJ = + **22.176 kJmole^{-1}**

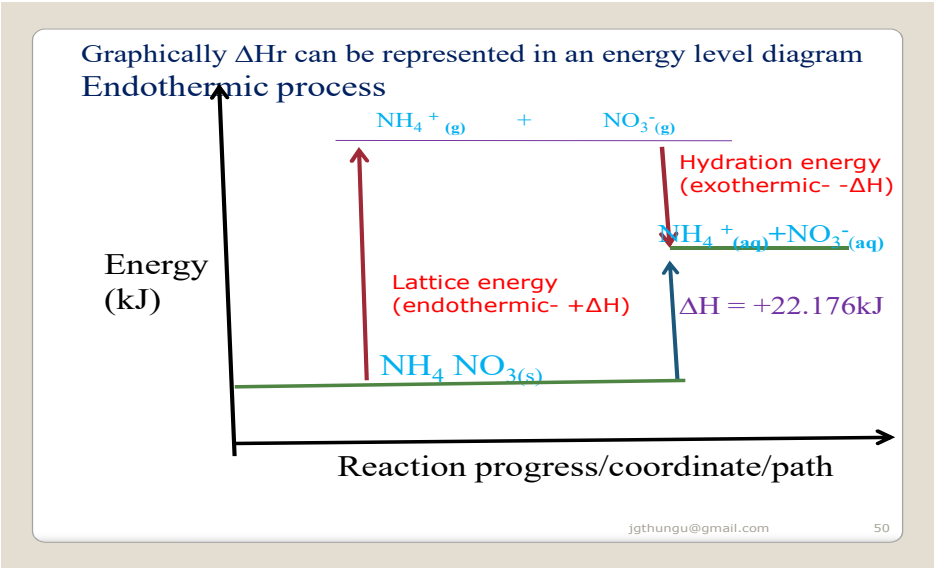
Number of mole 0.0625 moles

(e)What would happen if the distilled water is heated before experiment was performed .

The ammonium nitrate(V) would take less time to dissolve.

Increase in temperature reduces lattice energy causing endothermic dissollution to be faster.

(e)Illustrate the above process on an energy level diagram



c)Chemical Kinetic/Rate of reaction

The rate of a chemical reaction can be defined as the **time** taken for a known amount of reactants to form known amount of products.

Some reactions are **too slow** to be determined e.g weathering others are **instantaneous**

The SI unit of time is **seconds**. Minutes and hours are also common .

Time is determined using a **stop watch/clock**

Candidates using stop watch/clock should learn to:

(i)Press start button concurrently with starting off determination of a reaction using one hand each.

(ii)Press stop button when the reaction is over.

(iii)Record all times in seconds unless specified.

(iv)Press reset button to begin another timing

(v)Ignore time beyond seconds for stop clock/watch beyond this accuracy

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(vi)Avoid accidental pressing of any button before recording
It can be very **frustrating** repeating a whole procedure
The following factors theoretically and practically alter/influence/affect/determine the rate of a chemical reaction:
(a)Concentration
(b)Temperature
(a)Concentration
An **increase** in concentration increases the rate the rate of reaction by **reducing** the **time** taken to completion.
Theoretically, increase in concentration is a decrease in distance between reacting particles which increases their **collision frequency**.
Practically **decreasing** concentration is **diluting/adding water**
To demonstrate the effect of concentration on reaction rate
You are provided with
(i) sodium thiosulphate containing $40gdm^{-3}$ solution labeled A
(ii) 2M hydrochloric acid labeled solution B
You are required to determine the rate of reaction between solution A and B
Procedure
Measure 40cm3 of solution A into 100 cm3 glass beaker. Place it on top of a pen-mark "X". Measure another 40cm3 of solution B. Simultaneously put solution B into solution A and start off a stop watch/clock. Determine the time taken for the pen-mark "X" to be invisible/obscurred from above. Repeat the procedure by measuring 35cm3 of solution B and adding 5cm3 of water. Complete the table 1 below by using other values os solution B and water

Sample results: Table 1

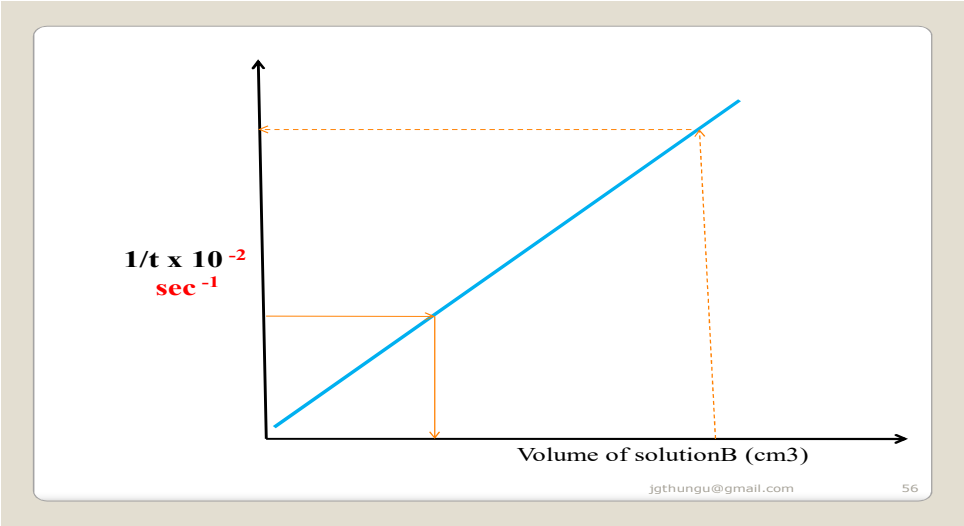
Volume of solution A (cm3)	40	40	40	40	40	40
Volume of solution B (cm3)	40	35	30	25	20	15
Volume of water (cm3)	0	5	10	15	20	25
Time taken for x to be invisible	14	17	21	25	50	120
1/t	0.0714	0.0588	0.0351	0.04	0.02	0.0083

Calculate 1/t in each case

Plot a graph of 1/t (y-axis) against volume of solution B.

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Sample questions

(i)Explain the shape of the graph

(Straight line graph from the origin)

Decrease in concentration decreases the rate of reaction. The higher the concentration of solution B the less time taken for mark x to be obscurred/invisible due to increased collision frequency between the reacting particles.

(ii)From the graph determine the time taken for the mark to be invisible at 37cm3

At 37cm3 then 1/t => 1/ 37 = 0.027

From a well plotted graph:

1/t = 0.027 => **16.2602 seconds**

(ii)From the graph determine the volume of solution B at 100 seconds

100 seconds => 1/t = 1 / 1000 = 0.01

From a well plotted graph:

At 1/t = 0.01 => the volume of B = **17.0cm3**

(iii) State another factor that would alter the rate of the above reaction.

Temperature

(iii) State another factor that would not alter the rate of the above reaction.

Surface area

Pressure

Catalyst

(b) Temperature

An **increase** in temperature increases the rate of reaction.

An increase of 10 °C/10K practically doubles the rate of a chemical reaction/reduces time of completion by $^{1}/_{2}$.

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An increase in temperature increase the kinetic energy of reacting particles increasing their collision frequency
Practically ,increase in temperature involves **heating the reactants**
The results and presentation should be as in the effect of concentration.
Increased temperature reverses the table I time results
i.e less time as temperature increases.

d)Qualitative analysis

Process of identifying unknown compounds
Compounds may be:
(i)Inorganic
(ii)organic

Inorganic analysis:

This involve mainly identification of ionic compounds containing **cations** and **anions.**
Cations present in an ionic compounds are identified by adding a **precipitating reagent** that forms a precipitate unique to the cation/s in the compound.
The main **precipitating reagents** used are:
2M NaOH and/or **2M NH_3(aq)**
When using 2M sodium hydroxide:
(i)No white precipitate is formed if **K^+** and **Na^+** ions are present
(ii) No white precipitate is formed if **NH_4^+** ions are present but a clourless gas with pungent smell of urine is produced which may not be recognized in a school laboratory examination setting.
(iii)White precipitate that dissolves / soluble in excess if **Zn^{2+} Pb^{2+} Al^{3+}** ions are present.
(iv)White precipitate that do not dissolves/insoluble in excess if **Ba^{2+} Mg^{2+} Ca^{2+}** ions are present.
(v)Blue precipitate that do not dissolves /insoluble in excess if **Cu^{2+}** ions are present.
(vi)Green precipitate that do not dissolves/insoluble in excess if **Fe^{2+}** ions are present.
(vii)Brown precipitate that do not dissolves/insoluble in excess if **Fe^{3+}** ions are present.
When using 2M aqueous ammonia
(i)No white precipitate is formed if **K^+** ,**NH_4^+** **Na^+** ions are present
(ii)White precipitate that dissolves / soluble in excess if **Zn^{2+}** ions are present.
(iii)White precipitate that do not dissolves/insoluble in excess if **Ba^{2+} Mg^{2+} Ca^{2+} Pb^{2+} Al^{3+}** ions are present.
(iv)Blue precipitate that **dissolves** /soluble in excess to form a **deep/royal blue** solution in excess if **Cu^{2+}** ions are present.

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(v)Green precipitate that do not dissolves/insoluble in excess if Fe^{2+} ions are present.
(vi)Brown precipitate that do not dissolves/insoluble in excess if Fe^{3+} ions are present.
Anions present in an ionic compounds are identified by adding a specific **precipitating reagent** that forms a precipitate unique to the specific anion/s in the compound.
(i)Lead(II)nitrate(V) solution
Lead forms **insoluble** $PbSO_4$,$PbSO_3$,$PbCO_3$, PbS, PbI_2,$PbCl_2$
PbS is a **black** precipitate,
PbI_2 is a yellow precipitate.
All the others are white precipitates.
(a)If a Lead(II)nitrate(V) solution is added to a substance/ solution/ compound :
(i)A yellow ppt shows presence of I^- ions
(ii)A **black** ppt shows presence of S^{2-} ions
(iii) A white ppt shows presence of SO_4^{2-} ,SO_3^{2-} ,CO_3^{2-} Cl^-
(b)If the white precipitate is added dilute nitric(V) acid:
(i)It dissolves to show presence of SO_3^{2-} ,CO_3^{2-}
(ii)It persist/remains to show presence of SO_4^{2-}, Cl^-
(c)If the white precipitate in b(i) is added acidified potassium manganate(VII)/ dichromate(VI)
(i) acidified potassium manganate(VII) is decolorized /orange colour of acidified potassium dichromate(VI) turns to green to show presence of SO_3^{2-}
(ii) acidified potassium manganate(VII) is not decolorized /orange colour of acidified potassium dichromate(VI) does not turn to green/remains orange to show absence of SO_3^{2-} /presence of CO_3^{2-}
(c)If the white precipitate in b(ii) is **boiled**:
(i)It dissolves to show presence of Cl^-
(ii)It persist/remains to show presence of SO_4^{2-}
(ii)Barium(II)nitrate(V)/Barium chloride solution
Barium(II)nitrate(V)/Barium chloride solution precipitates $BaSO_4$,$BaSO_3$, $BaCO_3$, from SO_4^{2-} ,SO_3^{2-} ,CO_3^{2-} ions.
Inorganic qualitative analysis require continous practice discussion
Sample presentation of results
You are provided with solid Y(aluminium (III)sulphate(VI)hexahydrate).Carry out the following tests and record your observations and inferences in the space provided.
1(a) Appearance

Observations	inference
(1mark)	

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White crystalline solid	Coloured ions Cu^{2+} , Fe^{2+} ,Fe^{3+} absent

(b)Place about a half spatula full of the solid into a clean dry boiling tube. Heat gently then strongly.

Observations (1mark)	inference
Colourless droplets formed on the cooler part of the test tube	Hydrated compound/compound containing water of crystallization
Solid remains a white residue	

(c)Place all the remaining portion of the solid in a test tube .Add about 10cm3 of distilled water. Shake thoroughly. Divide the mixture into five portions.

Observation	Inference (1mark)
Solid dissolves to form a colourless solution	Polar soluble compound Cu^{2+} , Fe^{2+} ,Fe^{3+} absent

(i)To the first portion, add three drops of sodium hydroxide then add excess of the alkali.

Observation	Inference (1mark)
White ppt, soluble in excess	Zn^{2+} , Pb^{2+} , Al^{3+}

(ii)To the second portion, add three drops of aqueous ammonia then add excess of the alkali.

Observation	Inference (1mark)
White ppt, insoluble in excess	Pb^{2+} , Al^{3+}

(iii)To the third portion, add three drops of sodium sulphate(VI)solution.

Observation	Inference (1mark)
No white ppt	Al^{3+}

(iv)I.To the fourth portion, add three drops of Lead(II)nitrate(IV)solution. **Preserve**

Observation	Inference (1mark)
White ppt	CO_3^{2-}, SO_4^{2-}, SO_3^{2-}, Cl^-,

II.To the portion in (iv) I above , add five drops of dilute hydrochloric acid.

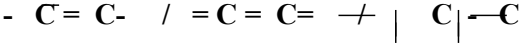
Observation	Inference (1mark)
White ppt persist/remains	SO_4^{2-}, Cl^-,

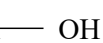
III.To the portion in (iv) II above, heat to boil.

Observation	Inference (1mark)
White ppt persist/remains	SO_4^{2-},

Organic analysis:

This involve mainly identification of the functional group:

(i) 

(ii) 

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(iii) R—— COOH / H^+

These functional groups can be identified by:
(i)burning-a substance which "catches fire" must reduce in amount.
Candidates should not confuse burning with flame coloration/test
(ii)Decolorization of bromine water/chlorine water/acidified $KMnO_4$ / to show **presence** of

- $C \equiv C$ - / - $\underset{|}{C} = \underset{\backslash}{C}$ - **and** R – OH

(iii)Turning orange acidified $K_2Cr_2O_7$ to green to show **presence** as in above.
(iii)pH 1/2/3 for strongly acidic solutions. pH 4/5/6 for weakly acidic solutions
(iv)Turning blue litmus paper red. red litmus paper remaining red show **presence** of **H^+ ions**

d)Flame test

The colour change on a clear colourless Bunsen flame is useful in identifying some cations / metals.
A very clean metallic spatula is recommended since dirt obscures /changes the correct coloration distinct flame coloration of some compounds

Barium/barium salts	orange
Sodium/ sodium salts	yellow
Potassium/potassium salts	Purple/lilac
Lithium/Lithium salts	**Deep red**/crimson
Calcium/ calcium salts	red
Copper/copper salts	Blue/ green

(e)Physical chemistry

Chemistry is a science subject that **incorporate** many scientific techniques.
Examining body/council, require tabulated results/data from the candidate.
This tabulated results is usually then put in a graph.
The general philosophy of methods of **presentation of chemistry practical data** is therefore availability of **evidence** showing:
(i)Practical done(complete table)
(ii)Accuracy of apparatus used(decimal point)
(iii)Accuracy/care in doing experiment to get collect trend(against teachers results)

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(iv)Graphical work(use of mathematical science)
(v)Calculations (Scientific mathematical integration)

(f)Sample practicals
Name ……..…….Class……..Index No……..
Candidate's signature……..
Date done…….Date marked…… …….Date revised…….…..….
233/3
CHEMISTRY Paper 3
PRACTICAL.
Pre-KCSE Practice **1**: 2013
MARKS SCHEME

Instruction to Candidate
Write your name and index number in the spaces provided above.
Sign and write the date of examination in the spaces provided above
*Answer **all** questions in the spaces provided.*
Mathematical tables and electronic calculators may be used.
*All working **must** be clearly shown where necessary.*
This paper consist of 8 printed pages.
Candidates should check the question paper to ascertain that all the pages are printed and indicated and that no questions are missin

For examiners use only

Question	**Maximum score**	**Candidates core**
1	**20**	20
2	**10**	10
3	**10**	10
Total score	**40**	**40**

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1.You are provided with:

(i)solution L containing 5.0g per litre of a dibasic organic acid $H_2X.2H_2O$.
(ii)solution M which is acidified potassium manganate(VII)
(iii)solution N a mixture of sodium ethanedioate and ethanedioic acid
(iv)0.1M sodium hydroxide solution P
(v)1.0M sulphuric(VI)

You are required to:

(i)standardize solution M using solution L
(ii)use standardized solution M and solution P to determine the % of sodium ethanedioate in the mixture.

Procedure 1

Fill the burette with solution M. Pipette 25.0cm3 of solution L into a conical flask. Heat this solution to about 70°C(**but not to boil**).Titrate the hot solution L with solution M until a permanent pink colour just appears .Shake thoroughly during the titration. Repeat this procedure to complete table 1.

Table 1

	1	2	3
Final burette reading (cm3)	**20.0**	**20.0**	**20.0**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of N used (cm3)	**20.0**	**20.0**	**20.0**

Table 1
CT=½ mk
DP=½ mk
AC=½ mk
AV= 1 mk
FA=½ mk
Total=3 mk

(2marks)

(a)Calculate the average volume of solution L used
(1mk)

$$\frac{20.0 + 20.0 + 20.0}{3} = 20.0\text{cm3}$$

(b)Given that the concentration of the dibasic acid is 0.05molesdm^{-3}.determine the value of x in the formula $H_2X.2H_2O$ (H=1.0,O=16.0)
(1mark)

Molar mass $H_2X.2H_2O$ = mass / litre => moles / litre

5.0g/litre = 100 g
0.05molesdm^{-3}

$H_2X.2H_2O$ =100

X = 100 – ((2 x1) + 2 x (2 x1) + (2 x 16) => 100 – 34 = 62

(c) Calculate the number of moles of the dibasic acid $H_2X.2H_2O$. (1mark)

Moles = molarity x pipette volume / 1000 => 0.05 x 25 = 0.00125 / 1.25 x10 $^{-3}$ moles

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1000

(d)Given the mole ratio manganate(VII)(MnO_4^-): acid H_2X is 2:5, calculate the number of moles of manganate(VII) (MnO_4^-) in the average titre. (1mark)

Moles H_2X = $^2/_5$ moles of MnO_4^-

=> $^2/_5$ x 0.0125/1.25 x10 $^{-2}$ moles = 0.0005 / 5.0 x 10 $^{-4}$ moles

(e)Calculate the concentration of the manganate(VII)(MnO_4^-) in moles per litre. (1mark)

Moles per litre/molarity = moles x 1000
average burette volume

=> 0.0005/5.0 x 10 $^{-4}$moles x 1000 = 0.02083 moles l^{-1} / M
24.0

Procedure 2

With solution M still in the burette ,pipette 25.0cm3 of solution N into a conical flask. Heat the conical flask containing solution N to about 70°C.Titrate while hot with solution M. Repeat the experiment to complete table 2. Table 2 (2marks)

	1	2	3
Final burette reading (cm3)	**12.5**	**12.5**	**12.5**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of N used (cm3)	**12.5**	**12.5**	**12.5**

Table 1
CT=$^1/_2$ mk
DP=$^1/_2$ mk
AC=$^1/_2$ mk
AV= 1 mk
FA=$^1/_2$ mk
Total=3 mk

(a)Calculate the average volume of solution L used (1mk)

12.5 + 12.5 + 12.5 = 12.5 cm3
3

(b)Calculations:

(i)How many moles of manganate(VII)ions are contained in the average volume of solution M used? (1mark)

Moles = molarity of solution M x average burette volume
1000

=> 0.02083 molesl^{-1}/ M x 12.5 = 0.00026 / 2.6 x 10^{-4} moles
1000

(ii)The reaction between manganate(VII)ions and ethanedioate ions that reacted with is as in the equation:

$2MnO_4^-(aq) + 5C_2O_4^{2-}(aq) + 16H^+(aq) \rightarrow 2Mn^{2+}(aq) + 10CO_2(g) + 8H_2O(l)$

Calculate the number of moles of ethanedioate ions that reacted with manganate (VII) ions in the average volume of solution M. (1mark)

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From the stoichiometric/ionic equation:

mole ratio MnO_4^- (aq): $C_2O_4^{2-}$ (aq) = 2:5

=> moles $C_2O_4^{2-}$ = $^5/_2$ moles MnO_4^-

=> $^5/_2$ x 0.00026 / 2.5 x 10^{-3} moles

= 0.00065 / 6.5 x10^{-4} moles

(iii)Calculate the number of moles of ethanedioate ions contained in 250cm3 solution N. (1mark)

25cm3 pipette volume -> 0.00065 /6.5 x10^{-4} moles

250cm3 -> 0.0065 /6.5 x10^{-3} moles x 250 = 0.0065 / 6.5 x10^{-3} moles

25

Procedure 3

Remove solution M from the burette and rinse it with distilled water. Fill the burette with sodium hydroxide solution P. Pipette 25cm3 of solution N into a conical flask and add 2-3 drops of phenolphthalein indicator. Titrate this solution N with solution P from the burette. Repeat the procedure to compl

Table 3

	1	2	3
Final burette reading (cm3)	**12.5**	**12.5**	**12.5**
Initial burette reading (cm3)	**0.0**	**0.0**	**0.0**
Volume of N used (cm3)	**12.5**	**12.5**	**12.5**

Table 1
CT=$^1/_2$ mk
DP=$^1/_2$ mk
AC=$^1/_2$ mk
AV= 1 mk
FA=$^1/_2$ mk
Total=3mk

(2 mark)

(a)Calculate the average volume of solution L used (1mk)

$\frac{12.5 + 12.5 + 12.5}{3}$ = 12.5 cm3

(b)Calculations:

(i)How many moles of sodium hydroxide solution P were contained in the average volume?

(1mark)

Moles = $\frac{\text{molarity of solution P x average burette volume}}{1000}$

=> 0.1 molesl^{-1} x $\frac{24.9}{1000}$ = 0.00249 / 2.49 x 10^{-3} moles

(ii)Given that NaOH solution P reacted with the ethanedioate ions from the acid only and the equation for the reaction is:

$2NaOH(aq) + H_2C_2O_4(aq) \rightarrow Na_2C_2O_4(g) + 2H_2O(l)$

Calculate the number of moles of ethanedioic acid that were used in the reaction.(1 mk)

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From the stoichiometric equation,mole ratio

$NaOH(aq) : H_2C_2O_4(aq) = 2:1$

=> moles $H_2C_2O_4$ = $^1/_2$ moles NaOH

=> $^1/_2$ x 0.00249 / 2.49 x 10^{-3} moles

=0.001245/1.245 x10^{-3} moles.

(iii)How many moles of ethanedioic acid were contained in 250cm3 of solution N? (1mark)

25cm3 pipette volume -> 0.001245/1.245 x10^{-3}

250cm3 -> 0.001245/1.245 x10^{-3} moles x 250 / 25

= 0.01245/1.245 x10^{-2} moles

(iii)Determine the % by mass of sodium ethanedioate in the mixture (H= 1.0,O=16.0,C=12.0 and total mass of mixture =2.0 g in 250cm3 solution) (1mark)

Molar mass $H_2C_2O_4$ = 90.0g

Mass of $H_2C_2O_4$ in 250cm3 =

moles in 250cm3 x molar mass $H_2C_2O_4$

=>0.01245/1.245 x10^{-2} moles x 90.0

= 1.1205g

% by mass of sodium ethanedioate

=(Mass of mixture - mass of $H_2C_2O_4$) x 100%

Mass of mixture

=> 2.0 - 1.1205 g = 43.975%

2.0

2. You are provided with 5.0 g solid B. You are to determine the molar mass of solid B.

Procedure

Place 100cm3 of liquid L into a plastic beaker. Determine its temperature and record it at time = 0 in Table 2 below. Stir continuously using the thermometer and record the highest temperature change to the nearest 0.5°C after every 30 seconds. After 120 seconds, add all solid B. Continue stirring and recording the to complete table 2.

Table 2
CT=$^1/_2$ mk
DP=$^1/_2$ mk
AC=$^1/_2$ mk
TR=$^1/_2$ mk
Total=2 mk

Table 2

Time (seconds)	0.0	30	60	90	120	150	180	210	240	270	300
Temperature(°C)	**20**	**20**	**20**	**20**		**18**	**16**	**14**	**14**	**15**	**16**

(2mark)

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(a)Plot a graph of temperature against time(x-axis)(3marks)

Graph
Scale(plots cover over ½ graph paper)=½ mk
Labelling(both axis)=½ mk
Plotting all points= 1 mk
Shape(Extrapolated graph)=1mk
Total =3 mk

(b)From the graph show and determine (2 mark

(i) the highest temperature change ΔT

$\Delta T = T_2 - T_1$ => 13.4 -20 = 6.6° C

Note ΔT is not - 6.6°C

(ii) the temperature of the mixture at 130 seconds

From extrapolation at 130 seconds = 19.2 °C

(iii)the time when all the solid first dissolved

From extrapolation of the lowest temperature = 220 Seconds

(d) Calculate the heat change for the reaction.(Assume density of liquid L is 1.0gcm^{-3}) specific heat capacity is 4.2Jkg^{-1}K^{-1}(1mark)

ΔH = mass of liquid L x c x ΔT =>100 x 4.2 x 6.6 = + 2772 J = + 2.772 kJ

1000

(e) Given the molar enthalpy of dissolution of Solid B in liquid L is + 22.176kJ mole^{-1},determine the number of moles of B used(1mark)

Moles of B = ΔH => + 2.772 kJ = 0.125 moles

ΔHs + 22.176kJ mole^{-1}

(f)Calculate the molar mass of B (1mark)

Molar mass of B = Mass used => 5.0 =>40 g

Moles used 0.125 moles

3(a)You are provided with solid Y. Carry out the following tests and record your observations and inferences in the space provided.

(i) Appearance

Observations	**inference**
(1mark)	
White crystalline solid	**Coloured Fe^{2+} ,Fe^{3+} , Cu^{2+} ions absent**

(ii)Place about a half spatula full of the solid into a clean dry boiling tube. Heat gently then strongly.

Observations	**inference**
(1mark)	
Colourless droplets forms on the cooler parts of test tube	**Hydrated compound/salt**
~~Solid remain white~~	

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(ii)Place all the remaining portion of the solid in a test tube .Add about 10cm3 of distilled water. Shake thoroughly. Divide the mixture into five portions.

Observation	Inference (1mark)
Solid dissolves to form a colourless solution	**Coloured Fe^{2+} , Fe^{3+} , Cu^{2+} ions absent**

I. To the first portion add three drops of universal indicator. (1mark)

Observation	Inference
pH= 4	**weakly acidic solution**

II.To the second portion, add three drops of aqueous ammonia then add excess of the alkali.

Observation (1mark)	Inference
White ppt, insoluble in excess	**Al^{3+} , Pb^{2+}**

III.To the third portion, add three drops of sodium sulpide solution.

Observation (1mark)	Inference
No black ppt	**Al^{3+}**

IV.To the fourth portion, add three drops of acidified Lead(II)nitrate(IV)solution. Heat to boil

Observation (1mark)	Inference
White ppt , persist/remains on boiling	**SO_4^{2-}**

(b)You are provided with solid P. Carry out the following tests and record your observations and inferences in the space provided.

(i)Place a portion of solid P on a clean metallic spatula and introduce it on a Bunsen flame.

($^1/_2$ mark)	
Solid burns with a yellow sooty flame	**bonds**

(ii)Add all the remaining solid to about 10cm3 of water in a test tube and shake well. Divide the mixture into 4 portions. ($^1/_2$ mark)

Solid dissolves to form a clourless solution	**Polar organic compound**

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I. To the 1st portion, test with litmus papers (1/2 mark) **Red litmus paper remain red** **Blue litmus paper turn blue**	**H^+ ions**
II. To the 2nd portion, add a little sodium hydrogen carbonate(1/2 mark)	
Effervescence/fizzing/bubbles **Colourless gas produced**	**H^+ ions**
III.To the 3rd portion, and three drops of solution M. Warm(1/2 mark) **Acidified $KMnO_4$ is decolorized** **// solution M is decolorized**	**OH, C C C C bonds**
IV.To the 4th portion, add three drops of bromine water (1/2marks) **Bromine water is decolorized**	**C C // C C bonds**

233/3 CHEMISTRY
Pre-KCSE 2013
Practice 1
Moi High School-Mbiruri
Requirements for each Candidates:

0.05 M Oxalic acid labeled **Solution L**
0.01M Potassium manganate (VII) labeled **Solution M**
0.03M oxalic acid labeled **Solution N**
0.1M Sodium hydroxide labeled **Solution P**
1.0M sulphuric(VI)acid.
15 0cm3 distilled water labeled Liquid L
50cm3 burette
25cm3 pipette
Two clean conical flasks
Pipette filler
-10 -110^{0}C Thermometer
Stop watch/clock
200cm3 clean beaker

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5.4g ammonium chloride/**8.0**g Ammonium nitrate(V) labeled **Solid B weighed accurately**
About 2.0 g of hydrated Aluminium sulphate labeled **Solid Y**
About 2.0g of Citric acid labeled **Solid P**
One boiling tube
Six clean dry test tubes
Pair of litmus papers(red and blue)
Clean metallic spatula

Access to bench reagents/apparatus

Means of heating
2M aqueous ammonia
0.1M acidified Lead(II)nitrate(V)
Bromine water
0.1M sodium sulphide
About 0.1g Sodium hydrogen carbonate
Universal indicator solution
pH chart

Name …….Index Number……..
233/3 **Candidates signature……**
CHEMISTRY **Date……..**
Paper 3
PRACTICAL
Practice 2012
$2^{1}/_{4}$ hours

You are provided with :

- Solution A containing an oxidizing agent A;
- Solution B ,0.05M aqueous sodium thiosulphate;
- Solution C containing a reducing agent C;
- Aqueous Potassium iodide;
- Solution D, starch solution.

You are required to determined the:

Concentration of solution A
Rate of reaction between the oxidizing agent A and the reducing agent C.

Procedure 1

1. Using a pipette and **pipette filter** ,place 25.0cm3 of solution A into a 250ml conical flask.
2. Meassure 10cm3 of aqueous potassium iodide and add it to solution A in the conical flask.
   Shake the mixture .Add 10cm3 of 2M sulphuric(VI)acid to the mixture and shake.

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3. Fill a burette with solution **B** and use it to titrate the mixture in the conical flask until it just turns **orange yellow**.Add 2cm3 of solution D to the mixture in the conical flask .Shake thoroughly. Continue titrating until the mixture **just turns colourless**. Record your results in **table 1** below.
4. Repeat the procedure and complete table 1.**Retain the remainder** of solution **A** and **D** for use in procedure II

Table I

	I	II	III
Final burette reading	**20.0**	**20.0**	**20.0**
Initial burette reading	**0.0**	**0.0**	**0.0**
Volume of solution B used (cm3)	**20.0**	**20.0**	**20.0**

(4mks)

(a)Calculate the:

(i) average volume of solution B used(1mk)

$$\frac{20.0 + 20.0 + 20.0}{3} \surd = 20.0 \surd \text{cm3}$$

(ii)number of moles of sodium thisulphate(1mk)

Moles = molarity x burette volume

$$\frac{\text{burette volume}}{1000}$$

$$\Rightarrow \frac{0.05 \times 20.0}{1000} \surd = 0.001 / 1.0 \times 10^{-3} \surd \text{ moles}$$

(b)Given that one mole of A reacts with six moles of sodium thiosulphate ,calculate the:

(i)number of moles of A that were used (1mk)

Mole ratio A:B = 1:6 √

$$\Rightarrow \text{Moles A} = \frac{0.001 / 1.0 \times 10^{-3} \text{ moles}}{6} = 0.00016/1.6 \times 10^{-4} \surd \text{ moles}$$

(ii)concentration of solution A in moles per litre(2mk)

Molarity of solution A = moles x 1000

$$\frac{\text{Pipette volume}}{1000}$$

$$\Rightarrow \frac{0.00016/1.6 \times 10\text{-}4 \text{ moles} \times 1000}{20} \surd = 0.008/8.0 \times 10^{-4} \text{ M} \surd$$

Procedure II

1. Label six test tubes as 1,2,3,4 ,5 and 6 and place them on a test tube rack.
2. Using a clean burette, measure the volumes of distilled water as shown in table 2 into the labeled test tubes.

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3. Using a burette ,measure the volumes of solution A shown in table 2 into each of the test tubes .
4. Clean the burette and rinse it with about 5cm3 of solution C
5. Using the burette ,measure 5cm3 of solution C and place it into a 100ml beaker.
6. Using a 10ml measuring cylinder ,measure 5cm3 of solution D and add it to the beaker containing solution C .Shake the mixture.
7. Pour the contents of test tube number 1 to the mixture in the beaker and immediately start off stop watch/clock. Swirl the contents of the beaker.Record the time taken for a blue colour to appear in table 2.
8. Repeat steps 5 to7 using the contents of test-tube 2,3,4,5 and 6.
9. Complete table 2 by computing Rate = $\frac{1}{\text{Time}}$ (S^{-1})

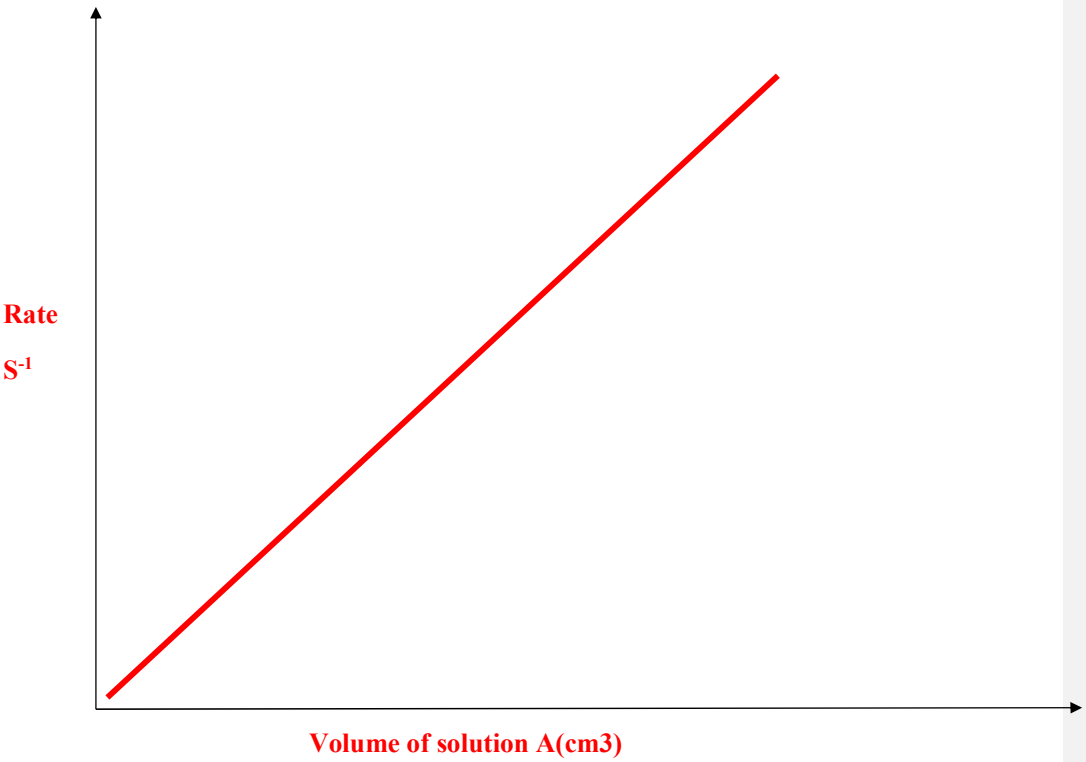
Table 2(Sample results)

Test-tube number	**1**	**2**	**3**	**4**	**5**	**6**
Volume of distilled water(cm3)	**0**	**2**	**3**	**5**	**6**	**7**
Volume of solution A(cm3)	**10**	**8**	**7**	**5**	**4**	**3**
Time(seconds)	**40.0**	**60.0**	**70.0**	**90.0**	**100.0**	**110.0**
Rate = $\frac{1}{\text{time}}$ (S^{-1})	**0.025 2.5 x 10^{-2}**	**0.0167 1.67 x 10^{-2}**	**0.0143 1.43 x 10^{-2}**	**0.0111 1.11 x 10^{-2}**	**0.0.1 1.11 x 10^{-2}**	**0.0083 8.3 x 10^{-3}**

Plot a graph of rate(y-axis)against volume of solution A(3mk)

Sketch graph of rate against time

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(b)What time would be taken for the blue colour to appear if the experiment was repeated using 4cm3 of distilled water and 6cm3 of solution A?(2mk)

From a correctly plotted graph

1/t at 6cm3 = 0.0125 √ **=>** **t = 1/0.0125 = 80seconds**√

2. You are provided with solid E. Carry out the experiments below. Write your observations and inferences in the spaces provided

(a)Place all solid in a boiling tube .Add 20cm3 of distilled water and shakeuntil all the solid dissolves.Label the solution as solution E.Use solution E for experiment (i)and (ii)

(i)To 2cm3 of solution E in a test tube in each of experiment I,II,III and IV add:

I. two drops of aqueous sodium sulphate(VI)

Observation(1mk)	**Inferences**(1mk)
White precipitate	**Pb^{2+} Ba^{2+} Ca^{2+}**

II.five drops of aqueous sodium chloride

Observations (1mk)	**Inferences**(1mk)

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White ppt	**Ca^{2+} Ba^{2+}**

~~III.two drops of barium chloride~~

Observations(1mk)	**Inferences(1mk)**
No white ppt	**SO_4^{2-} SO_3^{2-} CO_3^{2-}**

~~IV .~~two drops of lead(II)nitrate(V)

Observations(1mk)	**Inferences(1mk)**
No white ppt	**Cl^-**

~~(ii) To 2cm3 of solution E in a test tube ,add 5 drops of aqueous sodium~~ hydroxide .Add the piece of aluminium foil provided to the mixture and shake. Warm the mixture and test any gas produced with both blue and red litmus papers

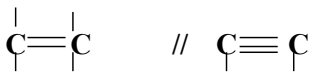
Observations(1mk)	**Inferences(1mk)**
Blue litmus paper remain blue	
Red limus paper turn blue	**NO_3^-**
Effervescence /fizzing/ bubbles	

Note:
Solid E is Calcium nitrate(V) / Barium nitrate(V)

3.You are provided with solid F. Carry out the following test.Write your observations and inferences in the spaces provided.
(a)Place all of solid F in a boiling tube. Add about 20cm3 of distilled water and shake until all the solid dissolves. Label the solution as solution F.
Add about half of the solid sodium hydrogen carbonate provided to 2cm3 of solution F

Observations(1mk)	**Inferences (1mk)**
No effervescence/fizzing	**H^+ absent**

(b)(i)Add about 10cm3 of dilute hydrochloric acid to the rest of solution F in the boiling tube. Filter the mixture. Wash the residue with about 2cm3 of distilled water.Dry the residue between filter papers. Place about one third of the dry residue on a metallic spatula and burn it in a Bunsen burner flame.

Observations(1mk)	**Inferences(1mk)**
Solid burns with a yellow sooty flame	 //

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(ii)Place all the remaining residue into a boiling tube. Add about 10cm3 of distilled water and shake thoroughly. **Retain** the mixture for the tests in (c)

Observations ($^1/_2$mk)	**Inferences($^1/_2$mk)**
	Solid
dissolves to a colourless solution	**Polar compound**

(c)~~Divide the mixture into two portions:~~

(i)to the first portion ,add the rest of the solid sodium hydrogen carbonate

Observations (1mk)	**Inferences(1mk)**
Effervescence/fizzing	**H^+**

(ii)to the second portion ,add two drops of bromine water.

Observations (1mk)	**Inferences(1mk)**
	Bromine
water decolorized	**C C C C bonds**

CHEMISTRY OF METALS

a)Introduction to metals

The rationale of studying metals cannot be emphasized.Since ages, the world over, metals like gold and silver have been used for commercial purposes.

The periodicity of alkali and alkaline earth metals was discussed in year 2 of secondary school education. This topic generally deals with:

(**a**)Natural occurrence of the chief ores of the most useful metals for industrial /commercial purposes.

(**b**)Extraction of these metals from their ores for industrial/ commercial purposes.

(**c**)industrial/ commercial uses of these metals.

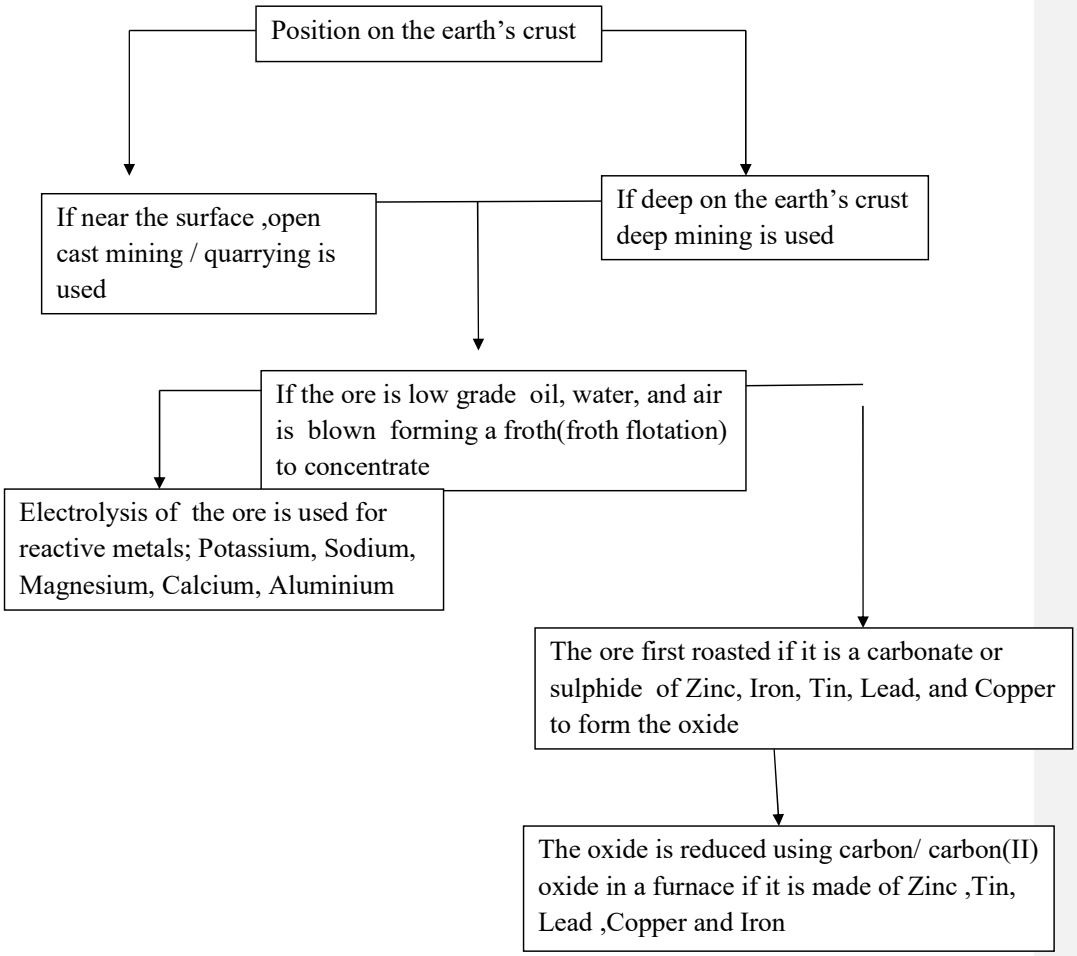
(**d**)main physical and chemical properties /characteristic of the metals.

The metals given detailed emphasis here are; **Sodium, Aluminium, Iron, Zinc, Lead** and **Copper.**

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The main criteria used in extraction of metals is based on its position in the electrochemical/reactivity series and its occurrence on the earth's crust.



1.**SODIUM**

a) Natural occurrence

Sodium naturally occurs as:

(i)Brine-a concentrated solution of sodium chloride(NaCl(**aq**)) in salty seas and oceans.

(ii)Rock salt-solid sodium chloride(NaCl(**s**)

(iii)Trona-sodium sesquicarbonate($NaHCO_3.Na_2CO_3.2H_2O$) especially in lake Magadi in Kenya.

(iv)Chile saltpeter-sodium nitrate($NaNO_3$)

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b)(i)

Extraction of Sodium from brine/Manufacture of Sodium hydroxide/The flowing mercury cathode cell/ TheCaster-Keller process

I.Raw materials

(i) Brine-concentrated solution of sodium chloride (NaCl (**aq**)) from salty seas and oceans.

(ii) Mercury

(iii) Water from river/lakes

II. Chemical processes

Salty lakes, seas and oceans contain large amount of dissolved sodium chloride (NaCl (aq)) solution.

This solution is concentrated to form brine which is fed into an electrolytic chamber made of **suspended** Carbon **graphite**/titanium as the **anode** and a **continuous** flow of Mercury as the **cathode**.Note

Mercury is the only naturally occurring known liquid metal at room temperature and pressure

Questions

I. Write the equation for the decomposition of the electrolyte during the electrolytic process.

$$\mathbf{H_2O(l) \longrightarrow H^+(aq) + OH^-(aq)}$$
$$\mathbf{NaCl(aq) \longrightarrow Na^+(aq) + Cl^-(aq)}$$

II. Name the ions present in brine that moves to the:

(i)Mercury cathode; $\mathbf{H^+(aq)}$, $\mathbf{Na^+(aq)}$

(ii)Titanium/graphite; $\mathbf{OH^-(aq)}$, $\mathbf{Cl^-(aq)}$

III. Write the equation for the reaction that take place during the electrolytic process at the;

Cathode; $\mathbf{2Na^+(aq) + 2e \longrightarrow 2Na(s)}$

Anode; $\mathbf{2Cl^-(aq) \longrightarrow Cl_2(g) + 2e}$

Note

(i) Concentration of $2Cl^-(aq)$ ions is higher than OH^- ions causing **overvoltage** thus **block**ing OH^- ions from being discharged at the anode.

(ii) Concentration of $Na^+(aq)$ ions is higher than H^+ ions causing **overvoltage** thus **block**ing H^+ ions from being discharged at the cathode.

IV. Name the products of electrolysis in the flowing mercury-cathode cell.

(i)Mercury cathode; **Sodium metal as grey soft metal/solid**

(ii)Titanium/graphite; **Chlorine gas as a pale green gas that turns moist blue/red litmus papers red then bleaches both.** Chlorine gas is a very useful by-product in;

(i)making (PVC)polyvinylchloride(polychloroethene) pipes.

(ii)chlorination/sterilization of water to kill germs.

(iii)bleaching agent

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(iv)manufacture of hydrochloric acid.

Sodium produced at the cathode immediately reacts with the mercury at the cathode forming **sodium amalgam(NaHg)** liquid that flow out of the chamber.

$$Na(s) + Hg(l) \longrightarrow Na\,Hg\,(l)$$

Sodium amalgam is added distilled water and reacts to form sodium hydroxide solution, free mercury and Hydrogen gas.

$$2Na\,Hg\,(l) + 2H_2O(l) \longrightarrow 2NaOH\,(aq) + 2Hg(l) + H_2(g)$$

Hydrogen gas is a very useful by-product in;

(i)making ammonia gas in the Haber process
(ii)manufacture of hydrochloric acid
(iii)in weather balloons to forecast weather
(iv)as rocket fuel

As the electrolysis of brine continues, the concentration of Cl-ions decreases and oxygen gas start being liberated. **Continuous** feeding of the electrolyte is therefore very necessary.

III.Uses of sodium hydroxide

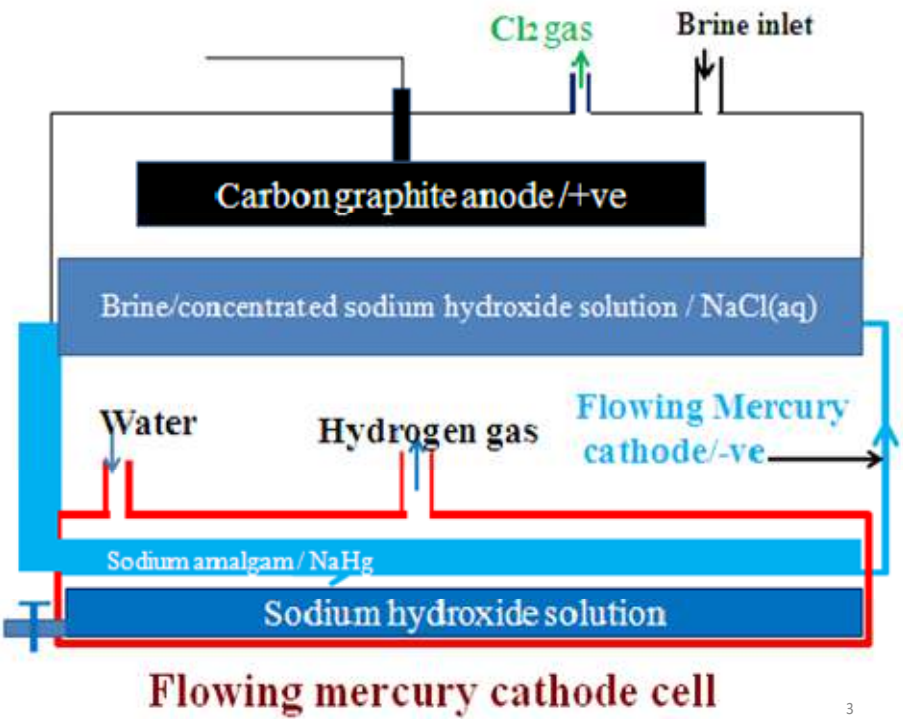
The sodium hydroxide produced is very pure and is used mainly in:

(i)Making soapy and soapless detergents.
(ii)making cellulose acetate/rayon

IV. Diagram showing the **Manufacture of Sodium hydroxide from the flowing Mercury-cathode cell.**

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Flowing mercury cathode cell

3

V. Environmental effects of Manufacture of Sodium hydroxide from the flowing Mercury-cathode cell.

1.Most of the Mercury used at the cathode is recycled ;

(i)to reduce the cost because mercury is expensive

(ii)to reduce pollution because mercury kills marine life.

(iii)because it causes chromosomal/genetic mutation to human beings.

2.Chlorine produced at the anode;

(i)has a pungent irritating smell that causes headache to human beings.

(ii)bleaches any wet substance.

(iii)dissolves water to form both hydrochloric acid and chloric(I)acid

Both cause marine pollution and stomach upsets.

b)(ii)

Extraction of sodium from rock salt/The Downs cell/process

I. Raw materials

(i)Rock salt/solid sodium chloride

(ii)calcium(II)chloride

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II. Chemical processes.

Rock salt/ solid sodium chloride is heated to molten state in a chamber lined with fire bricks on the outside.

Sodium chloride has a melting point of about 800°C. A little calcium (II) chloride is added to lower the melting point of the electrolyte to about 600°C.

The molten electrolyte is the electrolyzed in a carbon graphite anode suspended at the centre and surrounded by steel cathode.

Questions

I. Write the equation for the decomposition of the electrolyte during the electrolytic process.

$$NaCl(l) \longrightarrow Na^+(l) + Cl^-(l)$$

Note: In absence of water, the ions are in liquid state.

II. Name the ions present in molten rock salt that move to the;

(i)Steel cathode **-$Na^+(l)$**

(ii)Carbon graphite anode- **$Cl^-(l)$**

III. Write the equation for the reaction that take place during the electrolytic process at the;

(i)Steel cathode

$$2Na^+(l) + 2e \longrightarrow 2Na(l)$$

(ii)Carbon graphite anode

$$2Cl^-(l) \longrightarrow Cl_2(g) + 2e$$

IV. Name the products of electrolysis in the Downs cell at;

(i)Cathode:

Grey solid Sodium metal is less dense **than the molten electrolyte and therefore** float **on top of the cathode to be periodically tapped off.**

(ii)Anode:

Pale green chlorine gas that turns moist/damp/wet blue/red litmus papers red then bleaches/decolorizes both. Chlorine gas is again a very useful by-product in;

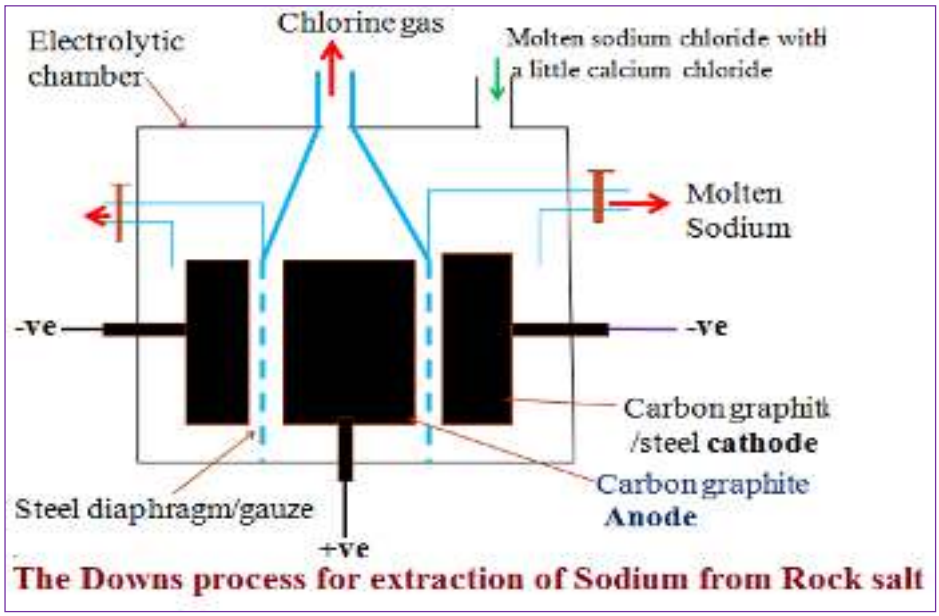
(i)making (PVC)polyvinylchloride(polychloroethene) pipes.
(ii)chlorination/sterilization of water to kill germs.
(iii)bleaching agent
(iv)manufacture of hydrochloric acid.

A **steel diaphragm/gauze** is suspended between the electrodes to **prevent recombination** of sodium at the cathode and chlorine gas at the anode **back** to sodium chloride.

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III. Diagram showing the **Downs cell/process for extraction of sodium**



The Downs process for extraction of Sodium from Rock salt

IV. Uses of sodium.

1.Sodium vapour is used as sodium lamps to give a yellow light in street lighting.

2.Sodium is used in making very useful sodium compounds like;

(i)Sodium hydroxide(NaOH)

(ii)Sodium cyanide(NaCN)

(iii)Sodium peroxide(Na_2O_2)

(iv)Sodamide($NaNH_2$)

3.An alloy of Potassium and Sodium is used as **coolant** in nuclear reactors.

V. Environmental effects of Downs cell.

1.Chlorine produced at the anode;

(i)has a pungent irritating smell that causes headache to human beings.

(ii)bleaches any wet substance.

(iii)dissolves water to form both hydrochloric acid and chloric(I)acid

Both cause marine pollution and stomach upsets.

2.Sodium metal rapidly react with traces of water to form alkaline Sodium hydroxide(NaOH(aq))solution. This raises the pH of rivers/lakes killing aquatic lifein case of leakages.

VI. Test for presence of Na.

If a compound has Na^+ ions in solid/molten/aqueous state then it changes a non-luminous clear/colourless flame to a **yellow coloration** but does not burn

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Experiment

Scoop a portion of sodium chloride crystals/solution in a clean metallic spatula. Introduce it to a clear /colourless Bunsen flame.

Observation	Inference
Yellow coloration	Na^+

Practice

(i)Calculate the time taken in hours for 230kg of sodium to be produced in the Downs cell when a current of 120kA is used.

(ii)Determine the volume of chlorine released to the atmosphere.

(Na=23.0),Faraday constant=96500C.I mole of a gas =24dm3 at r.t.p)

Working:

Equation at the cathode:

$2Na^+$ **(l)** + **2e** -> 2Na**(l)**

2 mole of electrons = 2 Faradays = 2 x 96500 C deposits a mass = molar mass of Na = 23.0g thus;

23.0 g -> 2 x 96500 C

(230 **x 1000**)g -> $\frac{230 \times 1000 \times 2 \times 96500}{23}$

= **1,930,000,000 / 1.93 x 10^9C**

Time(**t**) in seconds = $\frac{\text{Quantity of electricity}}{\text{Current(I) in amperes}}$

Substituting

$= \frac{1,930,000,000 / 1.93 \times 10^9 C}{120 \times 1000A}$

= 16,083,3333seconds / 268.0556 minutes

= 4.4676hours

Volume of Chlorine

Method 1

Equation at the anode:

$2Cl^-$ **(l)** -> $Cl_2(g)$ + **2e**

From the equation:

2 moles of electrons = 2 Faradays =2 x 96500C

2 x 96500C -> 24dm3

1,930,000,000 / 1.93 x 10^9C->$\frac{1,930,000,000 / 1.93x10^9 C \times 24}{2 \times 96500C}$

Volume of Chlorine = **240,000dm3 /2.4 x 10^5dm3**

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Method 2
Equation at the anode: $Cl^-(l) \rightarrow Cl_2(g) + \mathbf{2e}$
Mole ratio of products at Cathode: anode = 1:1
Moles of sodium at cathode $= \frac{(230 \times 1000)}{23} g = \mathbf{10,000 moles}$
10,000moles of Na = **10,000moles** moles of Chlorine
1 moles of Chlorine gas = 24000cm3
10,000moles of Chlorine- > 10000 x 24
=240,000dm3 / 2.4x 10^5dm3
Method 3
Equation at the anode: $Cl^-(l) \rightarrow Cl_2(g) + \mathbf{2e}$
Ratio of Faradays of products at Cathode: anode = 2:2
=> 2 x 96500C produce 24000cm3 of chlorine gas Then: 1,930,000,000 / 1.93 x 10^9C ->
$\frac{1,930,000,000}{1.93 \times 10^9 C} \times 24 = \frac{\mathbf{240,000dm3}}{2 \times 96500}$

(iij)The sodium metal produced was reacted with water to form 25000dm3 solution in a Caster-Keller tank.
(a)Calculate the concentration of the resulting solution in moles per litre.
(b)The volume of gaseous products formed at s.t.p(1 mole of gas =22.4 dm3 at s.t.p)
Chemical equation at Caster-Keller tank
$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$
Mole ratio Na:NaOH = 2 : 2 => 1:1
Moles Na =10000moles=10000moles of NaOH
25000dm3 ->10000moles of NaOH
1dm3 -> $\frac{10000 \times 1}{25000}$ = **0.4M / 0.4 moles/dm3**
Mole ratio Na: $H_2(g)$ = 2 : 1
Moles Na = 10000moles = **5000moles of H_2 (g)**
Volume of H_2 (g) = moles x molar gas volume at s.t.p
=> 5000moles x 22.4 dm3
=120,000dm3

(iv)The solution formed was further diluted with water for a titration experiment. 25.0 cm3 of the diluted solution required 20.0cm3 of 0.2M sulphuric(VI)acid for complete neutralization. Calculate the volume of water added to the diluted solution before titration.
Chemical equation
$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$
Moles ratio NaOH : H_2SO_4 = 2 : 1
Moles ratio H_2SO_4 = molarity x volume => 0.2M x 20

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1000 1000

$= 4.0 \times 10^{-3}$ **moles**

Moles NaOH = 2 x 4.0 x 10^{-3} moles= **8.0×10^{-3} moles**

Molarity of NaOH= Moles x 1000=> 8.0 x 10^{-3} moles x 1000

volume 25

$=$ **0.16 molesdm^{-3} /M**

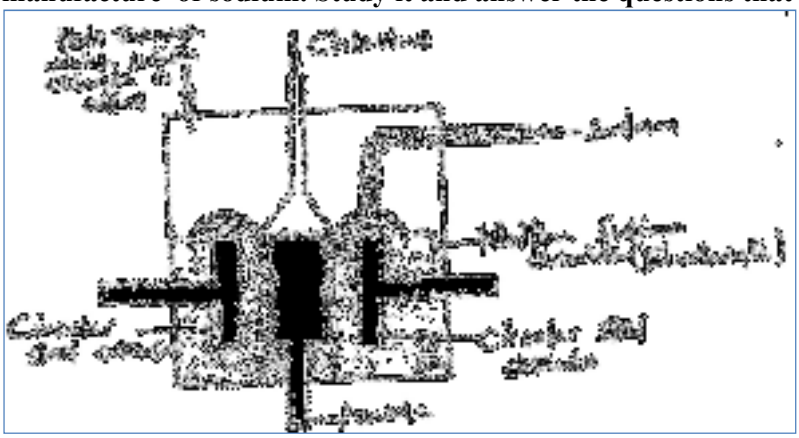
Volume used during dilution

$C_1V_1 = C_2V_2$ => 0.4M x V_1 = 0.16 M x 25

= 0.16 M x 25 = **10cm3**

0.4

(a) Below is a simplified diagram of the Downs Cell used for the manufacture of sodium. Study it and answer the questions that follow



(i)What material is the anode made of? Give a reason (2 mks)

Carbon graphite/Titanium

This because they are cheap and inert/do not influence/affect the products of electrolysis

(ii) What precaution is taken to prevent chlorine and sodium from re-combination? (1 mks)

Using a steel gauze/diaphragm separating the cathode from anode

(iii) Write an ionic equation for the reaction in which chlorine gas is formed (1mk)

$2Cl^-(l)$ -> $Cl_2(g)$ + 2e

(b) In the Downs process, (used for manufacture of sodium), a certain salt is added to lower the melting point of sodium chloride from about 800^{0}C to about 600^{0}C.

(i) Name the salt that is added (1mk)

Calcium chloride

(ii) State why it is necessary to lower the temperature(1mk)

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To reduce the cost of production

(c) Explain why aqueous sodium chloride is not suitable as an electrolyte for the manufacture of sodium in the Downs process(2mk)

The sodium produced react explosively/vigorously with water in the aqueous sodium chloride

(d) Sodium metal reacts with air to form two oxide. Give the formulae of two oxides (1mk)

Na_2O Sodium oxide(in limited air)

Na_2O_2 Sodium peroxide(in excess air)

2.ALUMINIUM

a)Natural occurrence

Aluminium is the most common naturally occurring metal. It makes 7% of the earths crust as:

(i)Bauxite ore- Hydrated aluminium oxide($Al_2O_3.2H_2O$)

(ii)Mica ore-Potassium aluminium silicate($K_2Al_2Si_6O_{16}$)

(iii)China clay ore- aluminium silicate ($Al_2Si_6O_{16}$)

(iv)Corrundum-Anhydrous aluminium oxide(Al_2O_3)

b)Extraction of aluminium from Bauxite/Halls cell/process)

The main ore from which aluminium is extracted is **Bauxite** ore- hydrated aluminium oxide($Al_2O_3.2H_2O$).

The ore is mined by open-caste mining method/quarrying where it is scooped together with silica/sand/silicon(IV)oxide (**SiO_2**) and soil/ iron(III)oxide (**Fe_2O_3**) as **impurities**.

The mixture is first dissolved in hot concentrated sodium/potassium hydroxide solution.

The alkalis dissolve both bauxite and silicon(IV)oxide.

This is because bauxite is **amphotellic** while silicon(IV)oxide is **acidic**.

Iron(III)oxide (**Fe_2O_3**) is filtered of /removed as a residue.

Carbon(IV)oxide is bubbled into the filtrate to precipitate aluminium (III) hydroxide ($Al(OH)_3$) as residue.

The aluminium (III) hydroxide (**$Al(OH)_3$**) residue is filtered off. Silicon (IV)oxide remain in the solution as filtrate. Aluminium (III) hydroxide ($Al(OH)_3$) residue is then heated to form pure aluminium (III)oxide(**Al_2O_3**)

$$2Al(OH)_3\ (s) \longrightarrow Al_2O_3\ (s) + 3H_2O(l)$$

Pure aluminium (III)oxide (**Al_2O_3**) has a very high melting point of **2015°C**.

Alot of energy is required to melt the oxide.

It is therefore dissolved first in molten **cryolite** /sodium hexafluoroaluminate (III)/Na_3AlF_6 to **lower** the melting point to about **800°C**.

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The molten electrolyte is put in the Hall cell made up of a steel tank lined with carbon graphite and an anode suspended into the electrolyte.
During the electrolysis:

(**i**)At the cathode;

$4Al^{3+}(\mathbf{l}) + 12e \longrightarrow 4Al(\mathbf{l})$

(**ii**) At the anode;

$6O^{2-}(l) \longrightarrow 3O_2(g) + 12e$

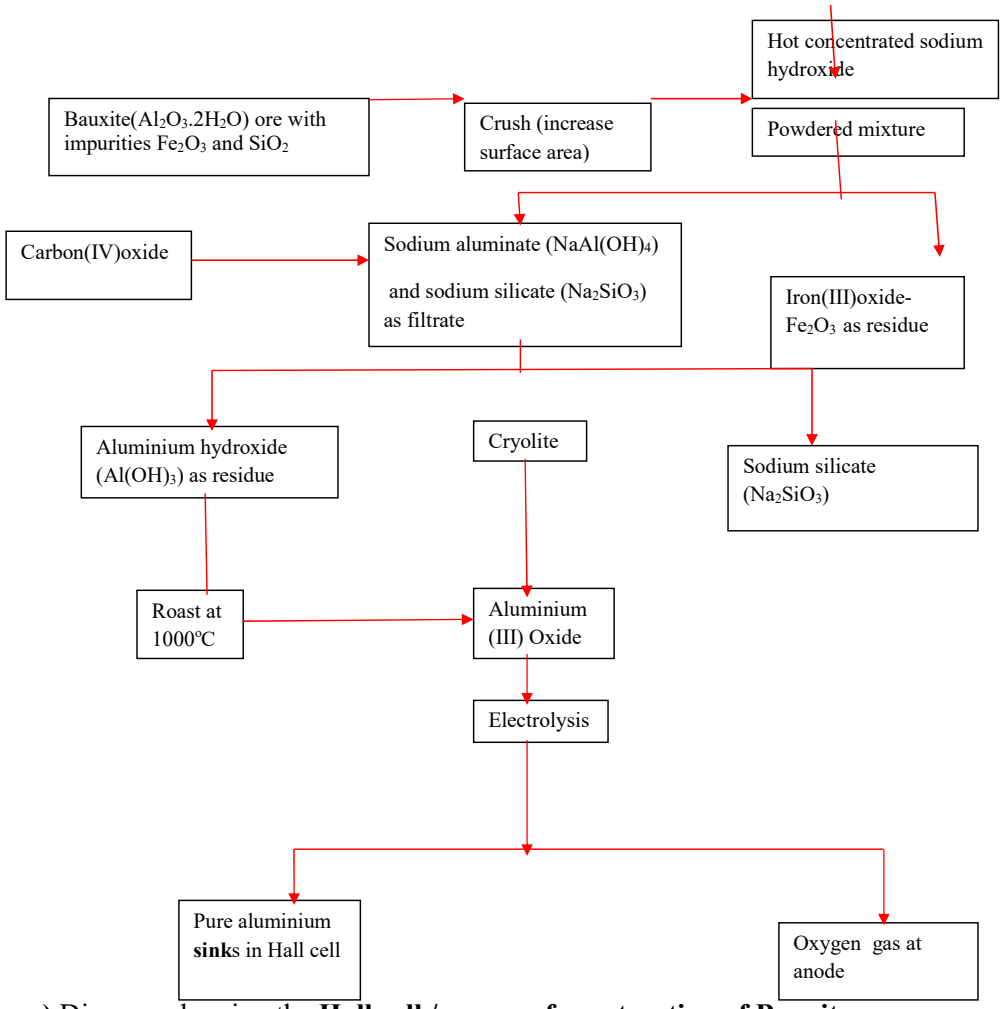
Aluminium is denser than the electrolyte therefore sink to the bottom of the Hall cell.
At this temperature ,the Oxygen evolved/produced at the anode **reacts** with carbon anode to form carbon(IV)oxide gas that **escape** to the atmosphere.

$C(s) + O_2(g) \longrightarrow CO_2(g)$

The anode thus should be continuously **replaced** from time to time.
Flow chart summary of extraction of aluminium from Bauxite

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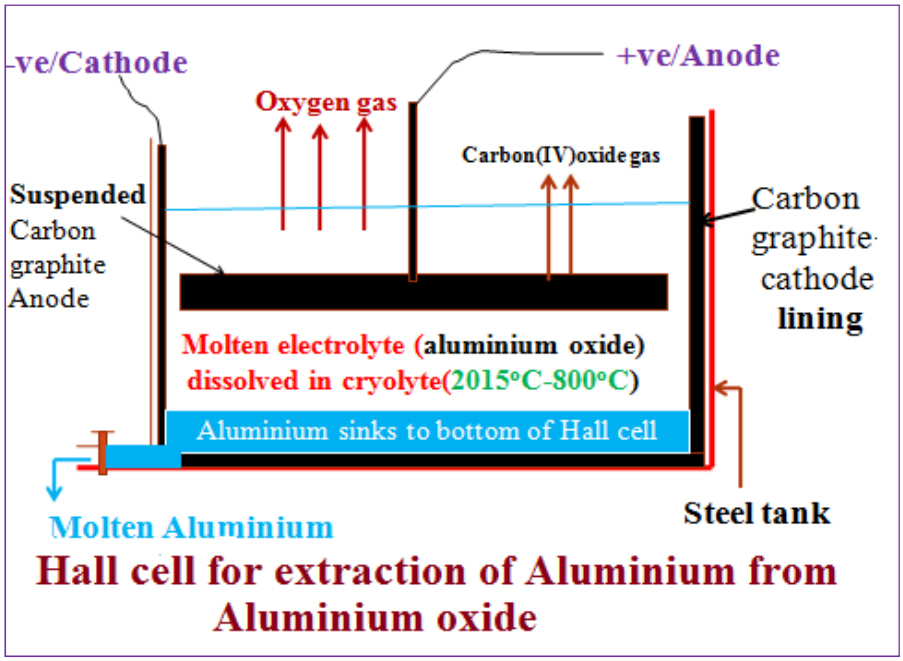
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c) Diagram showing the **Hall cell / process for extraction of Bauxite**

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d)Uses of aluminium

(i) In making aeroplane parts, buses, tankers, furniture because aluminium is very light.

(ii)Making duralumin-an alloy which is harder and has a higher tensile strength

(iii)Making utensils,sauce pans,spoons because it is light and good conductor of electricity.

(iv)Making overhead electric cables because it is light,ductile and good conductor of electricity.

(iv)Used in the thermite process for production of Manganese, Chromium amd Titanium.

e) Environmental effects of extracting aluminium from Bauxite.

Carbon(IV)oxide gas that **escape** to the atmosphere is a green house gas that causes global warming.

Bauxite is extracted by open caste mining that causes soil/environmental degradation.

f) Test for presence of Al^{3+}

If an ore is suspected to contain Al^{3+} it is;

(i)added hot concentrated sulphuric(VI)/Nitric(V)acid to free the ions present.

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(ii)the free ions are then added a precipitating reagent like 2M sodium hydroxide /2M aqueous ammonia.

Observation	Inference
White precipitate in excess 2M NaOH(aq)	Pb^{2+}, Al^{3+}, Zn^{2+}
White precipitate in excess 2M NH_3(aq)	Pb^{2+}, Al^{3+}
No black precipitate on adding Na_2S(aq)	Al^{3+}
No white precipitate on adding either NaCl(aq),HCl(aq),H_2SO_4(aq),Na_2SO_4(aq)	Al^{3+}

Practice

1.An unknown rock X was discovered in Ukraine. Test with dilute sulphuric (VI)acid shows rapid effervescence with production of a colourless gas A that forms a white precipitate with lime water and colourless solution B. On adding 3cm3 of 2M sodium hydroxide, a white precipitate C is formed that dissolves to form a colourless solution D on adding more sodium hydroxide. On adding 2M aqueous ammonia, a white precipitate E is formed which persist in excess aqueous ammonia.On which on adding 5cm3 of 1M Lead(II)nitrate(V) to F a white precipitate G is formed which remains on heating.

Identify:

A **Hydrogen/H_2**

B **Aluminium sulphate(VI)/$Al_2(SO_4)_3$**

C **Aluminium hydroxide/ $Al(OH_4)_3$**

D **Tetrahydroxoaluminate(III)/ $[Al(OH_4)_3]^-$**

E **Aluminium hydroxide/ $Al(OH)_3$**

F **Aluminium chloride/ $AlCl_3$**

2.Aluminium is obtained from the ore with the formula Al_2O_3. $2H_2O$. The ore is first heated and refined to obtain pure aluminium oxide (Al_2O_3). The oxide is then electrolysed to get Aluminium and oxygen gas using carbon anodes and carbon as cathode. Give the common name of the ore from where aluminium is extracted from ½ mark

What would be the importance of heating the ore first before refining it?1 mark

To remove the water of crystallization

The refined ore has to be dissolved in cryolite first before electrolysis. Why is this necessary? 1½ mark

To lower the melting point of aluminium oxide from about 2015°C to 900°C so as to lower /reduce cost of production

Why are the carbon anodes replaced every now and then in the cell for electrolysing aluminium oxide? 1 mark

Oxygen produced at anode react with carbon to form carbon(IV)oxide gas that escape

State two uses of aluminium

In making aeroplane parts, buses, tankers, utensils, sauce pans,spoons

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Making overhead electric cables
Making duralumin

3. IRON

a)Natural occurrence

Iron is the second most common naturally occurring metal. It makes 4% of the earths crust as:

(i)Haematite(Fe_2O_3)
(ii)Magnetite(Fe_3O_4)
(iii)Siderite($FeCO_3$)

b)The blast furnace for extraction of iron from Haematite and Magnetite

a)Raw materials:

(i)Haematite(Fe_2O_3)
(ii)Magnetite(Fe_3O_4)
(iii)Siderite($FeCO_3$)
(iv)Coke/charcoal/ carbon
(v)Limestone

b)Chemical processes:

Iron is usually extracted from Haematite (Fe_2O_3), Magnetite(Fe_3O_4) Siderite ($FeCO_3$).These ores contain silicon(IV)oxide(SiO_2) and aluminium(III)oxide (Al_2O_3) as impurities.

When extracted from siderite, the ore must first be roasted in air to decompose the iron(II)Carbonate to Iron(II)oxide with production of carbon(IV)oxide gas:

$$FeCO_3(s) \longrightarrow FeO(s) + CO_2(g)$$

Iron(II)oxide is then rapidly oxidized by air to iron(III)oxide(Haematite).

$$4FeO(s) + O_2(g) \longrightarrow 2Fe_2O_3(s)$$

Haematite (Fe_2O_3), Magnetite(Fe_3O_4), coke and limestone are all then fed from top into a tall (about 30metres in height) tapered steel chamber lined with refractory bricks called a blast furnace.

The furnace is covered with **inverted** double cap to prevent/reduce amount of any gases escaping .

Near the base/bottom, blast of hot air at about 1000K (827ºC) is driven/forced into the furnace through small holes called **Tuyeres**.

As the air enters ,it reacts with coke/charcoal/carbon to form carbon(IV)oxide gas. This reaction is highly exothermic.

$$C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -394kJ$$

This raises the temperature at the bottom of the furnace to about 2000K(1650ºC).As Carbon(IV)oxide gas rises up the furnace it reacts with more coke to form carbon(II)oxide gas.This reaction is endothermic.

$$CO_2(g) + C(s) \longrightarrow 2CO(g) \quad \Delta H = +173kJ$$

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Carbon(II)oxide gas is a strong reducing agent that reduces the ores at the upper parts of the furnace where temperatures are about 750K(500°C) i.e.

For Haematite;

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + CO_2(g)$$

For Magnetite;

$$Fe_3O_4(s) + 4CO(g) \longrightarrow 3Fe(s) + 4CO_2(g)$$

Iron is denser than iron ore. As it falls to the hotter base of the furnace it melts and can easily be tapped off.

Limestone fed into the furnace decomposes to quicklime/calcium oxide and produce more carbon(IV)oxide gas.

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

Quicklime/calcium oxide reacts with the impurities silicon(IV)oxide(SiO_2) and aluminium(III)oxide(Al_2O_3)in the ore to form calcium silicate and calcium aluminate.

$$CaO(s) + SiO_2(s) \longrightarrow CaSiO_3(l)$$
$$CaO(s) + Al_2O_3(s) \longrightarrow Ca\,Al_2O_4(l)$$

Calcium silicate and calcium aluminate mixture is called **slag**.Slag is **denser** than iron ore but **less dense** than iron therefore **float** on the pure iron. It is tapped at different levels to be tapped off for use in:

(i)tarmacing roads
(ii) cement manufacture
(iii)as building construction material

(c)Uses of Iron

Iron obtained from the blast furnace is hard and brittle. It is called **Pig iron.** It is remelted, added scrap steel then cooled. This iron is called cast iron.

Iron is mainly used to make:

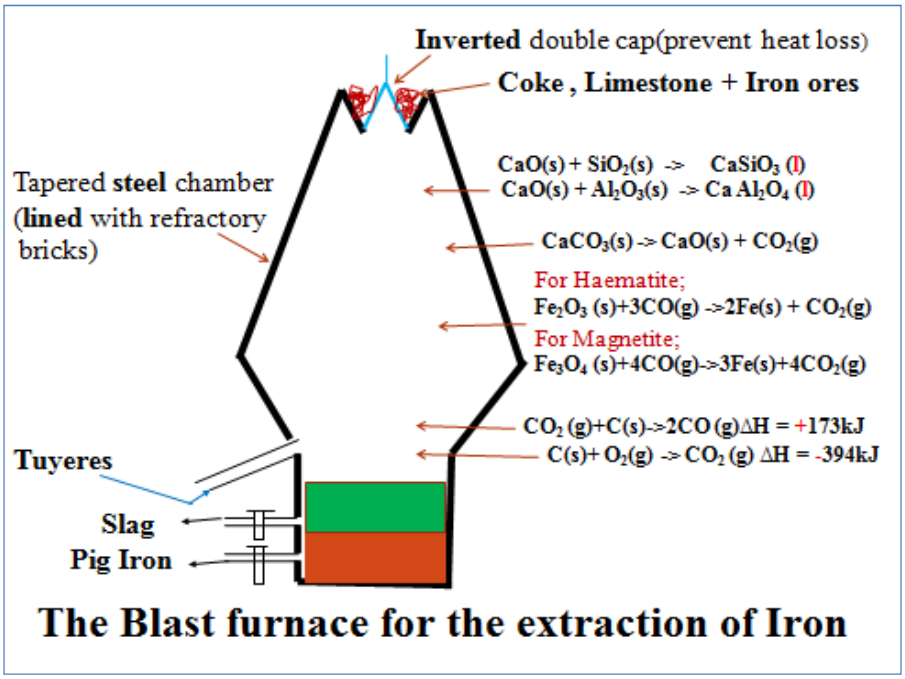
(i)gates ,pipes, engine blocks, rails, charcoal iron boxes,lamp posts because it is cheap.

(ii)nails, cutlery, scissors, sinks, vats, spanners,steel rods, and railway points from steel.

Steel is an alloy of iron with carbon, and/or Vanadium, Manganese, Tungsten, Nickel ,Chromium. It does **not rust**/corrode like iron.

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The Blast furnace for the extraction of Iron

e) Environmental effects of extracting Iron from Blast furnace

(i)Carbon(IV)oxide(CO_2) gas is a green house gas that causes/increases global warming if allowed to escape/leak from the furnace.

(ii)Carbon(II)oxide(CO)gas is a highly poisonous/toxic odourless gas that can kill on leakage.

It is **preferentially** absorbed by the haemoglobin in mammals instead of Oxygen to form a **stable** compound that reduce free hemoglobin in the blood.

(iii) Haematite (Fe_2O_3), Magnetite(Fe_3O_4) and Siderite ($FeCO_3$) are extracted through quarrying /open cast mining that cause soil / environmental degradation .

f) Test for the presence of Iron

Iron naturally exist in its compound as Fe^{2+} /Fe^{3+}

If an ore is suspected to contain Fe^{2+} /Fe^{3+} it is;

(i)added hot concentrated sulphuric(VI)/Nitric(V)acid to free the ions present.

(ii)the free ions are then added a precipitating reagent like 2M sodium hydroxide /2M aqueous ammonia which forms;

I) an insoluble green precipitate in excess of 2M sodium hydroxide /2M aqueous ammonia if Fe^{2+} ions are present.

I) an insoluble brown precipitate in excess of 2M sodium hydroxide /2M aqueous ammonia if Fe^{2+} ions are present.

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Observation	Inference
green precipitate in excess 2M NaOH(aq)	Fe^{2+}
green precipitate in excess 2M NH_3(aq)	Fe^{2+}
brown precipitate in excess 2M NaOH(aq)	Fe^{3+}
brown precipitate in excess 2M NH_3(aq)	Fe^{3+}

4. COPPER

a)Natural occurrence

Copper is found as uncombined element/metal on the earths crust in Zambia, Tanzania, USA and Canada .The chief ores of copper are:

(i)Copper pyrites($CuFeS_2$)
(ii)Malachite($CuCO_3.Cu(OH)_2$)
(iii)Cuprite(Cu_2O)

b)Extraction of copper from copper pyrites.

Copper pyrites are first crushed into fine powder. The powdered ore is the added water and oil. The purpose of water is to dissolve hydrophilic substances/particle. The purpose of oil is to make cover copper ore particle so as to make it hydrophobic

Air is blown through the mixture. Air creates bubbles that stick around hydrophobic copper ore. The air bubbles raise through buoyancy small hydrophobic copper ore particles to the surface. A concentrated ore floats at the top as froth. This is called froth flotation. The concentrated ore is then skimmed off.The ore is then roasted in air to form copper(I)sulphide ,sulphur(IV)oxide and iron (II) oxide.

$$2CuFeS_2(s) + 4O_2(g) \longrightarrow Cu_2S(s) + 3SO_2(g) + 2FeO(s)$$

Limestone ($CaCO_3$) and silicon(IV)oxide (SiO_2) are added and the mixture heated in absence of air.Silicon(IV)oxide (SiO_2) reacts with iron (II) oxide to form Iron silicate which constitutes the slag and is removed.

$$FeO(s) + SiO_2(s) \longrightarrow FeSiO_3(s)$$

The slag separates off from the copper(I)sulphide. Copper(I)sulphide is then heated in a regulated supply of air where some of it is converted to copper (I) oxide.

$$2Cu_2S(s) + 3O_2(g) \longrightarrow 2Cu_2S(s) + 2SO_2(g)$$

The mixture then undergo self reduction in which copper(I)oxide is reduced by copper(I)sulphide to copper metal.

$$Cu_2S(s) + 2Cu_2O(s) \longrightarrow 6Cu(s) + SO_2(g)$$

The copper obtained has Iron, sulphur and traces of silver and gold as impurities.It is therefore about 97.5% pure. It is refined by electrolysis/electrolytic method. During the electrolysis of refining copper, the impure copper is made the anode and a small pure strip is made the cathode.

Electrode ionization takes place where:

At the anode;

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$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e$

Note: Impure copper anode dissolves/erodes into solution and decreases in size.

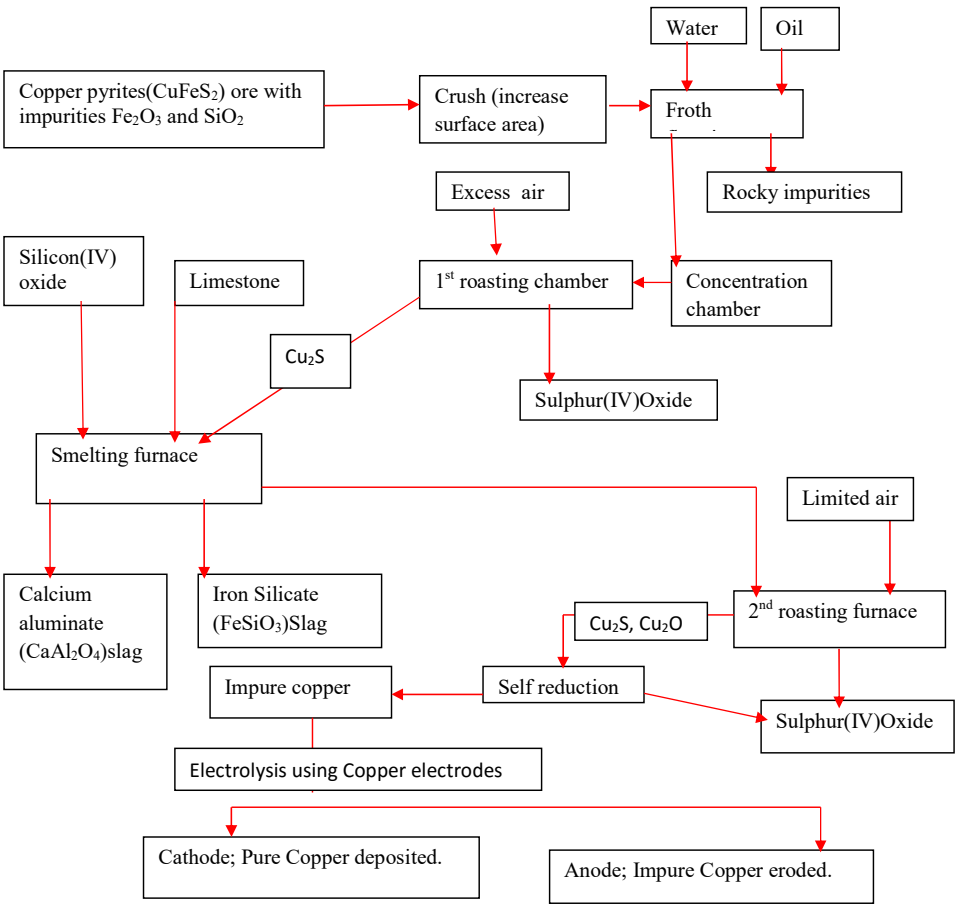
At the Cathode;

$Cu^{2+}(aq) + 2e \longrightarrow Cu(s)$

Note: The copper ions in the electrolyte($CuSO_4$) are reduced and deposited as copper metal at the cathode. The copper obtained is 99.98% pure.

Valuable traces of silver and gold collect at the bottom of the electrolytic cell as sludge. It is used to finance the extraction of copper pyrites.

(c)Flow chart summary of extraction of copper from Copper pyrites



Electrolytic purification of impure copper

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d) Uses of copper

Copper is mainly used in:

(i)making low voltage electric cables,contact switches, cockets and plugs because it is a good conductor of electricity.

(ii)Making solder because it is a good thermal conductor.

(iii)Making useful alloys e.g.

-Brass is an alloy of copper and Zinc(Cu/Zn)

-Bronze is an alloy of copper and Tin(Cu/Sn)

-German silver is an alloy of copper ,Zinc and Nickel(Cu/Zn/Ni)

(iv)Making coins and ornaments.

e) Environmental effects of extracting copper from Copper pyrites

(i)Sulphur(IV)oxide is a gas that has a pungent poisonous smell that causes head ache to human in high concentration.

(ii)Sulphur(IV)oxide gas if allowed to escape dissolves in water /rivers/rain to form weak sulphuric(IV)acid lowering the pH of the water leading to marine pollution, accelerated corrosion/rusting of metals/roofs and breathing problems to human beings.

(iii)Copper is extracted by open caste mining leading to land /environmental /soil degradation.

f) Test for the presence of copper in an ore

Copper naturally exist in its compound as Cu^{2+} /Cu^{+}

Copper (I) / Cu^{+} is readily oxidized to copper(II)/ Cu^{2+}

If an ore is suspected to contain Cu^{2+} /Cu^{+} it is;

(i)added hot concentrated sulphuric(VI)/Nitric(V)acid to free the ions present.

(ii)the free ions are then added a precipitating reagent; 2M sodium hydroxide /2M aqueous ammonia which forms;

I) an insoluble blue precipitate in excess of 2M sodium hydroxide if Cu^{2+} ions are present.

I) an insoluble blue precipitate in 2M aqueous ammonia that dissolve to **royal/deep blue solution** in excess if **Cu^{2+}** ions are present.

Observation	Inference
blue precipitate in excess 2M NaOH(aq)	Cu^{2+}
blue precipitate,dissolve to **royal/deep blue solution** in excess 2M NH_3(aq)	**Cu^{2+}**

g)Sample questions

Copper is extracted from copper pyrites as in the flow chart outlined below. Study it and answer the questions that follow

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5.**ZINC and LEAD**

a)Natural occurrence

Zinc occurs mainly as:

(i)Calamine-Zinc carbonate($ZnCO_3$)

(ii)Zinc blende-Zinc sulphide(ZnS)

Lead occurs mainly as Galena-Lead(II)Sulphide mixed with Zinc blende:

b)Extraction of Zinc/Lead from Calamine ,Zinc blende and Galena.

During extraction of Zinc , the ore is first roasted in air:

For Calamine Zinc carbonate decompose to Zinc oxide and carbon(IV) oxide gas.

$ZnCO_3(s) \longrightarrow ZnO(s) + CO_2(g)$

Zinc blende does not decompose but reacts with air to form Zinc oxide and sulphur(IV) oxide gas.

Galena as a useful impurity also reacts with air to form Lead(II) oxide and sulphur(IV) oxide gas.

$2ZnS(s) + 3O_2(g) \longrightarrow 2ZnO(s) + 2SO_2(g)$

(Zinc blende)

$2PbS(s) + 3O_2(g) \longrightarrow 2PbO(s) + 2SO_2(g)$

(Galena)

The oxides are mixed with coke and limestone/Iron(II)oxide/ Aluminium (III) oxide and heated in a blast furnace.

At the furnace temperatures limestone decomposes to quicklime/CaO and produce Carbon(IV)oxide gas.

$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$

Carbon(IV)oxide gas reacts with more coke to form the Carbon(II)oxide gas.

$C(s) + CO_2(g) \longrightarrow 2CO(g)$

Both Carbon(II)oxide and carbon/coke/carbon are reducing agents.

The oxides are reduced to the metals by either coke or carbon (II)oxide.

$ZnO(s) + C(s) \longrightarrow Zn(\mathbf{g}) + CO(g)$

$PbO(s) + C(s) \longrightarrow Pb(\mathbf{l}) + CO(g)$

$PbO(s) + CO(s) \longrightarrow Pb(\mathbf{l}) + CO_2(g)$

$PbO(s) + CO(s) \longrightarrow Pb(\mathbf{g}) + CO_2(g)$

At the furnace temperature:

(i)Zinc is a gas/vapour and is collected at the **top** of the furnace. It is condensed in a spray of molten lead to prevent reoxidation to Zinc oxide. On further cooling , Zinc collects on the surface from where it can be tapped off

(ii)Lead is a liquid and is ale to trickle to the bottom of the furnace from where it is tapped off.

Quicklime/CaO, Iron(II)Oxide, Aluminium(III)oxide are used to remove silica/silicon(IV)oxide as silicates which float above Lead preventing its reoxidation back to Lead(II)Oxide.

$CaO(s) + SiO_2(s) \longrightarrow CaSiO_3(s/l)$

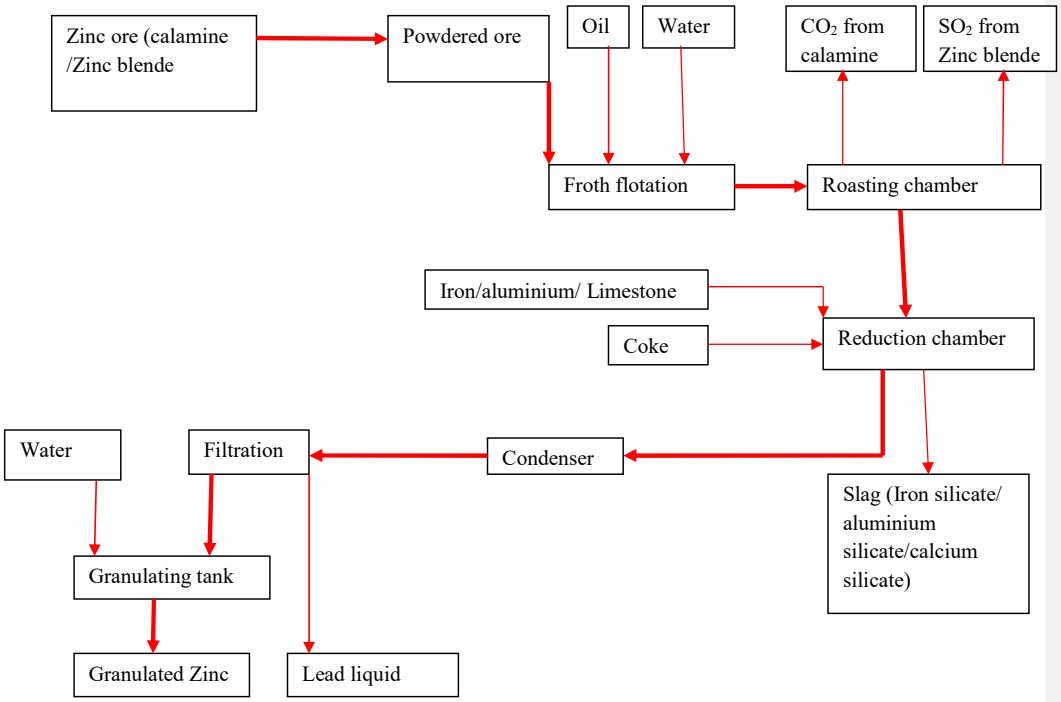
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$FeO(s) + SiO_2(s) \longrightarrow FeSiO_3(s/l)$ (Slag-Calcium silicate)

$Al_2O_3(s) + SiO_2(s) \longrightarrow Al_2SiO_4(s/l)$ (Slag-Aluminium silicate)

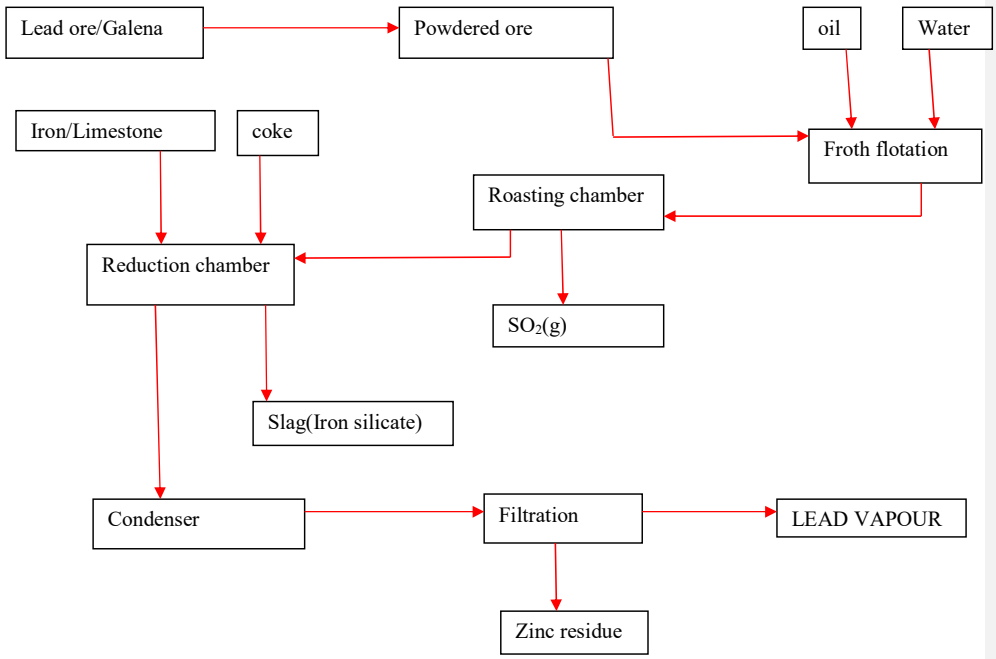
c)Flow chart on extraction of Zinc from Calamine ,Zinc blende.



d) Flow chart on extraction of Lead from Galena

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e) Uses of Lead

Lead is used in:

(i)making gun-burettes.

(ii)making protective clothes against nuclear (alpha rays/particle) radiation in a nuclear reactor.

(iii)Mixed with tin(Sn) to make **solder** alloy

f) Uses of Zinc

Zinc is used in:

(i)Galvanization-when iron sheet is dipped in molten Zinc ,a thin layer of Zinc is formed on the surface.Since Zinc is more reactive than iron ,it reacts with elements of air(CO_2/ O_2 / H_2O) to form basic Zinc carbonate($ZnCO_3.Zn(OH)_2$).This **sacrificial** method protects iron from corrosion/rusting.

(ii)As negative terminal and casing in dry/Laclanche cells.

(iii)Making brass alloy with copper(Cu/Zn)

g) Environmental effects of extracting Zinc and Lead.

(i) Lead and Lead salts are carcinogenic/causes cancer

(ii)Carbon(IV)oxide is a green house gas that causes/accelerate global warming.

(iii)Carbon(II)oxide is a colourless odourless poisonous /toxic gas that combines with haemoglobin in the blood to form stable carboxyhaemoglobin reducing free haemoglobin leading to death.

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(iv) Sulphur(IV)oxide is a gas that has a pungent poisonous smell that causes headache to human if in high concentration.
(v)Any leakages in Sulphur(IV)oxide gas escapes to the water bodies to form weak sulphuric(VI)acid lowering the pH of the water. This causes marine pollution /death of aquatic life, accelerated rusting/corrosion of metals/roofs and breathing problems to human beings.

h) Test for presence of Zinc/ Lead.

If an ore is suspected to contain Zinc/Lead it is:
I.added hot concentrated Nitric(V)acid to free the ions present.
Note:
Concentrated Sulphuric(VI)acid forms insoluble $PbSO_4$ thus cannot be used to free the ions in Lead salts.
II.the free ions are then added a precipitating reagent mostly 2M sodium hydroxide or 2M aqueous ammonia with the formation of;

(i)a soluble precipitate in excess of 2M sodium hydroxide if Zn^{2+}, Pb^{2+}, Al^{3+}ions are present.

(ii)a white precipitate that dissolves to form a colorless solution in excess 2M aqueous ammonia if Zn^{2+}ions are present.

(iii)an insoluble white precipitate in excess 2M aqueous ammonia if Pb^{2+}, Al^{3+}ions are present.

(iv) Pb^{2+} ions form a white precipitate when any soluble SO_4^{2-}, SO_3^{2-}, CO_3^{2-}, Cl^-, is added while Al^{3+} ions do not form a white precipitate

(v) Pb^{2+} ions form a yellow precipitate when any soluble I^-(e.g. Potassium/sodium Iodide) is added while Al^{3+} ions do not form a yellow precipitate.

(vi) Pb^{2+} ions form a black precipitate when any soluble S^-(e.g. Potassium/sodium sulphide) is added while Al^{3+} ions do not form a black precipitate.i.e;

Observation	Inference
White precipitate in excess 2M NaOH (aq)	Zn^{2+}, Pb^{2+}, Al^{3+} ions
White precipitate that dissolves to form a colourless solution in excess 2M NH_3(aq)	Zn^{2+} ions
White precipitate in excess 2M NH_3(aq)	Pb^{2+}, Al^{3+} ions
White precipitate on adding about 4 drops of either Na_2CO_3(aq), Na_2SO_3(aq), Na_2SO_4(aq), H_2SO_4(aq), HCl(aq), NaCl(aq)	Pb^{2+}ions
Yellow precipitate on adding about 4 drops of of KI(aq).NaI (aq)	Pb^{2+} ions
Black precipitate on adding aout 4 drops of Na_2S(aq)/K_2S(aq)	Pb^{2+} ions

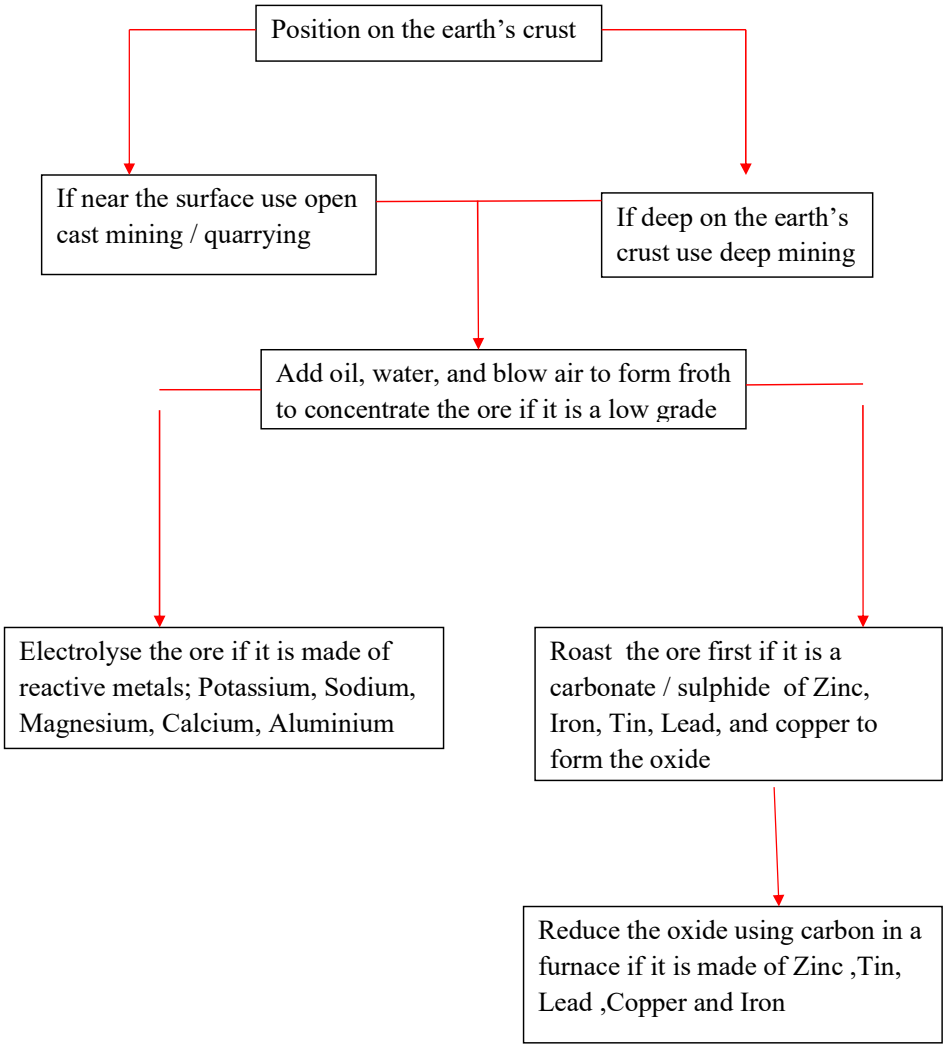
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6.GENERAL SUMMARY OF METALS

a) Summary methods of extracting metal from their ore

The main criteria used in extraction of metals is based on its position in the electrochemical/reactivity series and its occurrence on the earth's crust.



b) Summary of extraction of common metal.

Metal	Chief ore/s	Chemical formula of ore	Method of extraction	Main equation during extraction

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Sodium	Rock salt	NaCl(s)	**Downs process** Through electrolysis of molten NaCl ($CaCl_2$ lower m.pt from 800°C-> 600°C)	**Cathode:** $2Na^+(l) + 2e \rightarrow 2Na(l)$ **Anode:** $2Cl^-(l) \rightarrow Cl_2(g) + 2e$
Sodium/ sodium hydroxide	Brine	NaCl(aq)	**Flowing mercury cathode cell** Through electrolysis of concentrated NaCl(aq)	**Cathode:** $2Na^+(aq)+2e \rightarrow 2Na(aq)$ **Anode:** $2Cl^-(aq) \rightarrow Cl_2(g) + 2e$
Aluminium	Bauxite	$Al_2O_3.2H_2O$	**Halls process Through electrolysis of molten** Al_2O_3. (Cryolite lower m.pt from 2015°C -> 800°C)	**Cathode:** $4Al^{3+}(l) + 12e \rightarrow 4Al(l)$ **Anode:** $6O^{2-}(l) \rightarrow 3O_2(g) + 12e$
Iron	Haematite Magnetite	Fe_2O_3 Fe_3O_4	**Blast furnace** Reduction of the ore by carbon(II)oxide	$Fe_2O_3(s)+ 3CO(g) \rightarrow 2Fe(l) +3CO_2(g)$ $Fe_3O_4(s)+ 4CO(g) \rightarrow 3Fe(l) +4CO_2(g)$
Copper	Copper pyrites	$CuFeS_2$	**Roasting** the ore in air to get Cu_2S. **Heating** Cu_2S ore in regulated supply of air. **Reduction** of Cu_2O by Cu_2S	$2CuFeS_2 (s)+ 4O_2(g) \rightarrow Cu_2S(s)+3SO_2(g) +2FeO(s)$ $2Cu_2S (s)+ 3O_2(g) \rightarrow 2Cu_2O(s)+2SO_2(g)$ $Cu_2S (s)+ 2Cu_2O(s) \rightarrow 6Cu(s)+ SO_2(g)$
Zinc	Calamine	$ZnCO_3$	**Roasting** the ore in air	$ZnCO_3(s) \rightarrow$

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			to get ZnO **Blast furnace /reduction** of the oxide by Carbon(II)Oxide/Carbon	$ZnO(s) + CO_2(g)$ $2ZnS(s) + 3O_2(g) -> 2ZnO(s) + 2SO_2(g)$ $ZnO(s) + CO(g) -> Zn(s) + CO_2(g)$
Lead	Galena	PbS	**Blast furnace-Reduction of the oxide by carbon(II)oxide /Carbon**	$PbO(s) + CO(g) -> Pb(s) + CO_2(g)$

of metal.

Alloy name	**Constituents of the alloy**	**Uses of the alloy**
Brass	Copper and Zinc	Making scews and bulb caps
Bronze	Copper and Tin	Making clock springs,electrical contacts and copper coins
Soldier	Lead and Tin	Soldering, joining electrical contacts because of its low melting points and high thermal conductivity
Duralumin	Aluminium, Copper and Magnesium	Making aircraft , utensils ,windows frames because of its light weight and corrosion resistant.
Steel	Iron, Carbon ,Manganese and other metals	Railway lines , car bodies girders and utensils.
Nichrome	Nichrome and Chromium	Provide resistance in electric heaters and ovens
German silver	Copper,Zinc and Nickel	Making coins

d) Physical properties of metal.

Metals form giant metallic structure joined by metallic bond from electrostatic attraction between the metallic cation and free delocalized electrons.

This makes metals to have the following physical properties:

(i)High melting and boiling points

The giant metallic structure has a very close packed metallic lattice joined by strong electrostatic attraction between the metallic cation and free delocalized electrons.The more delocalized electrons the higher the melting/boiling points e.g.

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Aluminium has a melting point of about 2015°C while that of sodium is about 98°C.This is mainly because aluminium has more/three delocalized electrons than sodium/has one.
Aluminium has a boiling point of about 2470°C while that of sodium is about 890°C.This is mainly because aluminium has more/three delocalized electrons than sodium/has one.

(ii)High thermal and electrical conductivity

All metals are good thermal and electrical conductors as liquid or solids. The more delocalized electrons the higher the thermal and electrical conductivity. e.g.
Aluminium has an electrical conductivity of about 3.82×19^{-9} ohms per metre.
Sodium has an electrical conductivity of about 2.18×19^{-9} ohms per metre.

(iii)Shiny/Lustrous

The free delocalized electrons on the surface of the metal absorb, vibrate and then scatter/re-emit/lose light energy. All metals are therefore usually shades of grey in colour except copper which is shiny brown.e.g.
Zinc is bluish grey while iron is silvery grey.

(iv)High tensile strength

The free delocalized electrons on the surface of the metal atoms binds the surface immediately when the metal is coiled/folded preventing it from breaking /being brittle.

(v)Malleable.

Metals can be made into thin sheet. The metallic crystal lattice on being beaten/pressed/hammered on two sides extend its length and width/bredth and is then immediately bound by the delocalized electrons preventing it from breaking/being brittle.

(vi)Ductile.

Metals can be made into thin wires. The metallic crystal lattice on being beaten/pressed/hammered on all sides extend its length is then immediately bound by the delocalized electrons preventing it from breaking/being brittle.

Revision questions

1.Given some soil , dilute sulphuric(VI)acid,mortar,pestle,filter paper,filter funnel and 2M aqueous ammonia,describe with explanation,how you would show that the soil contain Zinc.

Place the soil sample in the pestle. Crush using the mortar to reduce the particle size/increase its surface area.
Add dilute sulphuric(VI)acid to free the ions in soil sample.
Filter to separate insoluble residue from soluble filtrate
To filtrate,add three drops of aqueous ammonia as precipitating reagent. A white precipitate of $Zn(OH)_2$, $Pb(OH)_2$ or $Al(OH)_3$ is formed

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Add excess aqueous ammonia to the white precipitate. If it dissolves the Zn^{2+} ions are present. $Zn(OH)_2$ react with excess ammonia to form soluble $[Zn(OH)_4]^{2+}$ complex.

2.In the extraction of aluminium,the oxide is dissolved in cryolite.

(i)What is the chemical name of cryolite?

Sodium hexafloroaluminate/Na_3AlF_6

(ii)What is the purpose of cryolite?

To lower the melting point of the electrolyte/Aluminium oxide from about 2015°C to 900°C

(iii)Name the substance used for similar purpose in the Down cell

Calcium chloride/$CaCl_2$

(iv)An alloy of sodium and potassium is used as coolant in nuclear reactors.Explain.

Nuclear reactors generate a lot of heat energy. sodium and potassium alloy reduce/lower the high temperature in the reactors.

(v)Aluminium metal is used to make cooking utensils in preference to other metals.Explain.

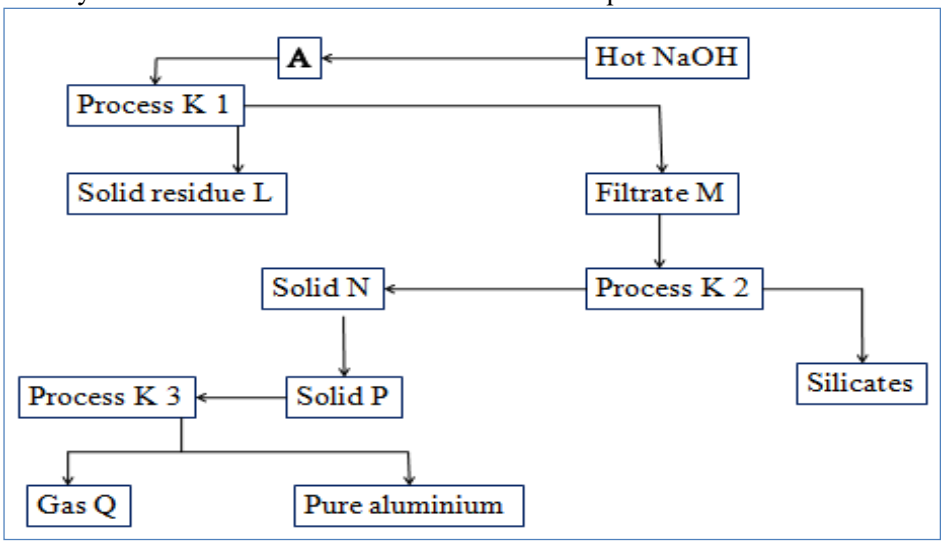
Aluminium

(i) is a very good conductor of electricity because it has three delocalized electrons in its metallic structure

(ii)is cheap,malleable,ductile and has high tensile strength

(iii)on exposure to fire/heat form an impervious layer that prevent it from rapid corrosion.

3.Study the scheme below and use it to answer the questions that follow.



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(a)Identify:

(i)solid residue L

Iron(III)Oxide/Fe_2O_3

(ii)Solid N

Aluminium hydroxide /$Al(OH)_3$

(iii)Filtrate M

Sodium tetrahydroxoaluminate/ $NaAl(OH)_4$ and sodium silicate/ $NaSiO_3$

(iv)Solid P

Aluminium oxide/ Al_2O_3

(v)Gas Q

Oxygen/O_2

(vi)Process K1

Filtration

(vii)Process K2

Electrolysis

(b)Write the equation for the reaction taking place in the formation of solid P from solid N

$2Al(OH)_3 \rightarrow Al_2O_3(s) + 3H_2O(l)$

(c)Name a substance added to solid N before process Process K2 take place.

Cryolite/Sodium tetrahydroxoaluminate/ $NaAl(OH)_4$

(d)State the effect of evolution of gas Q on

(i)process K2

Oxygen produced at the anode reacts with the carbon anode to form carbon(IV) oxide which escape. The electrolytic process needs continuous replacement of the carbon anode.

(ii)the environment

Oxygen produced at the anode reacts with the carbon anode to form carbon(IV) oxide which escape to the atmosphere.CO_2 is a green house gas that cause global warming**.**

(e)An aluminium manufacturing factory runs for 24 hours. If the total mass of aluminium produced is 27000kg,

(i)Calculate the current used. (Faraday constant=96500Coulombs, Al=27.0).

(ii)assuming all the gas produced react with 200kg of anode ,calculate the loss in mass of the electrode.(Molar gas volume at room temperature = 24dm3,C=12.0)

Working

Equation at Cathode

$Al^{3+}(l) + 3e \rightarrow Al(l)$

27g Al -> 3 Faradays = 3 x 96500C

(27000kg x 1000) g -> $\frac{(27000kg \times 1000)\ g \times 3 \times 96500C}{27g}$

=**289500000000 Coulombs**

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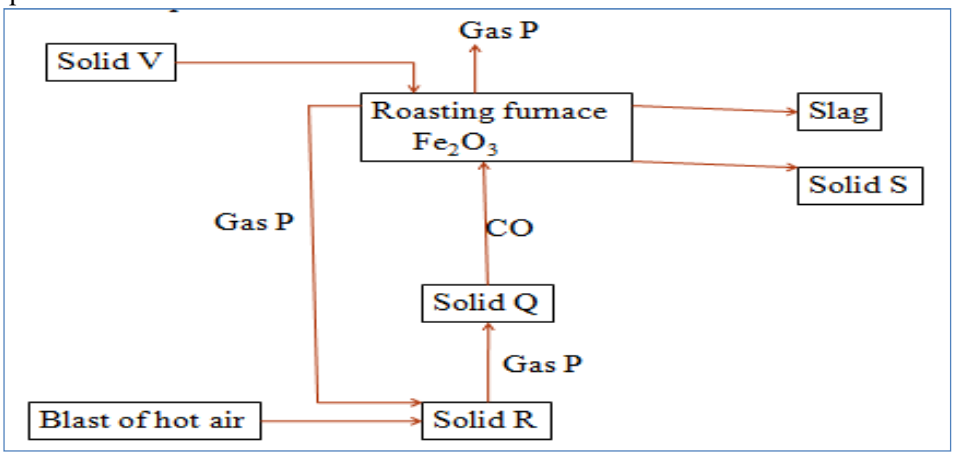
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Current = Quantity of electricity =>289500000000 Coulombs
Time in seconds 24 x 60 x 60
3350690Ampheres

Working
Equation at Anode
$2O^{2-}(l)$ + 4e -> $O_2(g)$
4 Faradays -> 4 x 96500C24dm3 $O_2(g)$ -
289500000000 Coulombs -> 289500000000 Coulombs x 24dm3
4 x 96500C
18,000,000dm3

Chemical equation at anode
$O_2(g)$ + C (s) -> $CO_2(g)$
Method 1

24dm3 of $O_2(g)$ -> 12.0g Carbon
18,000,000dm3 of $O_2(g)$ -> 18,000,000dm3 x 12 = 9000000g = **9000kg**
24dm3 1000g
Loss in mass of the carbon graphite anode = **9000kg**
NB: Mass of the carbon graphite anode remaining =27000kg - 9000kg **=18000kg**
The flow chart below shows the extraction of iron metal.Use it to answer the questions that follow.



(a)Identify:
(i)gas P
Carbon(IV)oxide/CO_2
(ii)Solid Q
Carbon/coke/charcoal

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(iii)Solid R

Carbon/coke/charcoal

(iv)Solid V

Limestone/calcium carbonate/$CaCO_3$

(v)Solid S

Iron/Fe

(b)Write the chemical equation for the reaction for the formation of:

(i)Solid S

$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

(ii)Carbon(II)oxide

$C(s) + CO_2(g) \rightarrow 2CO(g)$

(iii)Slag

$SiO_2(s) + CaO(s) \rightarrow CaSiO_3(s)$

$Al_2O_3(s) + CaO(s) \rightarrow Ca\ Al_2O_4(s)$

(iv)Gas P

$C(s) + O_2(g) \rightarrow CO_2(g)$

(c)State two uses of:

(i)Solid S

Iron is used in making:

(i)gates ,pipes, engine blocks, rails, charcoal iron boxes, lamp posts because it is **cheap**.

(ii)nails, cutlery, scissors, sinks, vats, spanners, steel rods, and railway points from steel.

Steel is an alloy of iron with carbon, and/or Vanadium, Manganese, Tungsten, Nickel ,Chromium.

It does **not rust**/corrode like iron.

(ii)Slag

(i) tarmacing roads

(ii) cement manufacture

(iii) as building construction material

3.You are provided with sulphuric(VI)acid ,2M aqueous ammonia and two ores suspected to contain copper and iron. Describe with explanation how you would differentiate the two ores.

Crush the two ores separately in using a mortar and pestle to reduce the particle size/increase the surface area.

Add sulphuric(VI)acid to separate portion of the ore. Filter.

To a portion of the filtrate,add three drops of 2M aqueous ammonia then axcess

Results

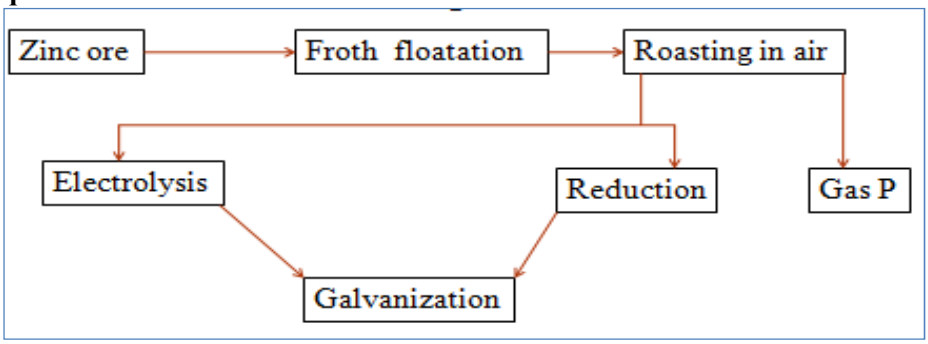
A green precipitate insoluble in excess 2M aqueous ammonia confirms the ore contain Fe^{2+} ion.

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A brown precipitate insoluble in excess 2M aqueous ammonia confirms the ore contain Fe^{3+} ion.
A blue precipitate that dissolve in excess 2M aqueous ammonia to form a deep/royal blue solution confirms the ore contain Cu^{2+} ion.

4. Use the flow chart below showing the extraction of Zinc metal to answer the questions that follow



(a)Name:

(i)two ores from which Zinc can be extracted

Calamine($ZnCO_3$)
Zinc blende(ZnS)

(ii)two possible identity of gas P

Sulphur(IV)oxide(SO_2) from roasting Zinc blende
Carbon(IV)oxide(CO_2) from decomposition of Calamine.

(b)Write a possible chemical equation taking place in the roasting chamber.

$2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$
$ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$

(c)Explain the effect of the by-product of the roating on the environment.

Sulphur (IV)oxide from roasting Zinc blende is an acidic gas that causes “acid rain” on dissolving in rain water.
Carbon(IV)oxide(CO_2) from decomposition of Calamine is a green house gas that causes global warming.

(d)(i)Name a suitable reducing agent used in the furnace during extraction of Zinc.

Carbon(II)oxide

(ii)Write a chemical equation for the reduction process

$ZnO(s) + CO(g) \rightarrow Zn(s) + CO_2(g)$

(e)(i)Before electrolysis, the products from roasting is added dilute sulphuric (VI)acid. Write the equation for the reaction with dilute sulphuric(VI)acid.

$ZnO(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

(ii)During the electrolysis for extraction of Zinc,state the

I. Anode used

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Aluminium sheet

II. Cathode used

Lead plate coated with silver

(ii)Write the equation for the electrolysis for extraction of Zinc at the:

I.Cathode;

$Zn^{2+}(aq) + 2e \rightarrow Zn(s)$

II.Anode;

$4OH^-(aq) \rightarrow 2H_2O(l) + O_2(s) + 4e$

(f)(i)What is galvanization

Dipping Iron in molten Zinc to form a thin layer of Zinc to prevent iron from rusting.

(ii)Galvanized iron sheet rust after some time. Explain

The thin layer of Zinc protect Iron from rusting through sacrificial protection. When all the Zinc has reacted with elements of air, Iron start rusting.

(g)State two uses of Zinc other than galvanization.

Making brass(Zinc/copper alloy)

Making german silver(Zinc/copper/nickel alloy)

As casing for dry cells/battery

(h)Calculate the mass of Zinc that is produced from the reduction chamber if 6400kg of Calamine ore is fed into the roaster. Assume the process is 80% efficient in each stage(Zn=64.0,C=12.0,O=16.0)

Molar mass $ZnCO_3(s)$ =124g

Molar mass Zn = 64g

Molar mass ZnO = 80g

Chemical equation

$ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$

Method 1

124g $ZnCO_3$ => 80g ZnO

(6400kg x1000)g $ZnCO_3$ => $\frac{(6400 \times 1000) \times 80}{124}$ = **512,000,000** g of ZnO

100% => 512,000,000 g of ZnO

80% => $\frac{80 \times 512,000,000}{100}$ g = **409600000**g of ZnO

Chemical equation

$ZnO(s) + CO(g) \rightarrow Zn(s) + CO_2(g)$

80g ZnO(s) => 64g Zn(s)

409600000g of ZnO => $\frac{409600000 \times 64}{80}$ = **327680000 g Zn**

100% => 327680000 g Zn

80% => $\frac{80 \times 327680000}{100}$ g Zn = **262144000**g of Zn

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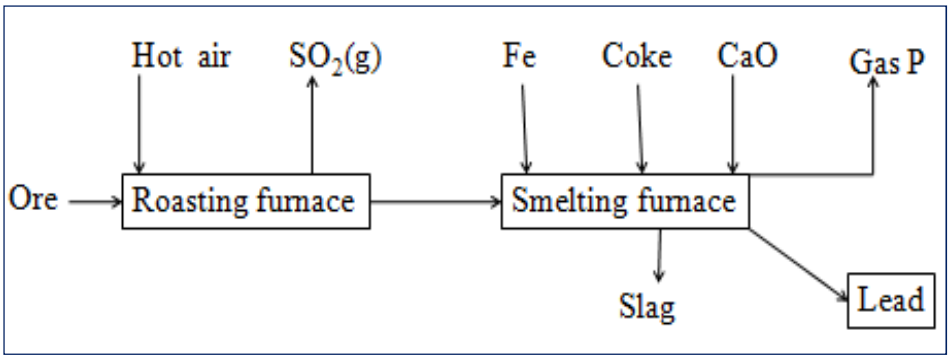
Mass of Zinc produced = **262144000g of Zn**

5.An ore is suspected to bauxite. Describe the process that can be used to confirm the presence of aluminium in the ore.

Crush the ore to fine powder to increase surface area/reduce particle size.
Add hot concentrated sulphuric(VI)/nitric(V) acid to free the ions.
Filter. Retain the filtrate
Add excess aqueous ammonia to a sample of filtrate.
A white precipitate confirms presence of either Al^{3+} or Pb^{2+}.
Add sodium sulphate,dilute sulphuric(VI)to another portion of filtrate.
No white precipitate confirms presence of Al^{3+}

Or Add potassium iodide to another portion of filtrate.
No yellow precipitate confirms presence of Al^{3+}

6.The flow chart below illustrate the industrial extraction of Lead metal



(a)(i)Name the chief ore that is commonly used in this process

Galena(PbS)

(ii)Explain what take place in the roasting furnace

ELECTROCHEMISTRY

Electrochemistry can be defined as the study of the effects of electricity on a substance/ compound and how chemical reactions produce electricity.
Electrochemistry therefore deals mainly with:

i) Reduction and oxidation
ii) Electrochemical (voltaic) cell
iii)Electrolysis (electrolytic) cell

(i)REDUCTION AND OXIDATION (REDOX)

1. In teams of oxygen transfer:
   i) Reduction is **removal** of oxygen.
   ii) Oxidation is **addition** of oxygen.
   iii) Redox is **simultaneous** addition and removal of oxygen.

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iv) Reducing agent is the species that undergoes **oxidation**, therefore **gains** oxygen.

v) Oxidizing agent is the species that undergoes **reduction**, therefore **looses/donates** oxygen.

e.g. When hydrogen is passed through heated copper (II) oxide, it is **oxidised** to copper metal as in the equation below:

$CuO_{(s)}$ + $H_{2(g)}$ -> $Cu_{(s)}$ + $H_2O_{(l)}$

(Oxidising agent) (Reducing agent)

2. In terms of hydrogen transfer:

i) Oxidation is the **removal** of hydrogen.

ii) Reduction is the **addition** of hydrogen.

iii) Redox is **simultaneous** addition and removal of hydrogen.

iv) Reducing agent is the species that undergoes **oxidation**, therefore **looses/ donates** hydrogen.

v) Oxidizing agent is the species that undergoes **reduction,** therefore **gains** hydrogen.

e.g. When hydrogen sulphide gas is bubbled into a gas jar containing chlorine gas it is oxidized (loose the hydrogen) to sulphur (yellow solid). The chlorine is reduced (gain hydrogen) to hydrogen chlorine gas.

$Cl_{2(g)}$ + $H_2S_{(g)}$ -> $S_{(S)}$ + $2HCl_{(g)}$

(Oxidizing agent) (Reducing agent)

3. In terms of electron transfer:

i) Oxidation is **donation/ loss/ removal** of electrons.

ii) Reduction is **gain/ accept/ addition** of electrons.

iii) Redox is **simultaneous gain/ accept/ addition** and **donation/ loss/ removal** of electrons.

iv) Reducing agent is the species that undergoes **oxidation,** therefore **looses/ donates** electrons.

v) Oxidizing agent is the species that undergoes **reduction**, therefore **gains/ accepts** electrons.

Example

a) Displacement of metals from their solutions:

Place $5cm^3$ each of Iron (II) sulphate (VI) solution into three different test tubes. Add about 1g of copper tunings / powder into one test tube then zinc and magnesium powders separately into the other test tubes. Shake thoroughly for 2 minutes each. Record any colour changes in the table below.

Metal added to Iron (II) sulphate (VI) solution	Colour changes
Copper	**Solution remains green**
Zinc	**Green colour fades**

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Magnesium	**Green colour fades**

Explanation

-When a more reactive metal is added to a solution of less reactive metal, it displaces it from its solution.

-When a less reactive metal is added to a solution of a more reactive metal, it does not displace it from its solution.

-Copper is less reactive than iron therefore cannot displace iron its solution.

-Zinc is more reactive than iron therefore can displace iron from its solution.

-Magnesium is more reactive than iron therefore can displace iron from its solution.

In terms of electron transfer:

- the more reactive metal undergoes oxidation (reducing agent) by **donating/loosing** electrons to form **ions**

-the less reactive metal undergoes reduction (oxidizing agent) by its ions in solution gaining /accepting/acquiring the electrons to form the metal.

-displacement of metals involves therefore electron transfer from a more reactive metal to ions of another less reactive metal.

Examples

1. $Zn(s) \rightarrow Zn^{2+}(aq) + 2e$ (oxidation/donation of electrons)
   $Fe^{2+}(aq) + 2e \rightarrow Fe(s)$ (reduction/gain of electrons)
   $Fe^{2+}(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + Fe(s)$ (redox/both donation and gain of electrons)
2. $Mg(s) \rightarrow Mg^{2+}(aq) + 2e$ (oxidation/donation of electrons)
   $Fe^{2+}(aq) + 2e \rightarrow Fe(s)$ (reduction/gain of electrons)
   $Fe^{2+}(aq) + Mg(s) \rightarrow Mg^{2+}(aq) + Fe(s)$ (redox/both donation and gain of electrons)
3. $Zn(s) \rightarrow Zn^{2+}(aq) + 2e$ (oxidation/donation of electrons)
   $Cu^{2+}(aq) + 2e \rightarrow Cu(s)$ (reduction/gain of electrons)
   $Cu^{2+}(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + Cu(s)$ (redox/both donation and gain of electrons)
4. $Fe(s) \rightarrow Fe^{2+}(aq) + 2e$ (oxidation/donation of electrons)
   $2Ag^{+}(aq) + 2e \rightarrow 2Ag(s)$ (reduction/gain of electrons)
   $2Ag^{+}(aq) + Fe(s) \rightarrow Fe^{2+}(aq) + 2Ag(s)$ (redox/both donation and gain of electrons)
5. **Zn**$(s) \rightarrow Zn^{2+}(aq) + 2e$ (oxidation/donation of electrons)
   $Cl_2(g) + 2e \rightarrow 2Cl^{-}(aq)$ (reduction/gain of electrons)
   $Cl_2(g) +$ **$Zn(s) \rightarrow Zn^{2+}(aq)$** $+ 2Cl^{-}(aq)$ (redox/both donation and gain of electrons)
6. **2Mg**$(s) \rightarrow 2Mg^{2+}(aq) + 4e$ (oxidation/donation of electrons)
   $O_2(g) + 4e \rightarrow 2O^{2-}(aq)$ (reduction/gain of electrons)
   $O_2(g) +$ **$2Mg(s) \rightarrow 2Mg^{2+}(aq)$** $+ 2O^{2-}(aq)$ (redox/both donation and gain of electrons)

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Note

(i)The number of electrons donated/lost MUST be equal to the number of electrons gained/acquired.

(i)During displacement reaction, the colour of ions /salts fades but does not if displacement does not take place. e.g

a)Green colour of **Fe^{2+}(aq) fades** if Fe^{2+}(aq) ions are **displaced** from their solution. Green colour of **Fe^{2+}(aq) appear** if Fe/iron **displaces** another salt/ions from their solution.

b)Blue colour of **Cu^{2+}(aq) fades** if Cu^{2+}(aq) ions are **displaced** from their solution and **brown** copper deposits appear. Blue colour of **Cu^{2+}(aq) appear** if Cu/copper **displaces** another salt/ions from their solution.

c)Brown colour of **Fe^{3+}(aq) fades** if Fe^{3+}(aq) ions are **displaced** from their solution. Brown colour of **Fe^{3+}(aq) appear** if Fe/iron **displaces** another salt/ions from their solution to form **Fe^{3+}(aq)**.

(iii)Displacement reactions also produce **energy/heat**. The **closer/nearer** the metals in the reactivity/electrochemical series the **less** energy/heat of displacement.

(iv)The **higher** the metal in the reactivity series therefore the **easier** to loose/donate electrons and thus the **stronger** the reducing agent.

4. (a)In terms of oxidation number:

i) Oxidation is increase in oxidation numbers.

ii) Reduction is decrease in oxidation numbers.

iii) Redox is simultaneous increase in oxidation numbers of one species/substance and a decrease in oxidation numbers of another species/substance.

iv) Reducing agent is the species that undergoes oxidation, therefore increases its oxidation number.

v) Oxidizing agent is the species that undergoes reduction, therefore increases its oxidation number.

(b)The idea/concept of oxidation numbers uses/applies the following simple guideline rules:

Guidelines /rules applied in assigning oxidation number

1.Oxidation number of combined Oxygen is always **-2** except in peroxides (Na_2O_2/H_2O_2) where its Oxidation number is -1

2.Oxidation number of combined Hydrogen is always **+1**except in Hydrides (NaH/KH) where its Oxidation number is -1

3.All **atoms** and **molecules** of elements have oxidation number **0** (zero)

Atom	Oxidation number	**Molecule**	**Oxidation number**
Na	0	**Cl_2**	**0**
O	0	**O_2**	**0**
H	0	**H_2**	**0**

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Al	0	**N_2**	**0**
Ne	0	**O_3**	**0**
K	0	**P_3**	**0**
Cu	0	**S_8**	**0**

4.All **combined metals** and **non-metals** have oxidation numbers equal to their valency /oxidation state e.g.

Metal/non-metal ion	**Valency**	**Oxidation state**	**Oxidation number**
Fe^{2+}	2	-2	-2
Fe^{3+}	3	-3	-3
Cu^{2+}	2	-2	-2
Cu^{+}	1	+1	+1
Cl^{-}	1	-1	-1
O^{2-}	2	-2	-2
Na^{+}	1	+1	+1
Al^{3+}	3	+3	+3
P^{3-}	3	-3	-3
Pb^{2+}	2	+2	+2

5.Sum of oxidation numbers of atoms of elements making a compound is equal zero(**0**) e.g.

Using this rule ,an unknown oxidation number of an atom in a compound can be determined as below:

a) $CuSO_4$ has-

-one atom of Cu with oxidation number +2(refer to Rule 4)

-one atom of S with oxidation number +6 (refer to Rule 4)

-six atoms of O each with oxidation number -2(refer to Rule 4)

Sum of oxidation numbers of atoms in $CuSO_4$ = (**+2** + **+6** + (**-2** x 6)) = **0**

b) H_2SO_4 has-

-two atom of H each with oxidation number +1(refer to Rule 2)

-one atom of S with oxidation number +6 (refer to Rule 4)

-four atoms of O each with oxidation number -2(refer to Rule 4)

Sum of oxidation numbers of atoms in H_2SO_4 = (**+2** + **+6** + (**-2** x 4)) = **0**

c) $KMnO_4$ has-

-one atom of K with oxidation number +1(refer to Rule 4)

-one atom of Mn with oxidation number +7 (refer to Rule 4)

-four atoms of O each with oxidation number -2(refer to Rule 4)

Sum of oxidation numbers of atoms in $KMnO_4$ = (**+1** + **+7** + (**-2** x 4)) = **0**

Determine the oxidation number of:

I.Nitrogen in;

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-NO => x + -2 = 0 thus x = 0 – (-2) = **+ 2**
The chemical name of this compound is thus Nitrogen(**II**)oxide
-NO_2 => x + (-2 x2)= 0 thus x = 0 – (-4) = **+ 4**
The chemical name of this compound is thus Nitrogen(**IV**)oxide
-N_2O => 2x + -2 = 0 thus 2x = 0 – (-2) = **+2/2= +1**
The chemical name of this compound is thus Nitrogen(**I**)oxide

II. Sulphur in;

-SO_2 => x + (-2 x2)= 0 thus x = 0 – (-4) = **+ 4**
The chemical name of this compound is thus Sulphur(**IV**)oxide
-SO_3 => x + (-2 x3)= 0 thus x = 0 – (-6) = **+ 6**
The chemical name of this compound is thus Sulphur(**VI**)oxide
-H_2SO_4 = ((+1 x 2) + x + (-2 x 4)) thus x= 0-(+2 +-8) =**+6**
The chemical name of this compound is thus Sulphuric(**VI**)acid
-H_2SO_3 = ((+1 x 2) + x + (-2 x 3)) thus x= 0-(+2 +-6) =**+4**
The chemical name of this compound is thus Sulphuric(**IV**)acid

III. Carbon in;

-CO_2 => x + (-2 x2)= 0 thus x = 0 – (-4) = **+ 4**
The chemical name of this compound is thus carbon(**IV**)oxide
-CO => x + -2 = 0 thus x = 0 – -2 = **+ 2**
The chemical name of this compound is thus carbon(**II**)oxide
-H_2CO_3 = ((+1 x 2) + x + (-2 x 3)) thus x= 0-(+2 +-6) =**+4**
The chemical name of this compound is thus Carbonic(**IV**)acid

IV.Manganese in;

-MnO_2 => x + (-2 x2)= 0 thus x = 0 – (-4) = **+ 4**
The chemical name of this compound is thus Manganese(**IV**)oxide
-$KMnO_4$ = ((+1 + x + (-2 x 4)) thus x= 0-(+1 +-8) =**+7**
The chemical name of this compound is thus Potassium manganate(**VII**)

V.Chromium in;

- Cr_2O_3 => 2x + (-2 x 3)= 0 thus 2x = 0 – (-6) = **+6 / 2= +3**
The chemical name of this compound is thus Chromium(**III**)oxide
-$K_2Cr_2O_7$ => (+1 x 2) + 2x + (-2 x7)= 0
thus 2x = 0 – +2 +-14 = **+12 / 2= +6**
The chemical name of this compound is thus Potassium **di**chromate(**VI**)
-K_2CrO_4 => (+1 x 2) + x + (-2 x4)= 0
thus 2x = 0 – +2 +-8 = **+12 / 2= +6**
The chemical name of this compound is thus Potassium chromate(**VI**)

6.The sum of the oxidation numbers of atoms of elements making a charged radical/complex ion is equal to its charge.

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Using this rule ,the oxidation number of unknown atom of an element in a charged radical/complex ion can be determined as in the examples below;

a) SO_4^{2-} has-

-one atom of S with oxidation number +6(refer to Rule 4)

-four atoms of O each with oxidation number -2(refer to Rule 1)

Sum of oxidation numbers of atoms in SO_4^{2-} = (**+6** + (**-2** x 4)) **= -2**

The chemical name of this radical is thus sulphate(**VI**) ion

b) NO_3^- has-

-one atom of N with oxidation number +4(refer to Rule 4)

-three atoms of O each with oxidation number -2(refer to Rule 1)

Sum of oxidation numbers of atoms in NO_3^- = (**+4** + (**-2** x 3)) **= -1**

The chemical name of this radical is thus nitrate(**IV**) ion.

Determine the oxidation number of:

I.Nitrogen in;

$-NO_2^-$ => x + (-2 x2)= -1 thus x = -1 – (-4) = **+ 3**

The chemical name of this compound/ion/radical is thus Nitrate(**III**)ion

II. Sulphur in;

$-SO_3^{2-}$ => x + (-2 x3)= -2 thus x = -2 – (-6) = **+ 4**

The chemical name of this compound/ion/radical is thus Sulphate(**IV**)ion

III. Carbon in;

$-CO_3^{2-}$ = x + (-2 x 3) = -2 thus x = -2 – (-6) = **+ 4**

The chemical name of this compound/ion/radical is thus Carbonate(**IV**)ion

IV.Manganese in;

$-MnO_4^-$ = x + (-2 x 4)= -1 thus x= -1-(-2 +-8) =**+7**

The chemical name of this compound/ion/radical is thus manganate(**VII**) ion

V.Chromium in

$-Cr_2O_7^{2-}$ => 2x + (-2 x7)= -2

thus 2x = -2 – +2 +-14 = **+12 / 2= +6**

The chemical name of this compound/ion//radical is thus dichromate(**VI**) ion

$-CrO_4^{2-}$ => x + (-2 x4)= -2

thus x = -2 + (-2 x 4) **= +6**

The chemical name of this compound/ion//radical is thus chromate(**VI**) ion

(c)Using the concept/idea of oxidation numbers as increase and decrease in oxidation numbers , the oxidizing and reducing species/agents can be determined as in the following examples;

(i) **Cu^{2+} (aq) + Zn(s) -> Zn^{2+} (aq) + Cu(s)**

Oxidation numbers -> +2 0 +2 0

Oxidizing species/agents =>Cu^{2+} ;its oxidation number decrease from+2 to 0 in Cu(s)

Reducing species/agents => Zn^{2+} ;its oxidation number increase from 0 to +2 in Zn(s)

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(ii) **$2Br^-$ (aq) + Cl_2(g) -> $2Cl^-$ (aq) + Br_2 (l)**
Oxidation numbers -> -1 0 -1 0
Oxidizing agent =>Cl_2(g) ;its oxidation number decrease from 0 to-1 in $2Cl^-$ **(aq)**
Reducing agents => Zn^{2+} ;its oxidation number increase from -1 to 0 in Zn(s)
(iii) **Br_2 (l) + Zn(s) -> Zn^{2+} (aq) + $2Br^-$ (aq)**
Oxidation numbers -> 0 0 +2 -1
Oxidizing agent => Br_2 (l) ;its oxidation number decrease from 0 to-1 in $2Br^-$(aq)
Reducing agents => Zn(s) ;its oxidation number increase from 0 to +2 in Zn^{2+}
(iv) **2HCl (aq) + Mg(s) -> $MgCl_2$ (aq) + H_2 (g)**
Oxidation numbers -> 2 (+1 -1) 0 +2 2(-1) 0
Oxidizing agent => **H^+ in HCl**;its oxidation number decrease from +1to 0 in **H_2 (g)**
Reducing agents => Mg(s) ;its oxidation number increase from 0 to +2 in Mg^{2+}
(v) **$2H_2O$ (l) + 2Na(s) -> 2NaOH (aq) + H_2 (g)**
Oxidation numbers -> +1 -2 0 +1 -2 +1 0
Oxidizing agent => **H^+ in** H_2O;its oxidation number decrease from +1to 0 in **H_2 (g)**
Reducing agents => Na(s) ;its oxidation number increase from 0 to +1 in Na^+
(vi) $5Fe^{2+}$(aq) + $8H^+$(aq) + MnO_4^- -> $5Fe^{3+}$(aq) + Mn^{2+}(aq) + $4H_2O$ (l)
+2 +1 +7 -2 +3 +2 +1 -2
Oxidizing agent => **Mn in** MnO_4^- ;its oxidation number decrease from +7to+2 in **Mn^{2+}**
Reducing agents => **Fe^{2+}** ;its oxidation number increase from +2 to +3 in Fe^{3+}
(vii) $6Fe^{2+}$(aq) + $14H^+$(aq) + $Cr_2O_7^{2-}$(aq) -> $6Fe^{3+}$(aq) + Cr^{3+}(aq) + $7H_2O$ (l)
+2 +1 +6 -2 +3 +3 +1 -2
Oxidizing agent:
Cr in $Cr_2O_7^{2-}$;its oxidation number decrease from +6 to+3 in **Cr^{3+}**
Reducing agents => **Fe^{2+}** ;its oxidation number increase from +2 to +3 in Fe^{3+}
(viii) $2Fe^{2+}$(aq) + $2H^+$(aq) + H_2O_2(aq) -> $2Fe^{3+}$(aq) + $2H_2O$ (l)
+2 +1 +1 -1 +3 +1 -2
Oxidizing agent:
O in H_2O_2;its oxidation number decrease from -1 to -2 in H_2O
Reducing agents => **Fe^{2+}** ;its oxidation number increase from +2 to +3 in Fe^{3+}
(ix) $Cr_2O_7^{2-}$(aq) + $6H^+$(aq) + $5H_2O_2$(aq) -> $2Cr^{3+}$(aq) + $2H_2O$ (l) + $5O_2$(g)
+6 -2 +1 +1 -1 +3 +1 -2 0
Oxidizing agents:
O in H_2O_2;its oxidation number decrease from -1 to -2 in H_2O
Cr in $Cr_2O_7^{2-}$ its oxidation number decrease from +6 to +3 in Cr^{3+}
Reducing agents
O in H_2O_2;its oxidation number increase from -1 to O in O_2(g)
O in $Cr_2O_7^{2-}$ its oxidation number increase from -2 to O in O_2(g)

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(x) $2MnO_4^-(aq) + 6H^+(aq) + 5H_2O_2(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5O_2(g)$

+7 -2 +1 +1 -1 +2 +1 -2 0

Oxidizing agents:

O in H_2O_2;its oxidation number decrease from -1 to -2 in H_2O

Mn in MnO_4^- its oxidation number decrease from +7 to +2 in Mn^{2+}

Reducing agents

O in H_2O_2;its oxidation number increase from -1 to O in $O_2(g)$

O in MnO_4^- its oxidation number increase from -2 to O in $O_2(g)$

(ii)ELECTROCHEMICAL (VOLTAIC) CELL

1. When a metal rod/plate is put in a solution of its own salt, some of the metal ionizes and dissolve into the solution i.e.

$M(s) \rightarrow M^+(aq) + e$ (monovalent metal)

$M(s) \rightarrow M^{2+}(aq) + 2e$ (divalent metal)

$M(s) \rightarrow M^{3+}(aq) + 3e$ (Trivalent metal)

The ions move into the solution leaving electrons on the surface of the metal rod/plate.

2.The **metal** rod becomes therefore **negatively** charged while its own **solution positively** charged. As the positive charges of the solution increase, some of them recombine with the electrons to form back the metal atoms

$M^+(aq) + e \rightarrow M(s)$ (monovalent metal)

$M^{2+}(aq) + 2e \rightarrow M(s)$ (divalent metal)

$M^{3+}(aq) + 3e \rightarrow M(s)$ (Trivalent metal)

3. When a metal rod/plate is put in a solution of its own salt, it constitutes/forms a **half-cell**. The tendency of metals to ionize differ from one metal to the other. The difference can be measured by connecting two half cells to form an electrochemical/voltaic cell as in the below procedure:

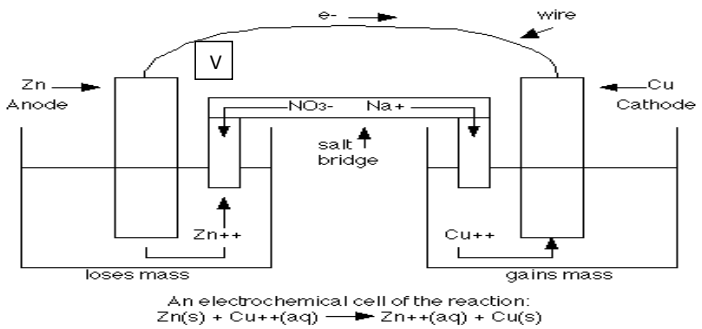
To set up an electrochemical /voltaic cell

To compare the relative tendency of metals to ionize

Place 50cm3 of 1M Zinc(II) sulphate(VI) in 100cm3 beaker. Put a clean zinc rod/plate into the solution. Place 50cm3 of 1M Copper(II) sulphate(VI) in another 100cm3 beaker. Put a clean copper rod/plate of equal area (length x width) with Zinc into the solution. Connect/join the two metals(to a voltmeter) using connecting wires. Dip a folded filter paper into a solution of Potassium nitrate(V) or sodium(I) chloride(I) until it soaks. Use the folded soaked filter paper to connect/join the two solutions in the two beakers. The whole set up should be as below

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An electrochemical cell of the reaction:
$Zn(s) + Cu^{++}(aq) \longrightarrow Zn^{++}(aq) + Cu(s)$

Repeat the above procedure by replacing:
(i)Zinc half cell with Magnesium rod/plate/ribbon dipped in 50cm3 of IM magnesium (II) sulphate(VI) solution
(ii)Zinc half cell with Silver rod/plate/coin dipped in 50cm3 of IM silver(I) nitrate(V) solution
(iii)Copper half cell with Iron rod/plate/spoon dipped in 50cm3 of IM Iron (II) sulphate(VI) solution
Record the observations in the table below

Changes on the 1st metal rod (A)	Changes on the 2nd metal rod (B)	Changes on the 1st solution (A(aq))	Changes on the 2nd solution (**B**(aq))	Voltage/voltmeter reading(Volts)
Using Zn/Cu half cell -The rod decrease in size /mass /dissolves/ erodes	-copper rod /plate increase in size /mass/ deposited	Zinc(II)sulphate (VI)colour remain colourless	Blue Copper (II)sulphate (VI)colour fades. Brown solid/residue/ deposit	0.8 (Theoretical value=**1.10V**)
Using Mg/Cu half cell -The rod decrease in size /mass /dissolves/ erodes	-copper rod /plate increase in size /mass/ deposited	Magnesium(II) sulphate(VI) colour remain colourless	Blue Copper (II)sulphate (VI)colour fades Brown solid/residue/ deposit	1.5 (Theoretical value=**2.04V**)

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Using Ag/Cu half cell -The rod increase in size /mass /deposited	-silver coin/ rod /plate increase in size /mass/ deposited	Blue Copper (II)sulphate (VI)colour remains	Silver(I)nitrate (V)colour remain colourless	0.20 (Theoretical value=**0.46V**)
Using Fe/Cu half cell -The rod decrease in size /mass /dissolves/ erodes	-copper rod /plate increase in size /mass/ deposited	Iron(II)sulphate (VI)colour becomes more green	Blue Copper (II)sulphate (VI)colour fades.Brown solid/residue/ deposit	0.60 (Theoretical value=**0.78V**)

From the above observations ,it can be deduced that:
(i)in the Zn/Cu half-cell the;

-Zinc rod/plate ionizes /dissolves faster than the copper rod/plate to form Zn^{2+}

Ionic equation **$Zn(s) \rightarrow Zn^{2+}(aq) + 2e$**

-blue copper ions in the Copper (II)sulphate solution gains the donated electrons to form brown copper metal/atoms

Ionic equation **$Cu^{2+}(aq) + 2e \rightarrow Cu(s)$**

This reaction shows /imply the **Zinc** rod has a **higher** tendency to ionize than **copper.**The **Zinc** rod has a **higher net** accumulation of electrons and is more **negative** compared to the copper rod which has **lower** accumulation of electrons. The copper rod is therefore relatively more **positive** with respect to Zinc rod.
When the two half cells are connected , electrons therefore flow from the **negative** Zinc rod through the external wire to be gained by copper **ions**. This means a net accumulation/**increase** of Zn^{2+} positive ions on the negative half cell and a net **decrease** in Cu^{2+} positive ions on the positive half cell.
The purpose of the salt bridge therefore is:

(i)complete the circuit

(ii)maintain balance of charges /ions on both half cells.

For the negative half cell the NO_3^- /Cl^- from salt bridge **decrease/neutralise** the increased **positive**(Zn^{2+}) ion.
For the positive half cell the Na^+ / K^+ from salt bridge **increase** the decreased **positive**(Cu^{2+}) ion.

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The voltmeter should theoretically register/read a 1.10Volts as a measure of the electromotive force (**e.m.f**) of the cell .Practically the voltage reading is lowered because the connecting wires have some **resistance** to be overcomed.
A combination of two half cells that can **generate** an electric **current** from a **redox** reaction is called a voltaic/electrochemical cell.
By convention a voltaic/electrochemical cell is represented;

M(s) / M^{2+}(aq) // N^{2+} (aq) / N(s)

(metal rod of M)(solution ofM)(solution ofN)(metal rod ofN)

Note;

a)(i)Metal M must be the one **higher** in the reactivity series.
(ii)It forms the **negative** terminal of the cell.
(iii)It must diagrammatically be drawn **first** on the **left hand side** when illustrating the voltaic/electrochemical cell.

b)(i)Metal N must be the one **lower** in the reactivity series.
(ii)It forms the **positive** terminal of the cell.
(iii)It must diagrammatically be drawn second/**after**/ **right hand side** when illustrating the voltaic/electrochemical cell.

Illustration of the voltaic/electrochemical cell.

(i)Zn/Cu cell

1. Zinc rod ionizes /dissolves to form Zn^{2+} ions at the negative terminal

Zn(s) -> Zn^{2+}(aq) + 2e

2. Copper ions in solution gain the donated electrons to form copper atoms/metal

Cu^{2+}(aq) + 2e -> Cu(s)

3.Overall redox equation

Cu^{2+}(aq) + Zn(s) -> Zn^{2+}(aq) + Cu(s)

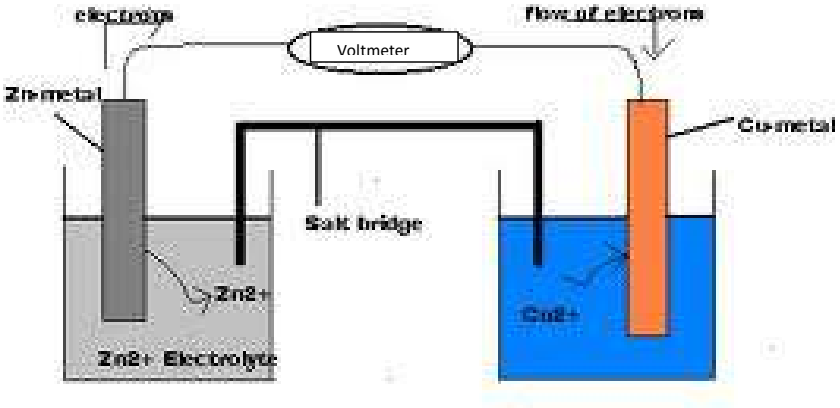
4.cell representation.

Zn(s) / 1M, Zn^{2+}(aq) // 1M,Cu^{2+}(aq) / Cu(s) E^0 = +1.10 V

5.cell diagram

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(ii)Mg/Cu cell

1. Magnesium rod ionizes /dissolves to form Mg^{2+} ions at the negative terminal

$Mg(s) \rightarrow Mg^{2+}(aq) + 2e$

2. Copper ions in solution gain the donated electrons to form copper atoms/metal

$Cu^{2+}(aq) + 2e \rightarrow Cu(s)$

3.Overall redox equation

$Cu^{2+}(aq) + Mg(s) \rightarrow Mg^{2+}(aq) + Cu(s)$

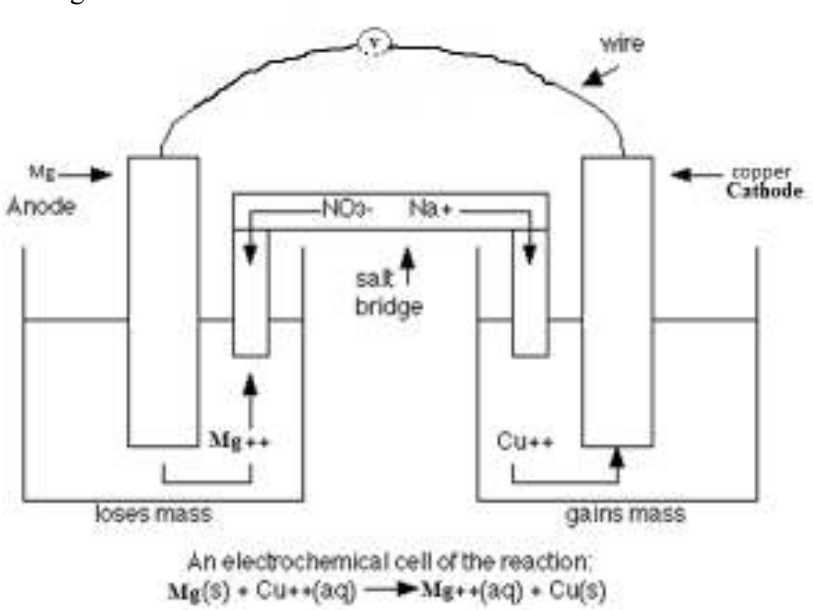
4.cell representation.

$Mg(s) / 1M, Mg^{2+}(aq) // 1M, Cu^{2+}(aq) / Cu(s)$ $E^0 = +2.04$ V

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5.cell diagram.



(iii)Fe/Cu cell

1. Magnesium rod ionizes /dissolves to form Mg^{2+} ions at the negative terminal

$Fe(s) \rightarrow Fe^{2+}(aq) + 2e$

2. Copper ions in solution gain the donated electrons to form copper atoms/metal

$Cu^{2+}(aq) + 2e \rightarrow Cu(s)$

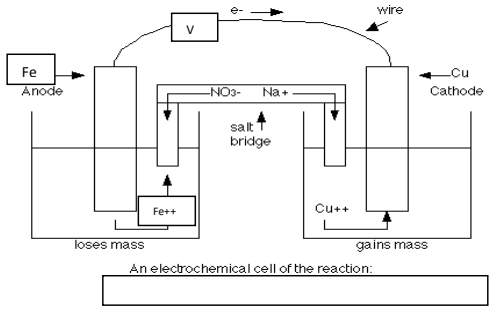
3.Overall redox equation

$Cu^{2+}(aq) + Fe(s) \rightarrow Fe^{2+}(aq) + Cu(s)$

4.cell representation.

$Fe(s) / 1M, Fe^{2+}(aq) // 1M, Cu^{2+}(aq) / Cu(s)$ $E^0 = +0.78$ V

5.cell diagram.



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(iv)Ag/Cu cell

1. Copper rod ionizes /dissolves to form Cu^{2+} ions at the negative terminal

Cu(s) -> Cu^{2+}(aq) + 2e

2. Silver ions in solution gain the donated electrons to form silver atoms/metal

$2Ag^{+}$(aq) + 2e -> 2Ag(s)

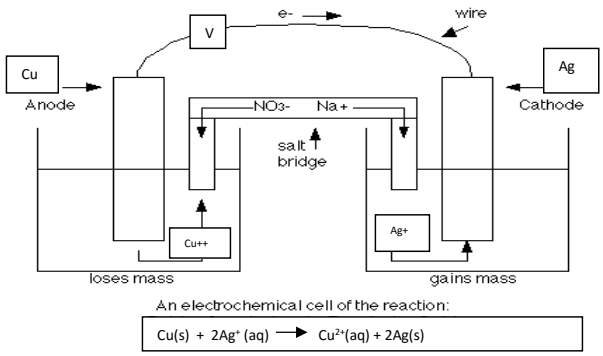
3.Overall redox equation

$2Ag^{+}$(aq) + Cu(s) -> Cu^{2+}(aq) + 2Ag(s)

4.cell representation.

Cu(s) / 1M, Cu^{2+}(aq) // 1M,$2Ag^{+}$(aq) / 2Ag(s) E^{0} = +0.46 V

5.cell diagram.



Standard electrode potential (E^{0})

The **standard** electrode potential (**E^{0}**) is obtained if the **hydrogen** half cell is used as **reference.** The standard electrode potential (E^{0}) consist of inert platinum electrode immersed/dipped in 1M solution of (sulphuric(VI) acid) H^{+} ions. Hydrogen gas is bubbled on the platinum electrodes at:

(i)a temperature of 25°C

(ii)atmospheric pressure of 101300Pa/101300Nm^{-2}/1atm/760mmHg/76cmHg

(iii)a concentration of 1M(1$moledm^{-3}$) of sulphuric(VI) acid/ H^{+} ions and 1M(1$moledm^{-3}$) of the other half cell.

Hydrogen is **adsorbed** onto the **surface** of the platinum. An **equilibrium/balance** exist between the adsorbed layer of molecular hydrogen and H^{+} ions in solution to form a half cell.

$\frac{1}{2} H_2$ (g) ==== H^{+} (aq) + e

The half cell representation is:

Pt,$\frac{1}{2} H_2$ (g) / H^{+} (aq), 1M

The standard electrode potential (**E^{0}**) is thus defined as **the potential difference for a cell comprising of a particular element in contact with1M solution of its own ions and the standard hydrogen electrode.**

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If the other electrode has a **higher/greater** tendency to lose electrons than the hydrogen electrode, the electrode is therefore **negative** with respect to hydrogen electrode and its electrode potential has **negative** (**E⁰**) values.
If the other electrode has a **lower/lesser** tendency to lose electrons than the hydrogen electrode, the electrode is therefore **positive** with respect to hydrogen electrode and its electrode potential has **positive** (**E⁰**) values.

Table showing the standard electrode potential (**E⁰**) of some reactions

Reaction	**(E⁰) values in volts**
F_2 (g)+ 2e -> $2F^-$ (aq)	+2.87
H_2O_2 (aq)+ H^+ (aq) +2e -> H_2O (l)	+1.77
MnO_4^- (aq)+ $4H^+$ (aq) +3e -> MnO_2 (s) +H_2O (l)	+1.70
2HClO (aq)+ $2H^+$ (aq) +2e -> Cl_2 (aq) +$2H_2O$ (l)	+1.59
MnO_4^- (aq)+ $4H^+$ (aq) +5e -> Mn^{2+} (aq) +H_2O (l)	+1.51
Cl_2 (g)+ 2e -> $2Cl^-$ (aq)	+1.36
MnO_2 (s)+ $4H^+$ (aq) +2e -> Mn^{2+} (aq) +$2H_2O$ (l)	+1.23
Br_2 (aq)+ 2e -> $2Br^-$ (aq)	+1.09
NO_3^- (aq)+ $2H^+$ (aq) + e -> NO_2 (g) + H_2O (l)	+0.80
Ag^+ (aq) + e -> Ag(s)	+0.80
Fe^{3+} (aq) + e -> Fe^{2+} (aq)	+0.77
$2H^+$ (aq)+ O_2 (g) -> H_2O_2 (aq)	+0.68
I_2 (aq)+ 2e -> $2I^-$ (aq)	+0.54
Cu^{2+} (aq) + 2e -> Cu(s)	+0.34
$2H^+$ (aq) + 2e -> H_2(g)	**+0.00**
Pb^{2+} (aq) + 2e -> Pb(s)	-0.13
Fe^{2+} (aq) + 2e -> Fe(s)	-0.44
Zn^{2+} (aq) + 2e -> Zn(s)	-0.77
Al^{3+} (aq) + 3e -> Al(s)	-1.66
Mg^{2+} (aq) + 2e -> Mg(s)	-2.37
Na^+ (aq) + e -> Na(s)	-2.71
K^+ (aq) + e -> K(s)	-2.92

Note:

(i)**E⁰** values generally show the **possibility/feasibility** of a **reduction** process/**oxidizing strength.**

(ii)The **element**/species in the half cell with the **highest** negative **E⁰** value easily **gain / acquire** electrons.

It is thus the **strongest oxidizing agent** and its reduction process is highly **possible/feasible.** The **element**/species in the half cell with the **lowest** positive **E⁰** value easily **donate / lose** electrons.

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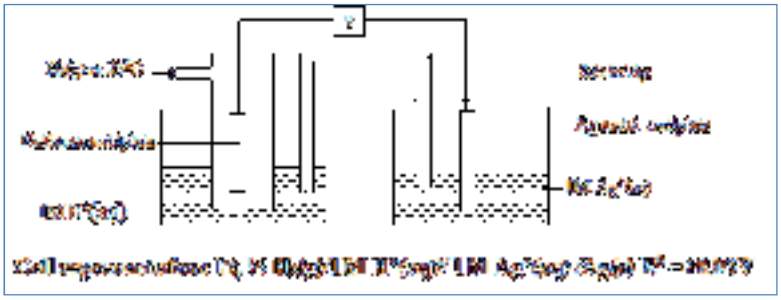
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It is thus the **strongest reducing agent** and its reduction process is the least **possible/feasible.**

(iii)The **overall** redox reaction is **possible**/feasible is it has a **positive** (+) **E^{θ}**.

If the **overall** redox reaction is **not** possible/ **not** feasible/ **forced**, it has a **negative** (-) **E^{θ}**

Sample standard electrochemical cell



Calculation examples on E^{θ}

Calculate the E^{θ} value of a cell made of:

a)Zn and Cu

From the table above:

$Cu^{2+}(aq) + 2e \rightarrow Cu(s)$ $E^{\theta} = +0.34V$(higher E^{θ} /**R**ight **H**and **S**ide diagram)

$Zn^{2+}(aq) + 2e \rightarrow Zn(s)$ $E^{\theta} = -0.77V$(lower E^{θ}/ **L**eft **H**and **S**ide diagram)

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e$ $E^{\theta} = +0.77$(reverse lower E^{θ} to derive cell reaction / representation)

Overall E^{θ} = **E^{θ} higher- E^{θ} lower / E^{θ} RHS - E^{θ} LHS/ E^{θ}oxidized- E^{θ} reduced**

Substituting:

Overall E^{θ} = +0.34 – (- 0.77) = **+1.10V**

Overall redox equation:

$Cu^{2+}(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + Cu(s)$ $E^{\theta} = +1.10V$

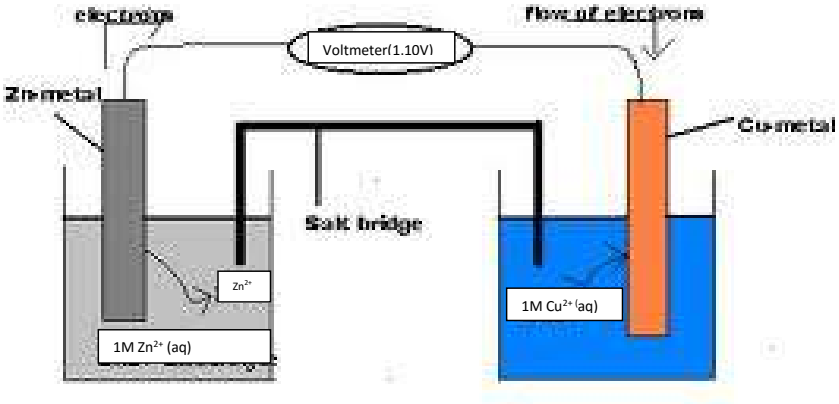
Overall conventional cell representation:

$Zn(s) / Zn^{2+}(aq)$ 1M, // 1M,$Cu^{2+}(aq)$ / Cu(s) $E^{\theta} = +1.10V$

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Overall conventional cell diagram:



Zinc and copper reaction has a **positive**(+) overall E^{θ} therefore is possible/feasible and thus Zinc can displace/reduce Copper solution.

b)Mg and Cu

From the table above:

$Cu^{2+}(aq) + 2e \rightarrow Cu(s)$ $E^{\theta} = +0.34V$(higher E^{θ} /**R**ight **H**and **S**ide diagram)

$Mg^{2+}(aq) + 2e \rightarrow Mg(s)$ $E^{\theta} = -2.37V$(lower E^{θ}/ **L**eft **H**and **S**ide diagram)

$Mg(s) \rightarrow Mg^{2+}(aq) + 2e$ $E^{\theta} = +2.37$(reverse lower E^{θ} to derive cell reaction / representation)

Overall E^{θ} = **E^{θ} higher- E^{θ} lower / E^{θ} RHS - E^{θ} LHS/ E^{θ} oxidized- E^{θ} reduced**

Substituting:

Overall E^{θ} = +0.34 – (- 2.37) = **+2.71V**

Overall redox equation:

$Cu^{2+}(aq) + Mg(s) \rightarrow Mg^{2+}(aq) + Cu(s)$ $E^{\theta} = +2.71V$

Overall conventional cell representation:

$Mg(s) / Mg^{2+}(aq)$ 1M, // 1M,$Cu^{2+}(aq) / Cu(s)$ $E^{\theta} = +2.71V$

c)Ag and Pb

From the table above:

$2Ag^{+}(aq) + 2e \rightarrow 2Ag(s)$ $E^{\theta} = +0.80V$(higher E^{θ} /**R**ight **H**and **S**ide diagram)

$Pb^{2+}(aq) + 2e \rightarrow Pb(s)$ $E^{\theta} = -0.13V$(lower E^{θ}/ **L**eft **H**and **S**ide diagram)

$Pb(s) \rightarrow Pb^{2+}(aq) + 2e$ $E^{\theta} = +0.13$(reverse lower E^{θ} to derive cell reaction / representation)

Overall E^{θ} = **E^{θ} higher- E^{θ} lower / E^{θ} RHS - E^{θ} LHS/ E^{θ} oxidized- E^{θ} reduced**

Substituting:

Overall E^{θ} = +0.80 – (- 0.13) = **+0.93V**

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Overall redox equation:

$2Ag^+(aq) + Pb(s) \rightarrow Pb^{2+}(aq) + 2Ag(s) \quad E^\theta = +0.93V$

Overall conventional cell representation:

$Pb(s) / Pb^{2+}(aq)\ 1M, // 1M, 2Ag^+(aq) / Ag(s) \quad E^\theta = +0.93V$

d)Chlorine and Bromine

From the table above:

$2e + Cl_2(g) \rightarrow 2Cl^-(aq) \quad E^\theta = +1.36V$(higher E^θ /**R**ight **H**and **S**ide diagram)

$2e + Br_2(aq) \rightarrow 2Br^-(aq) \quad E^\theta = +0.13V$(lower E^θ/ **L**eft **H**and **S**ide diagram)

$2Br^-(aq) \rightarrow Br_2(aq) + 2e \quad E^\theta = -0.13$(reverse lower E^θ to derive cell reaction / representation)

Overall E^θ = **E^θ higher- E^θ lower / E^θ RHS - E^θ LHS/ E^θ oxidized- E^θ reduced**

Substituting:

Overall E^θ = $-0.13 - (-1.36) =$ **+1.23V**

Overall redox equation:

$2Br^-(aq) + Cl_2(g) \rightarrow 2Cl^-(aq) + Br_2(aq) \quad E^\theta = +1.23V$

Overall conventional cell representation:

$Cl_2(g) / 2Cl^-(aq)\ 1M, // 1M, 2Br^-(aq) / Br_2(aq) \quad E^\theta = +1.23V$

Chlorine displaces bromine from bromine water. When chlorine gas is thus bubbled in bromine water, the pale **green** colour **fades** as displacement takes place and a **brown** solution containing dissolved bromine liquid is **formed**. This reaction is feasible /possible because the overall redox reaction has a **positive** E^θ value.

e)Strongest oxidizing agent and the strongest reducing agent.

From the table above:

$2e + F_2(g) \rightarrow 2F^-(aq) \quad E^\theta = +2.87V$(highest E^θ /strongest oxidizing agent)

$2e + 2K^+(aq) \rightarrow 2K(aq) \quad E^\theta = -2.92V$(lowest E^θ/ strongest reducing agent)

$2K(aq) \rightarrow 2K^+(aq) + 2e \quad E^\theta = +2.92V$ (reverse lower E^θ to derive cell reaction / representation)

Overall E^θ = **E^θ higher- E^θ lower / E^θ RHS - E^θ LHS/ E^θ oxidized- E^θ reduced**

Substituting:

Overall E^θ = $+2.87 - (-2.92) =$ **+5.79V**

Overall redox equation:

$F_2(g) + 2K(s) \rightarrow 2F^-(aq) + 2K^+(aq) \quad E^\theta = +5.79V$

Overall conventional cell representation:

$2K(s) / 2K^+(aq), 1M, // 1M, 2F^-(aq) / F_2(g) \quad E^\theta = +5.79V$

The redox reactions in an electrochemical/voltaic is commercially applied to make the:

(a)Dry /primary/Laclanche cell.

(b)Wet /secondary /accumulators.

(a)Dry/primary/Laclanche cell

Examine a used dry cell.

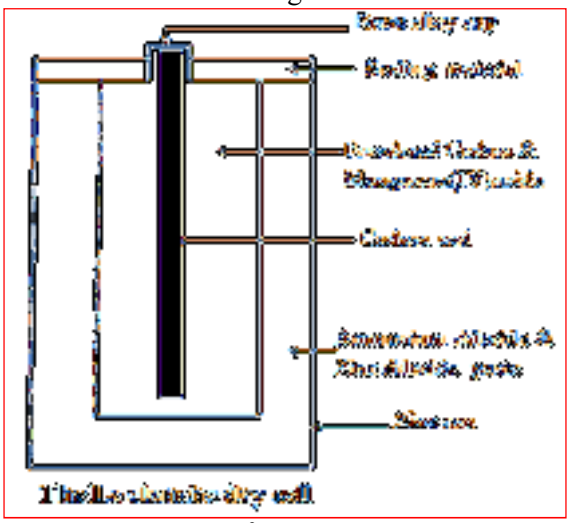
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Note the positive and the negative terminal of the cell. Carefully using a knife cut a cross section from one terminal to the other.
The dry cell consist of a **Zinc can** containing a **graphite rod** at the **centre** surrounded by a **paste** of;

-Ammonium chloride
-Zinc chloride
-powdered manganese (IV) oxide mixed with Carbon.

Zinc acts/serve as the negative terminal where it ionizes/dissociates:



$Zn(s) \rightarrow Zn^{2+}(aq) + 2e$

Ammonium ions in ammonium chloride serve as the positive terminal where it is converted to ammonia gas and hydrogen gas.

$2NH_4^+(aq) + 2e \rightarrow 2NH_3(g) + H_2(g)$

Ammonia forms a complex salt / compound /$(Zn(NH_3)_4)^{2+}$ (aq) / tetramminezinc(II) complex with the Zinc chloride in the paste.
Manganese (IV) oxide oxidizes the hydrogen produced at the electrodes to water preventing any bubbles from coating the carbon terminal which would reduce the efficiency of the cell.
Ammonium chloride is used as paste because the solid does not conduct electricity because the ions are fused/not mobile.
Since the reactants are used up, the dry /primary /Laclanche cell cannot provide continous supply of electricity.The process of restoring the reactants is called **recharging.**

b)Wet/Secondary/Accumulators

1. Wet/Secondary/Accumulators are **rechargeable** unlike dry /primary /Laclanche cells.Wet/Secondary/Accumulators are made up of:

(i)**Lead** plate that forms the **negative** terminal

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(ii)**Lead(IV) oxide** that forms the **positive** terminal

2.The two electrodes are dipped in concentrated sulphuric(VI) acid of a relative density 1.2/1.3

3.At the negative terminal,lead ionizes /dissolves;

$Pb(s) \rightarrow Pb^{2+} + 2e$

4.At the positive terminal,

(i) Lead(IV) oxide **reacts** with the hydrogen ions in sulphuric(VI)acid to form Pb^{2+} (aq) ions;

$PbO_2(s) + 4H^+(aq) + 2e \rightarrow Pb^{2+}(aq) + H_2O(l)$

(ii) Pb^{2+} (aq) ions formed **instantly** react with sulphate (VI) ions/ SO_4^{2-} (aq) from sulphuric (VI)acid to form **insoluble** Lead(II) sulphate (VI).

$Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$

5.The **overall** cell reaction is called **discharging**

$PbO_2(s) + Pb(s) + 4H^+(aq) + 2SO_4^{2-}(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$ $E^\theta = +2.0V$

6.The insoluble Lead(II) sulphate (VI) formed should not be left for long since fine Lead(II) sulphate (VI) will change to a course non-reversible and inactive form making the cell less efficient.

As the battery discharges ,lead and lead(IV)oxide are depleted/finished/reduced and the concentration of sulphuric(VI)acid decreases.

7. During **recharging**, the electrode reaction is reversed as below:

$2PbSO_4(s) + 2H_2O(l) \rightarrow PbO_2(s) + Pb(s) + 4H^+(aq) + 2SO_4^{2-}(aq)$

8. A car battery has six Lead-acid cells making a total of 12 volts.

(iii)ELECTROLYSIS (ELECTROLYTIC CELL)

1.Electrolysis is defined simply as the **decomposition** of a **compound** by an electric current/**electricity**.

A compound that is decomposed by an electric current is called an electrolyte.

Some electrolytes are **weak** while others are **strong**.

2.Strong electrolytes are those that are fully ionized/dissociated into (many) ions.

Common strong electrolytes include:

(i)all **mineral** acids

(ii)all strong **alkalis**/sodium hydroxide/potassium hydroxide.

(iii)all soluble **salts**

3.Weak electrolytes are those that are partially/partly ionized/dissociated into (few) ions.

Common weak electrolytes include:

(i)all **organic** acids

(ii)all **bases** except sodium hydroxide/potassium hydroxide.

(iii)**Water**

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4. A compound that is **not** decomposed by an electric current is called non-electrolyte. Non-electrolytes are those compounds /substances that exist as molecules and thus cannot ionize/dissociate into(any) ions .
Common non-electrolytes include:

(i) most organic solvents (e.g. petrol/paraffin/benzene/methylbenzene/ethanol)

(ii)all hydrocarbons(alkanes /alkenes/alkynes)

(iii)Chemicals of life(e.g. proteins, carbohydrates, lipids, starch, sugar)

5. An electrolytes in **solid** state have **fused** /joined ions and therefore do **not** conduct electricity but the **ions** (cations and anions) are **free** and **mobile** in **molten** and **aqueous** (solution, dissolved in water) state.

6.During electrolysis, the free ions are attracted to the **electrodes**. An electrode is a rod through which current enter and leave the electrolyte during electrolysis. An electrode that does not influence/alter the products of electrolysis is called an **inert electrode.**

Common inert electrodes include:

(i)**Platinum**

(ii)**Carbon graphite**

Platinum is not usually used in a school laboratory because it is very **expensive**. Carbon graphite is **easily**/readily and **cheaply** available (from used dry cells).

7.The **positive** electrode is called **Anode**.The anode is the electrode through which **current enter** the electrolyte/**electrons leave** the electrolyte

8.The **negative** electrode is called **Cathode**. The cathode is the electrode through which **current leave** the electrolyte / **electrons enter** the electrolyte

9. During the electrolysis, free **anions** are attracted to the **anode** where they **lose** /**donate** electrons to form **neutral** atoms/molecules. i.e.

$M(l) \rightarrow M^+(l) + e$ (for cations from molten electrolytes)

$M(s) \rightarrow M^+(aq) + e$ (for cations from electrolytes in aqueous state / solution / dissolved in water)

The neutral atoms /molecules form the **products** of electrolysis at the anode. This is called **discharge** at anode

10. During electrolysis, free **cations** are attracted to the **cathode** where they **gain** /**accept/acquire** electrons to form **neutral** atoms/molecules.

$X^+(aq) + 2e \rightarrow X(s)$ (for cations from electrolytes in aqueous state / solution / dissolved in water)

$2X^+(l) + 2e \rightarrow X(l)$ (for cations from molten electrolytes)

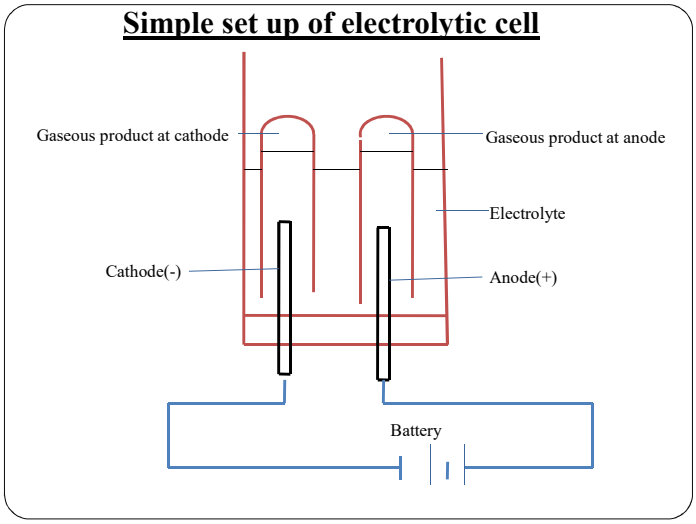
The neutral atoms /molecules form the **products** of electrolysis at the cathode. This is called **discharge** at cathode.

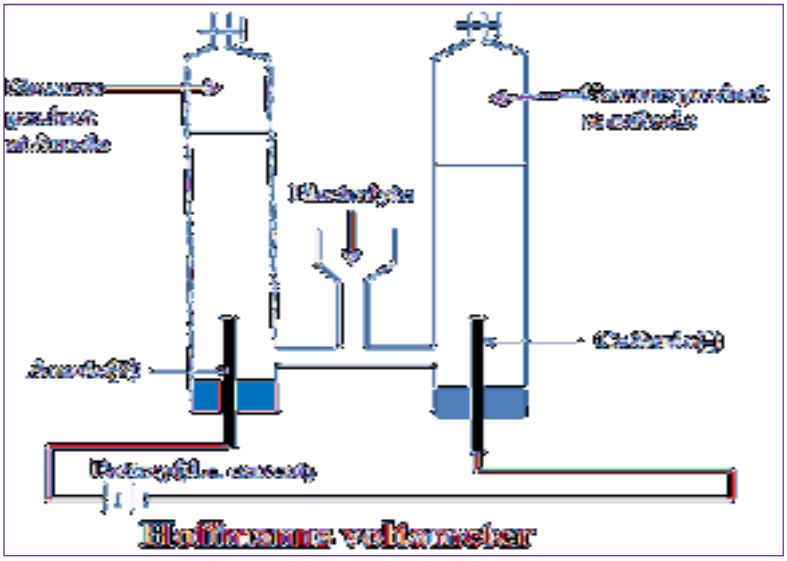
11. The below set up shows an electrolytic cell.

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Simple set up of electrolytic cell

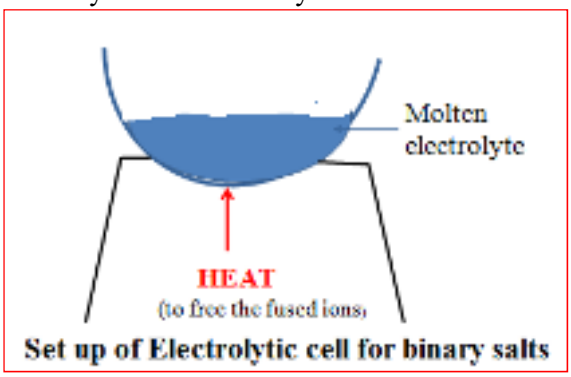




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12. For a compound /salt containing only two ion/binary salt the products of electrolysis in an electrolytic cell can be determined as in the below examples:



Set up of Electrolytic cell for binary salts

a)To determine the products of electrolysis of molten Lead(II)chloride

(i)Decomposition of electrolyte into free ions;

$PbCl_2$ (l) -> Pb^{2+}(l) + $2Cl^-$(l)

(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode(-);

Pb^{2+}(l) + 2e -> Pb (l)

(Cation / Pb^{2+} gains / accepts / acquires electrons to form free **atom)**

(iii)At the anode/positive electrode(+);

$2Cl^-$(l) -> Cl_2 (g) + 2e

(Anion / Cl^- donate/lose electrons to form free **atom** then a gas **molecule)**

(iv)Products of electrolysis therefore are;

I.At the cathode grey beads /solid lead metal.

II.At the anode pale green chlorine gas.

b)To determine the products of electrolysis of molten Zinc bromide

(i)Decomposition of electrolyte into free ions;

$ZnBr_2$ (l) -> Zn^{2+}(l) + $2Br^-$(l)

(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode(-);

Zn^{2+}(l) + 2e -> Zn(l)

(Cation / Zn^{2+} gains / accepts / acquires electrons to form free **atom)**

(iii)At the anode/positive electrode(+);

$2Br^-$(l) -> Br_2 (g) + 2e

(Anion / Br^- donate/lose electrons to form free **atom** then a liquid **molecule** which change to **gas** on heating)

(iv)Products of electrolysis therefore are;

I.At the cathode grey beads /solid Zinc metal.

II.At the anode **red** bromine **liquid** / **red/brown** bromine **gas.**

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c)To determine the products of electrolysis of molten sodium chloride

(i)Decomposition of electrolyte into free ions;

$NaCl\ (l) \rightarrow Na^{+}(l) + Cl^{-}(l)$

(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode(-);

$2Na^{+}(l) + 2e \rightarrow Na\ (l)$

(Cation / Na^{+} gains / accepts / acquires electrons to form free **atom)**

(iii)At the anode/positive electrode(+);

$2Cl^{-}(l) \rightarrow Cl_2\ (g) + 2e$

(Anion / Cl^{-} donate/lose electrons to form free **atom** then a gas **molecule)**

(iv)Products of electrolysis therefore are;

I.At the cathode grey beads /solid sodium metal.

II.At the anode pale green chlorine gas.

d)To determine the products of electrolysis of molten Aluminium (III)oxide

(i)Decomposition of electrolyte into free ions;

$Al_2O_3\ (l) \rightarrow 2Al^{3+}(l) + 3O^{2-}(l)$

(Compound decomposed into free cation and anion in **liquid** state)

(ii)At the cathode/negative electrode(-);

$4Al^{3+}\ (l) + 12e \rightarrow 4Al\ (l)$

(Cation / Al^{3+} gains / accepts / acquires electrons to form free **atom)**

(iii)At the anode/positive electrode(+);

$6O^{2-}(l) \rightarrow 3O_2\ (g) + 12e$

(Anion /$6O^{2-}$ donate/lose 12 electrons to form free **atom** then three gas **molecule)**

(iv)Products of electrolysis therefore are;

I.At the cathode grey beads /solid aluminium metal.

II.At the anode colourless gas that relights/rekindles glowing splint.

13. For a compound /salt mixture containing **many** ions in an electrolytic cell, the **discharge** of ions in the cell depend on the following **factors**:

a) Position of cations and anions in the electrochemical series

1. Most electropositive cations require more energy to reduce (gain electrons) and thus not readily discharged. The **higher** elements /metals in the electrochemical series the **less easily**/**readily** it is **discharged** at the cathode in the electrolytic cell.

Table I showing the relative ease of discharge of **cations** in an electrolytic cell

$K^{+}(aq) + e \rightarrow K(s)$	(**least** readily/easily discharged)
$Na^{+}(aq) + e \rightarrow Na(s)$	
$Ca^{2+}(aq) + 2e \rightarrow Ca(s)$	
$Mg^{2+}(aq) + 2e \rightarrow Mg(s)$	
$Al^{3+}(aq) + 3e \rightarrow Al(s)$	
$Zn^{2+}(aq) + 2e \rightarrow Zn(s)$	
$Fe^{2+}(aq) + 2e \rightarrow Fe(s)$	

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$Pb^{2+}(aq)$ + 2e -> Pb(s)
$2H^{+}(aq)$ + 2e -> $H_2(g)$ (hydrogen is usually "metallic")
$Cu^{2+}(aq)$ + 2e -> Cu(s)
$Hg^{2+}(aq)$ + 2e -> Hg(s)
$Ag^{+}(aq)$ + e -> Ag(s) (**most** readily/easily discharged)

2.The **OH^-** ion is the **most** readily/easily **discharged** anion . All the other anionic **radicals**(SO_4^{2-} ,SO_3^{2-} ,CO_3^{2-} ,HSO_4^- ,HCO_3^- ,NO_3^- ,PO_4^{3-})are not/never discharged. The ease of discharge of halogen ions increase down the group.

Table II showing the relative ease of discharge of **anions** in an electrolytic cell

$4OH^-(aq)$ -> $2H_2O(l)$ + $O_2(g)$ + 4e (**most** readily/easily discharged)
$2I^-(aq)$ -> $I_2(aq)$ + 2e
$2Br^-(aq)$ -> $Br_2(aq)$ + 2e
$2Cl^-(aq)$ -> $Cl_2(aq)$ + 2e
$2F^-(aq)$ -> $F_2(aq)$ + 2e

SO_4^{2-} ,SO_3^{2-} ,CO_3^{2-} ,HSO_4^- ,HCO_3^- ,NO_3^- ,PO_4^{3-} **not/never/rarely** discharged.

3.(a)When **two** or more **cations** are attracted to the **cathode**, the ion **lower** in the electrochemical series is discharged **instead** of that which is **higher** as per the table I above. This is called selective/preferential discharge at cathode.

(b)When **two** or more **anions** are attracted to the **anode**, the ion **higher** in the electrochemical series is discharged **instead** of that which is **lower** as per the table I above. This is called selective/preferential discharge at anode.

4.The following experiments shows the influence /effect of selective/preferential discharge on the products of electrolysis:

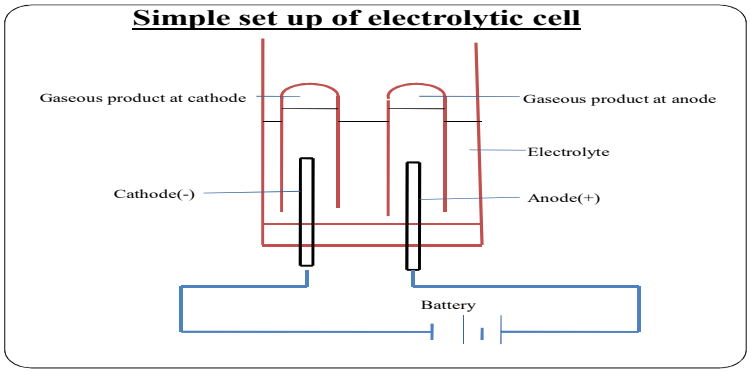
(i)Electrolysis of acidified water/dilute sulphuric(VI) acid

Fill the Hoffmann voltameter with dilute sulphuric(VI) acid. Connect the Hoffmann voltameter to a d.c. electric supply. Note the observations at each electrode.

Electrolytic cell set up during electrolysis of acidified water/dilute sulphuric(VI) acid

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Answer the following questions:

I. Write the equation for the decomposition of the electrolytes during the electrolytic process.

$H_2O(l)$ -> $OH^-(aq)$ + $H^+(aq)$

$H_2SO_4(aq)$ -> $SO_4^{2-}(aq)$ + $2H^+(aq)$

II. Name the ions in acidified water that are attracted/move to:

Cathode- $H^+(aq)$ from **either** sulphuric(VI) acid (H_2SO_4) or water (H_2O)

Anode- $SO_4^{2-}(aq)$ from sulphuric (VI) acid (H_2SO_4) and **OH^- (aq)** from water (H_2O)

III. Write the equation for the reaction during the electrolytic process at the:

Cathode $4H^+(aq)$ + 4e -> $2H_2(g)$

Anode $4OH^-(aq)$ -> $2H_2O(l)$ + $O_2(g)$ + 4e

($4OH^-$ ions selectively discharged instead of SO_4^{2-} ions at the anode)

IV. Name the products of electrolysis of acidified water.

Cathode-Hydrogen gas (colourless gas that extinguishes burning splint with explosion/ "pop" sound

Anode-Oxygen gas (colourless gas that relights /rekindles glowing splint)

V. Explain the difference in volume of products at the cathode and anode.

The **four**(4) electrons donated/lost by **OH^-** ions to form **1** molecule/**1**volume/**1**mole of **oxygen** (**O_2**)gas at the **anode** are gained/acquired/accepted by the four **$H^+(aq)$** ions to form **2** molecule/**2**volume/**2**mole of **Hydrogen** (**H_2**)gas at the **cathode.**

The volume of **Oxygen** gas at the anode is thus a **half** the volume of **Hydrogen** produced at the cathode/ The volume of **Hydrogen** gas at the cathode is thus a **twice** the volume of **Oxygen** produced at the anode.

VI. Why is electrolysis of dilute sulphuric(VI) acid called "electrolysis of (acidified) water"?

The ratio of **H_2 (g)**: **O_2 (g)** is **2:1** as they are combined in water. This implies/means that water in the electrolyte is being decomposed into hydrogen and Oxygen gases. The electrolysis of dilute sulphuric acid is therefore called "electrolysis of acidified water."

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VI. Explain the changes in concentration of the electrolyte during electrolysis of acidified water"

The concentration of dilute sulphuric (VI) acid **increases.** Water in the electrolyte is decomposed into Hydrogen and Oxygen gases that escape. The concentration /mole of acid present in a given volume of solution thus continue increasing/rising.

(ii)Electrolysis of Magnesium sulphate(VI) solution

Fill the Hoffmann voltameter with dilute sulphuric(VI) acid. Connect the Hoffmann voltameter to a d.c. electric supply. Note the observations at each electrode.

Answer the following questions:

I. Write the equation for the decomposition of the electrolytes during the electrolytic process.

$H_2O(l) \rightarrow OH^-(aq) + H^+(aq)$

$MgSO_4(aq) \rightarrow SO_4^{2-}(aq) + Mg^{2+}(aq)$

II. Name the ions in Magnesium sulphate(VI) solution that are attracted/move to:

Cathode- $Mg^{2+}(aq)$ from Magnesium sulphate(VI) solution ($MgSO_4$) and **$H^+(aq)$** from water (H_2O)

Anode- $SO_4^{2-}(aq)$ from Magnesium sulphate(VI) solution ($MgSO_4$) and **$OH^-(aq)$** from water (H_2O)

III. Write the equation for the reaction during the electrolytic process at the:

Cathode $4H^+(aq) + 4e \rightarrow 2H_2(g)$

H^+ ions selectively discharged instead of Mg^{2+} ions at the cathode)

Anode $4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e$

($4OH^-$ ions selectively discharged instead of SO_4^{2-} ions at the anode)

IV. Name the products of electrolysis of Magnesium sulphate(VI) solution

Cathode-Hydrogen gas (colourless gas that extinguishes burning splint with explosion/ "pop" sound

Anode-Oxygen gas (colourless gas that relights /rekindles glowing splint)

V. Explain the difference in volume of products at the cathode and anode.

The **four**(4) electrons donated/lost by **OH^-** ions to form **1** molecule/**1**volume/**1**mole of **oxygen** (**O_2**)gas at the **anode** are gained/acquired/accepted by the four **$H^+(aq)$** ions to form **2** molecule/**2**volume/**2**mole of **Hydrogen** (**H_2**)gas at the **cathode.** The volume of **Oxygen** gas at the anode is thus a **half** the volume of **Hydrogen** produced at the cathode/ The volume of **Hydrogen** gas at the cathode is thus a **twice** the volume of **Oxygen** produced at the anode.

VI. Explain the changes in concentration of the electrolyte during electrolysis of Magnesium sulphate(VI) solution

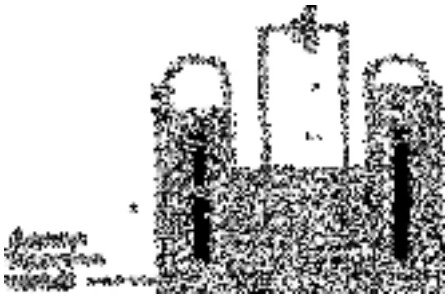
The concentration of dilute Magnesium sulphate(VI) solution **increases.**

The ratio of **H_2 (g)**: **O_2 (g)** is **2:1** as they are combined in water.

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Water in the electrolyte is decomposed into Hydrogen and Oxygen gases that escape as products.
The concentration /mole of acid present in a given volume of Magnesium sulphate(VI) solution thus continue increasing/rising.
The set – up below was used during the electrolysis of aqueous magnesium sulphate using inert electrodes.



Name a suitable pair of electrodes for this experiment
Identify the ions and cations in the solution
On the diagram label the cathode
Write ionic equations for the reactions that took place at the anode.
Explain the change that occurred to the concentration of magnesium sulphate solution during the experience.
During the electrolysis a current of 2 amperes was passed through the solution for 4 hours. Calculate the volume of the gas produced at the anode.(1 faraday 96500 coulombs and volume of a gas at room temperature is 24000cm3)
One of the uses of electrolysis is electroplating
What is meant by electroplating?
Give tow reasons why electroplating is necessary.

b) Concentration of the electrolytes

1.**High** concentrations of cations and/or anions at the electrodes **block** the ion/s that is likely to be discharged at the electrode. This is called **over voltage**. A concentrated solution therefore produces different products of electrolysis from a dilute one.
2. The following experiments show the influence/effect of concentration of electrolyte on the products of electrolysis.

(i)Electrolysis of dilute and concentrated(brine)sodium chloride solution

I. Dissolve about 0.5 g of pure sodium chloride crystals in 100cm3 of water. Place the solution in an electrolytic cell. Note the observations at each electrode for 10 minutes. Transfer the set up into a **fume chamber/open** and continue to make observations for a further 10 minute.
Answer the following questions:

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I. Write the equation for the decomposition of the electrolytes during the electrolytic process.

$H_2O(l)$ -> $OH^-(aq)$ + $H^+(aq)$

$NaCl(aq)$ -> $Cl^-(aq)$ + $Na^+(aq)$

II. Name the ions in sodium chloride solution that are attracted/move to:

Cathode- $Na^+(aq)$ from Sodium chloride solution (NaCl) and **$H^+(aq)$** from water (H_2O)

Anode- $Cl^-(aq)$ from sodiumchloride solution (NaCl) and **OH^- (aq)** from water (H_2O)

III. Write the equation for the reaction during the electrolytic process at the:

Cathode $4H^+(aq) + 4e \rightarrow 2H_2(g)$

H^+ ions selectively discharged instead of Na^+ ions at the cathode)

Anode $4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e$

($4OH^-$ ions selectively discharged instead of Cl^- ions at the anode)

IV. Name the products of electrolysis of dilute sodium chloride solution

Cathode-Hydrogen gas (colourless gas that extinguishes burning splint with explosion/ "pop" sound

Anode-Oxygen gas (colourless gas that relights /rekindles glowing splint)

V. Explain the difference in volume of products at the cathode and anode.

Four(4) electrons donated/lost by **OH^-** ions to form **1** molecule/**1**volume/**1**mole of **oxygen** (**O_2**)gas at the **anode** are gained/acquired/accepted by four **$H^+(aq)$** ions to form **2** molecule/**2**volume/**2**mole of **Hydrogen** (**H_2**)gas at the **cathode.**

The volume of **Oxygen** gas at the anode is **half** the volume of **Hydrogen** produced at the cathode/ The volume of **Hydrogen** gas at the cathode is **twice** the volume of **Oxygen** produced at the anode.

VI. Explain the changes in concentration of the electrolyte during electrolysis of sodium chloride solution

The concentration of dilute sodium chloride solution **increases.**

The ratio of **H_2 (g)**: **O_2 (g)** is **2:1** as they are combined in water. Water in the electrolyte is decomposed into Hydrogen and Oxygen gases that escape as products. The concentration /moles of salt present in a given volume of sodium chloride solution continue increasing/rising.

II. Dissolve about 20 g of pure sodium chloride crystals in 100cm3 of water. Place the solution in an electrolytic cell. Note the observations continuously at each electrode for 30 minutes in a **fume chamber/open**.

Answer the following questions:

I. Write the equation for the decomposition of the electrolytes during the electrolytic process.

$H_2O(l)$ -> $OH^-(aq)$ + $H^+(aq)$

$NaCl(aq)$ -> $Cl^-(aq)$ + $Na^+(aq)$

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II. Name the ions in sodium chloride solution that are attracted/move to:
Cathode- Na^+(aq) from Sodium chloride solution (NaCl) and **H^+(aq)** from water (H_2O)
Anode- Cl^-(aq) from sodium chloride solution (NaCl) and **OH^- (aq)** from water (H_2O)
III. Write the equation for the reaction during the electrolytic process at the:
Cathode $2H^+(aq) + 2e \rightarrow H_2(g)$
H^+ ions selectively discharged instead of Na^+ ions at the cathode)
Anode $2Cl^-(aq) \rightarrow Cl_2(g) + 4e$
(Cl^- ions with a **higher** concentration **block** the discharge of OH^- ions at the anode)
IV. Name the products of electrolysis of concentrated sodium chloride solution/brine
Cathode-Hydrogen gas (colourless gas that extinguishes burning splint with explosion/ "pop" sound
Anode-Chlorine gas(pale green gas that bleaches damp/moist/wet litmus papers)
V. Explain the difference in volume of products at the cathode and anode.
Two (2) electrons donated/lost by **Cl^-** ions to form **1** molecule/**1**volume/**1**mole of **Chlorine** (**Cl_2**)gas at the **anode** are gained/acquired/accepted by two **H^+(aq)** ions to form **1** molecule/**1**volume/**1**mole of **Hydrogen** (**H_2**)gas at the **cathode.**
The volume of **Chlorine** gas at the anode is **equal to** the volume of **Hydrogen** produced at the cathode/ The volume of **Hydrogen** gas at the cathode is **equal to** the volume of **Chlorine** produced at the anode.
VI. Explain the changes in concentration of the electrolyte during electrolysis of concentrated sodium chloride solution/brine
The concentration of concentrated sodium chloride solution/brine **increases.**
The ratio of **Cl_2 (g)**: **H_2 (g)** is **1:1** as they are combined in water.
Water in the electrolyte is decomposed into only Hydrogen gas that escapes as products at cathode.
The concentration /moles of OH^- (aq) and Na^+ ion (as NaOH) present in a given volume of electrolyte continue increasing/rising.
This makes the electrolyte strongly alkaline with **high pH.**
As the electrolysis of brine continues the concentration of **Cl^-** ions **decrease** and **oxygen** gas start being **liberated** at **anode**.
The electrolyte pH is thus lowered and the concentration of brine starts again increasing.
(ii)Electrolysis of dilute and concentrated Hydrochloric acid solution
I. Prepare about 50cm3 of 0.05 M of dilute Hydrochloric acid in 100cm3 solution. Place the solution in an electrolytic cell. Note the observations at each electrode for 10 minutes.
Answer the following questions:

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I. Write the equation for the decomposition of the electrolytes during the electrolytic process.

$H_2O(l)$ -> $OH^-(aq)$ + $H^+(aq)$

$HCl(aq)$ -> $Cl^-(aq)$ + $H^+(aq)$

II. Name the ions in dilute Hydrochloric acid solution that are attracted/move to:

Cathode- $H^+(aq)$ from dilute Hydrochloric acid (HCl) and **$H^+(aq)$** from water (H_2O)

Anode- $Cl^-(aq)$ from dilute Hydrochloric acid (HCl) and **OH^- (aq)** from water (H_2O)

III. Write the equation for the reaction during the electrolytic process at the:

Cathode $4H^+(aq) + 4e \rightarrow 2H_2(g)$

H^+ ions selectively discharged instead of Na^+ ions at the cathode)

Anode $4OH^-(aq) \rightarrow H_2O(l) + O_2 + 4e$

($4OH^-$ ions selectively discharged instead of Cl^- ions at the anode)

IV. Name the products of electrolysis of dilute Hydrochloric acid

Cathode-Hydrogen gas (colourless gas that extinguishes burning splint with explosion/ "pop" sound

Anode-Oxygen gas (colourless gas that relights /rekindles glowing splint)

V. Explain the difference in volume of products at the cathode and anode.

Four(4) electrons donated/lost by **OH^-** ions to form **1** molecule/**1**volume/**1**mole of **oxygen** (**O_2**)gas at the **anode** are gained/acquired/accepted by four **$H^+(aq)$** ions to form **2** molecule/**2**volume/**2**mole of **Hydrogen** (**H_2**)gas at the **cathode.**

The volume of **Oxygen** gas at the anode is **half** the volume of **Hydrogen** produced at the cathode/ The volume of **Hydrogen** gas at the cathode is **twice** the volume of **Oxygen** produced at the anode.

VI. Explain the changes in concentration of the electrolyte during electrolysis of dilute Hydrochloric acid

The concentration of dilute Hydrochloric acid **increases.**

The ratio of **H_2 (g)**: **O_2 (g)** is **2:1** as they are combined in water. Water in the electrolyte is decomposed into Hydrogen and Oxygen gases that escape as products. The concentration /moles of HCl present in a given volume of dilute Hydrochloric acid continue increasing/rising.

II. Prepare about 50cm3 of 2M of Hydrochloric acid in 100cm3 solution. Place the solution in an electrolytic cell. Note the observations at each electrode for 30 minutes

CautionThis experiment should be done in the open/fume chamber.

Answer the following questions:

I. Write the equation for the decomposition of the electrolytes during the electrolytic process.

$H_2O(l)$ -> $OH^-(aq)$ + $H^+(aq)$

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$HCl(aq) \rightarrow Cl^-(aq) + H^+(aq)$

II. Name the ions in 2M Hydrochloric acid solution that are attracted/move to:
Cathode- $H^+(aq)$ from dilute Hydrochloric acid (HCl) and **$H^+(aq)$** from water (H_2O)
Anode- $Cl^-(aq)$ from dilute Hydrochloric acid (HCl) and **OH^- (aq)** from water (H_2O)
III. Write the equation for the reaction during the electrolytic process at the:
Cathode $4H^+(aq) + 4e \rightarrow 2H_2(g)$
H^+ ions selectively discharged instead of Na^+ ions at the cathode)
Anode $2Cl^-(aq) \rightarrow Cl_2 + 2e$
(OH^- ions concentration is **low**.Cl^- ions concentration is **higher** at the anode thus cause **over voltage/block** discharge of OH^- ions)
IV. Name the products of electrolysis of 2M Hydrochloric acid
Cathode-Hydrogen gas (colourless gas that extinguishes burning splint with explosion/ "pop" sound
Anode-Chlorine gas (Pale green gas that bleaches blue/red moist/wet/damp litmus papers)
V. Explain the difference in volume of products at the cathode and anode.
Two(2) electrons donated/lost by **Cl^-** ions to form **1** molecule/**1**volume/**1**mole of **Chlorine** (**Cl_2**)gas at the **anode** are gained/acquired/accepted by two **$H^+(aq)$** ions to form **1** molecule/**1**volume/**1**mole of **Hydrogen** (**H_2**)gas at the **cathode.**
The volume of **Chlorine** gas at the anode is **equal to** the volume of **Hydrogen** produced at the cathode/ The volume of **Hydrogen** gas at the cathode is **twice** the volume of **Chlorine** produced at the anode.
VI. Explain the changes in concentration of the electrolyte during electrolysis of 2M Hydrochloric acid
The concentration of Hydrochloric acid **decreases.**
The ratio of **H_2 (g)**: **Cl_2 (g)** is **1:1** as they are combined in Hydrochloric acid.
Water in the electrolyte is decomposed only into Hydrogen gas that escapes as products at the cathode.
There is a net accumulation of excess OH^- (aq) ions in solution.
This makes the electrolyte strongly alkaline with high pH.

c) Nature of electrodes used in the electrolytic cell

Inert electrodes (carbon-graphite and platinum) **do not** alter the expected products of electrolysis in an electrolytic cell. If another/different electrode is used in the electrolytic cell it alters/influences/changes the expected products of electrolysis. The examples below illustrate the influence of the nature of electrode on the products of electrolysis:

(i)Electrolysis of copper(II) sulphate(VI) solution

I. Using carbon-graphite electrodes

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Weigh Carbon -graphite electrodes. Record the masses of the electrodes in table I below. Place the electrodes in 1M copper(II) sulphate(VI) solution in a beaker. Set up an electrolytic cell.
Close the switch and pass current for about 20 minutes. Observe each electrode and any changes in electrolyte. Remove the electrodes from the electrolyte. Wash with acetone/propanone and allow them to dry. Reweigh each electrode.
Sample results

Mass of **cathode before** electrolysis	23.4 g	Mass of **anode before** electrolysis	22.4 g
Mass of **cathode after** electrolysis	25.4 g	Mass of **anode after** electrolysis	22.4 g
Brown solid deposit at the cathode **after** electrolysis	-	Bubbles of colourless gas that relight splint	-
Blue colour of electrolyte **fades**/become less blue	-	Blue colour of electrolyte **fades** /become less blue	-

Answer the following questions:

I. Write the equation for the decomposition of the electrolytes during the electrolytic process.

$H_2O(l) \rightarrow OH^-(aq) + H^+(aq)$

$CuSO_4(aq) \rightarrow SO_4^{2-}(aq) + Cu^{2+}(aq)$

II. Name the ions in 1M copper(II) sulphate(VI) solution that are attracted/move to:

Cathode- Cu^{2+} (aq) from copper(II) sulphate(VI) solution and **H^+(aq)** from water (H_2O)

Anode- SO_4^{2-}(aq) from copper(II) sulphate(VI) solution and **OH^- (aq)** from water (H_2O)

III. Write the equation for the reaction during the electrolytic process at the:

Cathode $2Cu^{2+}(aq) + 4e \rightarrow 2Cu(g)$

Cu^{2+} ions are **lower** than H^+ ions in the electrochemical series therefore selectively discharged at the cathode.)

Anode $4OH^-(aq) \rightarrow H_2O(l) + O_2 + 4e$

(OH^- ions ions are **higher** than SO_4^{2-} ions in the electrochemical series therefore selectively discharged at the cathode.))

IV. Name the products of electrolysis of 1M copper(II) sulphate(VI) solution

Cathode-2 moles of **copper** metal as brown solid coat

Anode-Oxygen gas (Colourless gas that relights /rekindles glowing splint)

V. Explain the changes that take place at the cathode and anode.

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Four(4) electrons donated/lost by **OH^-** ions to form **1** molecule/**1**volume/**1**mole of **Oxygen** (**O_2**)gas at the **anode** are gained/acquired/accepted by two **Cu^{2+}(aq)** ions to form **2** moles of brown **copper** solid that deposit itself at the **cathode.**
The moles of **oxygen** gas at the anode is **equal to** the moles of **copper** produced at the cathode

VI. Explain the changes in electrolyte during electrolysis of 1M copper (II) sulphate(VI) solution.

(i)The **pH** of copper(II) sulphate(VI) solution lowers/**decreases.** The salt becomes **more acidic.** Water in the electrolyte is decomposed only into Oxygen gas (from the OH^- ions) that **escapes** as products at the **anode**. There is a net accumulation of **excess H^+** (aq) ions in solution. This makes the electrolyte strongly **acidic** with **low** pH.

(ii) **Cu^{2+}** (aq) ions are responsible for the **blue** colour of the electrolyte/ copper(II) sulphate (VI) solution. As electrolysis continues, **blue** Cu^{2+} (aq) ions gain electrons to form **brown Copper.** The blue colour of electrolyte therefore fades/become less blue.

(iii)Copper is deposited at the cathode. This increases the mass of the cathode.OH^- ions that produce Oxygen gas at anode come from water. Oxygen escapes out/away without increasing the mass of anode.

II. Using copper electrodes

Weigh clean copper plates electrodes. Record the masses of the electrodes in table I below. Place the electrodes in 1M copper(II) sulphate(VI) solution in a beaker. Set up an electrolytic cell.

Close the switch and pass current for about 20 minutes. Observe each electrode and any changes in electrolyte. Remove the electrodes from the electrolyte. Wash with acetone/propanone and allow them to dry. Reweigh each electrode.

Sample results

Mass of **cathode before** electrolysis	23.4 g	Mass of **anode before** electrolysis	22.4 g
Mass of **cathode after** electrolysis	25.4 g	Mass of **anode after** electrolysis	20.4 g
Brown solid deposit at the cathode **after** electrolysis	-	Anode decrease insize/erodes/wear off	-
Blue colour of electrolyte **remain** blue	-	Blue colour of electrolyte **remain** blue	-

Answer the following questions:

I. Write the equation for the decomposition of the electrolytes during the electrolytic process.

$H_2O(l) \rightarrow OH^-(aq) + H^+(aq)$

$CuSO_4(aq) \rightarrow SO_4^{2-}(aq) + Cu^{2+}(aq)$

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II. Name the ions in 1M copper(II) sulphate(VI) solution that are attracted/move to:

Cathode- Cu^{2+} (aq) from copper(II) sulphate(VI) solution and **H^{+}(aq)** from water (H_2O)

Anode- SO_4^{2-}(aq) from copper(II) sulphate(VI) solution and **OH^{-} (aq)** from water (H_2O)

III. Write the equation for the reaction during the electrolytic process at the:

Cathode $Cu^{2+}(aq) + 2e \rightarrow Cu(s)$

Cu^{2+} ions are **lower** than H^{+} ions in the electrochemical series therefore selectively discharged at the cathode.)

Anode $Cu(s) \rightarrow Cu^{2+}(aq) + 2e$

(Both OH^{-} ions and SO_4^{2-} ions move to the anode but none is discharged. The copper anode itself ionizes/dissolves/dissociate because less energy is used to remove an electron/ionize /dissociate copper atoms than OH^{-} ions.

IV. Name the products of electrolysis of 1M copper(II) sulphate(VI) solution using copper electrodes.

Cathode-1 moles of **copper** metal as brown solid coat (Cathode increase/deposits)

Anode-Anode erodes/decrease in size

V. Explain the changes that take place during the electrolytic process

(i)Cathode

-Cu^{2+} ions are **lower** than **H^{+}** ions in the electrochemical series therefore selectively discharged at the cathode. **Cu^{2+}** ions have greater tendency to accept/gain/acquire electrons to form brown **copper** atoms/solid that deposit itself and increase the mass/size of the **cathode.**The copper deposited at the cathode is **pure**

-H^{+} ions accumulate around the cathode. Electrolyte thus becomes strongly acidic around the cathode.

-Cu^{2+} ions in solution are responsible for the blue colour of electrolyte. Blue colour of electrolyte **fade** around the cathode.

(ii)**Anode**

Copper atom at the anode easily ionizes to release electrons. The anode therefore keeps decreasing in mass/eroding. The amount of copper that dissolve/erode is **equal** to the mass of copper deposited. This is called **electrode ionization.** Electrode ionization is where the anode erodes/decrease and the cathode deposits/increase during electrolysis. The overall **concentration** of the electrolyte remains **constant**

14. In industries electrolysis has the following uses/applications:

(a)Extraction of reactive metals from their ores.

Potassium, sodium ,magnesium, and aluminium are extracted from their ores using electrolytic methods.

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(b)Purifying copper after exraction from copper pyrites ores.

Copper obtained from copper pyrites ores is not pure. After extraction, the copper is refined by electrolysing copper(II)sulphate(VI) solution using the **impure** copper as **anode** and a thin strip of **pure** copper as **cathode**. Electrode ionization take place there:

(i)At the cathode; **Cu^{2+} (aq) + 2e -> Cu(s)** (Pure copper deposits on the strip

(ii)At the anode; **Cu(s) ->Cu^{2+} (aq) + 2e** (impure copper erodes/dissolves)

(c)Electroplating

The label EPNS(**E**lectro **P**lated **N**ickel **S**ilver) on some steel/metallic utensils mean they are plated/coated with silver and/or Nickel to **improve** their **appearance**(**add** their **aesthetic** value)and **prevent**/slow **corrosion**(**rusting** of iron). Electroplating is the process of coating a metal with another metal using an electric current. During electroplating,the **cathode** is made of the metal to be **coated**/impure.

Example:

During the electroplating of a spoon with silver

(i)the spoon/impure is placed as the cathode(negative terminal of battery)

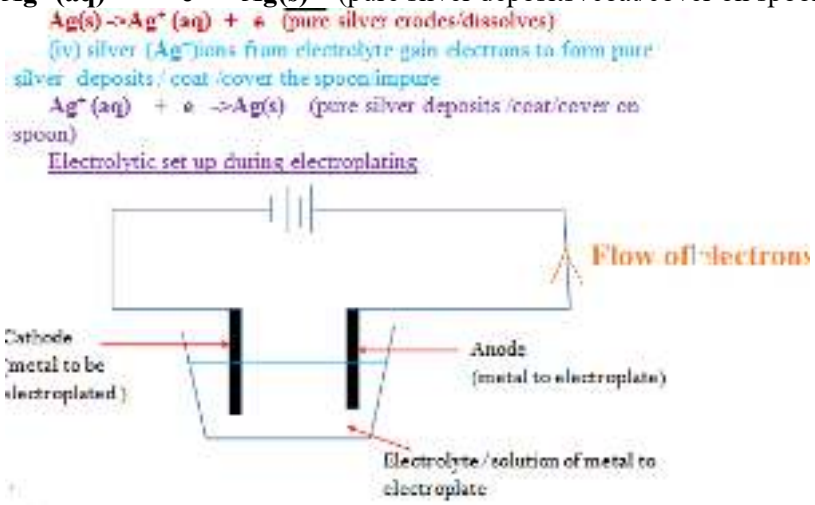
(ii)the pure silver is placed as the anode(positive terminal of battery)

(iii)the pure silver erodes/ionizes/dissociates to release electrons:

Ag(s) ->Ag^{+} (aq) + e (impure silver erodes/dissolves)

(iv) silver (**Ag^{+}**)ions from electrolyte gain electrons to form pure silver deposits / coat /cover the spoon/impure

Ag^{+} (aq) + e ->Ag(s) (pure silver deposits /coat/cover on spoon)



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15. The quantitative **amount** of **products** of electrolysis can be determined by applying **Faradays 1st law of electrolysis**.
Faradays 1st law of electrolysis states that **"the mass/amount of substance liberated/produced/used during electrolysis is directly proportional to the quantity of of electricity passed/used."**
(**a**)The **SI** unit of **quantity** of electricity is the coulomb(C). The coulomb may be defined as the quantity of electricity passed/used when a current of one ampere flow for one second.i.e;

1Coulomb = 1 Ampere x 1Second

The Ampere is the **SI** unit of current(I)
The Second is the **SI** unit of time(t) therefore;
Quantity of electricity(in Coulombs) = Current(**I**) x time(**t**)

Practice examples

1. A current of 2 amperes was passed through an electrolytic cell for 20 minutes. Calculate the quantity of electric charge produced.
Working:
Quantity of electricity(in Coulombs) = Current(**I**) x time(**t**)
Substituting /converting time to second = 2 x (20 x 60)
= **2400 C**

2. A current of 2 amperes was passed through an electrolytic.96500 coulombs of charge were produced. Calculate the time taken.
Working:

$$\text{Time(t) in seconds} = \frac{\text{Quantity of electricity(in Coulombs)}}{\text{Current(I) in amperes}}$$

Substituting $= \frac{96500}{2}$
$=$ **48250 seconds**

3. 96500 coulombs of charge were produced after 10 minutes in an electrolytic cell . Calculate the amount of current used.
Working:

$$\text{Current(I) in amperes} = \frac{\text{Quantity of electricity(in Coulombs)}}{\text{Time(t) in seconds}}$$

Substituting/converting time to second $= \frac{96500}{10 \times 60}$
$=$ **160.8333 Amperes**

(**b**)The quantity of electricity required for **one mole** of **electrons** at the anode/cathode is called the **Faraday constant**(F). It is about **96500 Coulombs.**i.e
The number of **Faradays** used /required is equal to the number of **electrons** used at cathode/anode during the electrolytic process. e.g.
Cu^{2+} require to gain **2** moles of electrons=**2** Faradays =**2** x 96500 coulombs of electricity at the cathode.

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Al^{3+} require to gain **3** moles of electrons=**3** Faradays =**3** x 96500 coulombs of electricity at the cathode
Na^{+} require to gain **1** moles of electrons=**1** Faradays =**1** x 96500 coulombs of electricity at the cathode
$2H^{+}$ require to gain **2** moles of electrons=**2** Faradays =**2** x 96500 coulombs of electricity at the cathode to form 1molecule of hydrogen gas
$2O^{2-}$ require to lose/donate **4** moles of electrons=**4** Faradays =**4** x 96500 coulombs of electricity at the anode to form 1molecule of Oxygen O_2 gas.
$4OH^{-}$ require to lose/donate **4** moles of electrons=**4** Faradays =**4** x 96500 coulombs of electricity at the anode to form 1molecule of Oxygen gas and 2 molecules of water.
(**c**)The mass/amount of products at the cathode/anode is related to the molar mass of the substance and/or the volume of gases at standard/room temperature and pressure as in the below examples:
Practice examples
1.Calculate the mass of copper deposited at the cathode when a steady current of 4.0 amperes is passed through copper(II)sulphate(VI) for 30 minutes in an electrolytic cell. (Cu=63.5, 1F = 96500C)
Working:
Quantity of electricity(in Coulombs) = Current(**I**) x time(**t**)
Substituting /converting time to second = 4 x (30 x 60)
= **7200 C**
Equation at the cathode: $Cu^{2+}(aq) + \mathbf{2e} \rightarrow Cu(s)$
2 mole of electrons = 2 Faradays = 2 x 96500 C produce a mass =molar mass of copper thus;
2 x 96500C -> 63.5 g
72000C -> $\frac{7200 \times 63.5}{2 \times 96500}$ = **2.3689 g** of copper
2.a)If 3.2 g of Lead were deposited when a current of 2.5 amperes was passed through an electrolytic cell of molten Lead(II)bromide for 20 minutes, determine the Faraday constant.(Pb = 207)
Working:
Quantity of electricity (in Coulombs) = Current(**I**) x time(**t**)
Substituting /converting time to second = 2.5 x (20 x 60)
= **3000 C**
If 3.2g of Lead -> 3000C
Then 207 g of Lead -> $\frac{207 \times 3000}{3.2}$ = **194062.5** C
Equation at the cathode: $Pb^{2+}(l) + \mathbf{2e} \rightarrow Pb(l)$
From the equation: 2 moles of electrons = 2 Faradays = 194062.5 C
1mole of electrons = 1 Faraday => 194062.5 = **97031.25 C**

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b)What is the volume of bromine vapour produced at the anode at room temperature(1mole of gas at room temperature and pressure = 24000cm3)

Method 1

Equation at the anode: $Br^-(l) \rightarrow Br_2(g) + 2e$

From the equation: 2 moles of electrons = 2 Faradays = 194062.5 C -> 24000cm3

3000 C -> $\frac{3000 \times 24000}{194062.5}$

=**371.0145cm3**

Method 2

Equation at the anode: $Br^-(l) \rightarrow Br_2(g) + 2e$

Mole ratio of products at Cathode: anode = 1:1

Moles of Lead at cathode = $\frac{3.2}{207}$ = 0.0155moles = moles of Bromine

1 moles of bromine vapour -> 24000cm3

0.0155moles of Bromine -> 0.0155 x 24000 = **372 cm3**

Method 3

Equation at the anode: $Br^-(l) \rightarrow Br_2(g) + 2e$

Ratio of Faradays used to form products at Cathode: anode = 2:2

=> 2 x 97031.25 C produce 24000cm3 of bromine vapour

Then: 3000 C -> $\frac{3000 \times 24000cm3}{2 \times 97031.25}$ = **371.0145cm3**

3.What mass of copper remain from 2.0 at the anode if a solution of copper(II)sulphate(VI) is electrolysed using a current of 1 ampere flowing through an electrolytic cell for 20 minutes.(Cu= 63.5, 1Faraday = 96487 coulombs)

Working:

Quantity of electricity (in Coulombs) = Current(**I**) x time(**t**)

Substituting /converting time to second = 1 x (20 x 60)

= **1200 C**

Equation at the cathode: $Cu^{2+}(aq) + 2e \rightarrow Cu(s)$

2 mole of electrons = 2 Faradays = 2 x 96500 C erode/dissolve a mass =molar mass of copper thus;

2 x 96500C -> 63.5 g

1200C -> $\frac{1200 \times 63.5}{2 \times 96500}$ = **0.3948g** of copper deposited

Mass of copper remaining = Original mass – mass dissolved/eroded

=> 2.0 -0.3948 = **1.6052 g** of copper remain

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4. Calculate the current passed if a mass of 0.234 g of copper is deposited in 4 minutes during electrolysis of a solution of copper (II)sulphate(VI).
(Cu= 63.5 ,1F = 96500C)
Working:
Equation at the cathode: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e$
2 mole of electrons = 2 Faradays = 2 x 96500 C produce a mass =molar mass of copper thus;

63.5 g -> 2 x 96500C
0.234 g -> $\frac{0.234}{63.5} \times 2 \times 96500$ = **711.2126 C**

Current(**I**) in amperes = $\frac{\text{Quantity of electricity(in Coulombs)}}{\text{Time(t) in seconds}}$

Substituting/converting time to second= $\frac{711.2126\ C}{4 \times 60}$

= **2.9634 A**mperes

5. (a)What quantity of electricity will deposit a mass of 2.43 g of Zinc during electrolysis of a solution of Zinc (II)sulphate(VI).
(Zn= 65 ,1F = 96500C)
Working:
Equation at the cathode: $Zn^{2+}(aq) + 2e \rightarrow Zn(s)$

2 mole of electrons = 2 Faradays = 2 x 96500 C erode/dissolve a mass =molar mass of Zinc thus;

65 g -> 2 x 96500
2.43 g -> $\frac{2.43 \times 2 \times 96500}{65}$ = **7215.2308 C**

(b)Calculate the time (in minutes) it would take during electrolysis of the solution of Zinc (II)sulphate(VI) above if a current of 4.0 Amperes is used.
Time(**t**) in seconds = $\frac{\text{Quantity of electricity(in Coulombs)}}{\text{Current(I) in amperes}}$

Substituting = $\frac{7215.2308}{4}$ = 1803.8077 seconds = $\frac{1803.8077}{60}$ = **30.0635** minutes

6.When a current of 1.5 amperes was passed through a cell containing M^{3+} ions of metal M for 15 minutes, the mass at cathode increased by 0.26 g.(Faraday constant = 96500C
a) Calculate the quantity of electricity used.
Quantity of electricity (in Coulombs) = Current(**I**) x time(**t**)
Substituting /converting time to second = 1.5 x (15 x 60)
= **1350 C**

b) Determine the relative atomic mass of metal M
Equation at the cathode: $M^{3+}(aq) + 3e \rightarrow M(s)$

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1350 C of electricity -> 0.26 g of metal M
3 mole of electrons = 3 Faradays = 3 x 96500 C produce a mass =molar mass of M thus;
RAM of M = 0.26 g x 3 x 96500 = **55.7556**(No units)
1350

7.An element "P" has a relative atomic mass 88.When a current of 0.5 amperes was passed through fused chloride of "P" for 32 minutes and 10seconds ,0.44 g of "P" was deposited at the cathode. Determine the charge on an ion of "P"(Faraday constant = 96500C)

Working:
Quantity of electricity (in Coulombs) = Current(**I**) x time(**t**)
Substituting /converting time to second = 0.5 x ((32 x 60) + 10)
= **965C**
0.44 g of metal "P" are deposited by 965C
88g of of metal "P" are deposited by: 88 x 965= **193000** C
0.44
96500 C = 1 mole of electrons = 1 Faradays = single charge
193000 C -> 193000 = **2** moles/Faradays/charges => symbol of ion = $\mathbf{P^{2+}}$
96500

8. During purification of copper by electrolysis 1.48 g of copper was deposited when a current was passed through aqueous copper (II)sulphate(VI) for 2 ½ hours. Calculate the amount of current that was passed. (Cu= 63.5 ,1F = 96500C)

Working:
Equation at the cathode: $Cu^{2+}(aq) + \mathbf{2e}^- \rightarrow Cu(s)$
2 mole of electrons = 2 Faradays = 2 x 96500 C produce a mass =molar mass of copper thus;
63.5 g -> 2 x 96500C
1.48 g -> 1.48 x 2 x 96500 = **4255.1181 C**
63.5
Current(**I**) in amperes = **Q**uantity of electricity(in Coulombs)
Time(**t**) in seconds
Substituting/converting time to second= **4255.1181**C
((2 x 60) + 30) x60
= **0.4728 A**mperes

17. Practically Faraday 1st law of electrolysis can be verified as below.

Verifying Faraday 1st law of electrolysis

Procedure.
Weigh clean copper plates electrodes. Record the masses of the electrodes in table I below. Place the electrodes in 1M copper(II) sulphate(VI) solution in a beaker. Set up an electrolytic cell.

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Close the switch and pass a steady current of 2 amperes by adjusting the rheostat for exactly 20 minutes.Remove the electrodes from the electrolyte. Wash with acetone/ propanone and allow them to dry. Reweigh each electrode.
Sample results

Mass of **cathode before** electrolysis	7.00 g	Mass of **anode before** electrolysis	7.75 g
Mass of **cathode after** electrolysis	8.25 g	Mass of **anode after** electrolysis	6.50 g
Change in mass at cathode **after** electrolysis	1.25 g	Change in mass at anode **after** electrolysis	1.25 g

Answer the following questions:

I. Write the equation for the decomposition of the electrolytes during the electrolytic process.

$H_2O(l) \rightarrow OH^-(aq) + H^+(aq)$

$CuSO_4(aq) \rightarrow SO_4^{2-}(aq) + Cu^{2+}(aq)$

II. Name the ions in 1M copper(II) sulphate(VI) solution that are attracted/move to:

Cathode- Cu^{2+} (aq) from copper(II) sulphate(VI) solution and **H^+(aq)** from water (H_2O)

Anode- SO_4^{2-}(aq) from copper(II) sulphate(VI) solution and **OH^- (aq)** from water (H_2O)

III. Write the equation for the reaction during the electrolytic process at the:

Cathode $Cu^{2+}(aq) + 2e \rightarrow Cu(s)$

Cu^{2+} ions are **lower** than H^+ ions in the electrochemical series therefore selectively discharged at the cathode.)

Anode $Cu(s) \rightarrow Cu^{2+}(aq) + 2e$

(Both OH^- ions and SO_4^{2-} ions move to the anode but none is discharged. The copper anode itself ionizes/dissolves/dissociate as less energy is used to remove an electron/ionize /dissociate copper atoms than OH^- ions.

IV. Name the products of electrolysis of 1M copper(II) sulphate(VI) solution using copper electrodes.

Cathode-1.25 g of **copper** metal as brown solid coat/deposits

Anode-1.25 g of **copper** metal erodes/decrease in size

V. (i)How many moles of electrons are used to deposit/erode one mole of copper metal at the cathode/anode?

From the equation at anode/cathode= **2 moles**

(ii)How many Faradays are used to deposit/erode one mole of copper metal at the cathode/anode?

From the equation at anode/cathode : 2 moles = **2 Faradays**

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(iii)Calculate the quantity of electric charge used

Working:

Quantity of electricity (in Coulombs) = Current(**I**) x time(**t**)

Substituting /converting time to second = 2 x 20 x 60

= **2400C**

VI. (i) Calculate the quantity of electricity required to deposit/erode one mole of copper at the cathode/anode(Cu=63.5)

Since 1.25 g of copper -> 2400C

Then 63.5 g (1mole of copper) -> $\frac{63.5 \times 2400}{1.25}$ = **121920** C

(ii)Determine the Faraday constant from the results in V(i) above

From the equation at;

Cathode $Cu^{2+}(aq) + 2e \rightarrow Cu(s)$

Anode $Cu(s) \rightarrow Cu^{2+}(aq) + 2e$

2 moles = 2 Faradays -> 121920 C

1 moles = 1 Faradays -> $\frac{121920}{2}$ = **60960** C

(iii) The faraday constant obtained above is far lower than theoretical.Explain

-high resistance of the wires used.

-temperatures at 25°C were not kept constant

-plates/electrodes used were not made of pure copper

-plates/electrodes used were not thoroughly clean copper

Further practice

1.**An element P has a relative atomic mass of 88. When a current of 0.5 amperes was passed through the fused chloride of P for 32 minutes and 10 seconds, 0.44g of P were deposited at the cathode. Determine the charge on an ion of P. (1 faraday = 96500 Coulombs).**

2.During electrolysis of aqueous copper (II) sulphate, 144750 coulombs of electricity were used. Calculate the mass of copper metal that was obtained

(Cu = 64 ;1 Faraday = 96500 coulombs) (3 mks)

3.A nitrate of a metal M was electrolysed .1.18 g of metal was deposited when a current of 4 ampheres flow for 16 minutes.Determine the formula of the sulphate(VI)salt of the metal.

(Faraday constant = 96500 , RAM of X = 59.0)

Working

Q = It =>(4 x 16 x 60) = **3840** C

1.18 g of X => 3840 C

59.0 g => $\frac{59.0 \times 3840}{1.18}$ = **192000** C

96500 C = 1Faraday

192000 C= $\frac{192000 \text{ C}}{1}$ x1 = **2F** thus charge of M = $\mathbf{M^{2+}}$

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96500 C

Valency of M is **2** thus formula of sulphate(VI)salt **MSO_4**

4. Below is the results obtained when a current of 2.0ampheres is passed through copper(II)sulphate(VI)solution for 15 minutes during electrolysis using copper electrode.

Initial mass of cathode = 1.0 g

Final mass of cathode = 1.6 g

Change in mass of cathode = 0.60 g

(i)Determine the change in mass at the anode. Explain your answer.

Mass decrease = 0.6g.

Electrode ionization take place where the cathode increase in mass form the erosion of the anode

(ii)Calculate the quantity of electricity required to deposit one mole of copper.(Cu =63.5)

Q =It => 2 x 15 x 60 = 1800 coulombs

Method 1

0.60 g of copper ->1800 coulombs

63.5 g -> $\frac{63.5 \times 1800}{0.60}$ = **190500** Coulombs

Method 2

Moles of Copper = $\frac{\text{Mass}}{\text{Molar mass}}$ => $\frac{0.60}{63.5}$ = **9.4488×10^{-3} moles**

9.4488×10^{-3} moles **->** 1800 coulombs

1 Mole -> $\frac{1 \times 1800 \text{ coulombs}}{9.4488 \times 10^{-3} \text{ moles}}$ = **190500.381** coulombs

(iii)Determine the oxidation number of copper produced at the cathode and hence the formula of its nitrate (V)salt (1 Faraday = 96500 Coulombs)

96500 Coulombs -> 1 Faraday

190500.381 coulombs -> $\frac{190500.381 \text{ coulombs} \times 1}{96500 \text{ Coulombs}}$

= 1.9741 Faradays => **2**F(**whole number**)

Charge of copper = **2+** = **Oxidation number**

=> Valency of copper = **2** hence chemical formula of nitrate (V)salt = **$Cu(NO_3)_2$**

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RADIOACTIVITY

A: INTRODUCTION / CAUSES OF RADIOCTIVITY

Radioactivity is the spontaneous disintegration/decay of an unstable nuclide.
A nuclide is an atom with defined mass number (number of protons and neutrons), atomic number and definite energy.
Radioactivity takes place in the nucleus of an atom unlike chemical reactions that take place in the energy levels involving electrons.
A nuclide is said to be stable if its neutron: proton ratio is equal to one (n/p = 1)
All nuclide therefore try to **attain n/p = 1** by undergoing radioactivity.
Examples
(i)Oxygen nuclide with $^{16}_{8}O$ has **8** neutrons and **8** protons in the nucleus therefore an **n/p = 1** thus **stable** and do **not** decay/disintegrate.
(ii)Chlorine nuclide with $^{35}_{17}Cl$ has **18** neutrons and **17** protons in the nucleus therefore an n/p = 1.0588 thus **unstable** and **decays**/disintegrates to try to attain n/p = 1.
(ii)Uranium nuclide with $^{237}_{92}U$ has 206 neutrons and 92 protons in the nucleus therefore an n/p = 2.2391 thus more unstable than $^{235}_{92}U$ and thus more readily decays / disintegrates to try to attain n/p = 1.
(iii) Chlorine nuclide with $^{37}_{17}Cl$ has **20** neutrons and **17** protons in the nucleus therefore an n/p = 1.1765 thus **more unstable** than $^{35}_{17}Cl$ and thus more **readily** decays / disintegrates to try to attain n/p = 1.
(iv)Uranium nuclide with $^{235}_{92}U$ has **143** neutrons and **92** protons in the nucleus therefore an n/p = 1.5543 thus more **stable** than $^{237}_{92}U$ but also **readily** decays / disintegrates to try to attain n/p = 1.
All **unstable** nuclides naturally try to attain nuclear **stability** with the production of:

(i)alpha(α) particle decay

The alpha (α) particle has the following main characteristic:
i)is **positive**ly charged(like protons)
ii) has mass number **4** and atomic number **2** therefore equal to a charged Helium atom ($^{4}_{2}He^{2+}$**)**
iii) have very **low** penetrating power and thus can be stopped /blocked/shielded by a thin sheet of **paper.**
iv) have **high** ionizing power thus cause a lot of **damage** to living cells.
v) a nuclide undergoing α-decay has its mass number **reduced** by **4** and its atomic number **reduced** by **2**
Examples of alpha decay

$^{210}_{84}Pb \rightarrow {}^{x}_{82}Pb + {}^{4}_{2}He^{2+}$

$\mathbf{^{210}_{84}Pb \rightarrow {}^{206}_{82}Pb + {}^{4}_{2}He^{2+}}$

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$^{226}_{88}Ra \rightarrow {}^{222}_{y}Rn + {}^{4}_{2}He^{2+}$

$\mathbf{^{226}_{88}Ra \rightarrow {}^{222}_{86}Rn + {}^{4}_{2}He^{2+}}$

$^{x}_{y}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He^{2+}$

$\mathbf{^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He^{2+}}$

$^{x}_{y}U \rightarrow {}^{230}_{88}Ra + 2\,{}^{4}_{2}He^{2+}$

$\mathbf{^{238}_{92}U \rightarrow {}^{230}_{88}Ra + 2\,{}^{4}_{2}He^{2+}}$

$^{210}_{84}U \rightarrow {}^{x}_{y}W + 10\,\alpha$

$\mathbf{^{210}_{84}U \rightarrow {}^{170}_{64}W + 10\,\alpha}$

$^{210}_{92}U \rightarrow {}^{x}_{y}W + 6\,\alpha$

$\mathbf{^{210}_{92}U \rightarrow {}^{186}_{80}W + 6\,\alpha}$

(ii)Beta (β) particle decay

The Beta (β) particle has the following main characteristic:

i)is negatively charged(like electrons)

ii)has no mass number and atomic number negative one(-1) therefore equal to a fast moving electron ($^{0}_{-1}e$)

iii) have medium penetrating power and thus can be stopped /blocked/shielded by a thin sheet of aluminium foil.

iv) have medium ionizing power thus cause less damage to living cells than the α particle.

v) a nuclide undergoing β -decay has its mass number **remain** the same and its atomic number **increase** by **1**

Examples of beta (β) decay

1. $^{23}\mathbf{x}\,Na \rightarrow {}^{23}_{12}Mg + {}^{0}_{-1}e$

$\mathbf{^{23}11\,Na \rightarrow {}^{23}_{12}Mg + {}^{0}_{-1}e}$

2. $^{234}\mathbf{x}\,Th \rightarrow {}^{y}_{91}Pa + {}^{0}_{-1}e$

$\mathbf{^{234}90\,Th \rightarrow {}^{y}_{91}Pa + {}^{0}_{-1}e}$

3. $^{207}_{70}Y \rightarrow \mathbf{x}\,{}^{y}Pb + 3^{0}_{-1}e$

$\mathbf{^{207}_{70}Y \rightarrow {}^{207}_{73}Pb + 3^{0}_{-1}e}$

4. $\mathbf{x}\,{}^{y}C \rightarrow {}^{14}_{7}N + {}^{0}_{-1}e$

$\mathbf{14_{6}C \rightarrow {}^{14}_{7}N + {}^{0}_{-1}e}$

5. $^{1}\mathbf{x}\,n \rightarrow {}^{y}_{1}H + {}^{0}_{-1}e$

$\mathbf{^{1}0\,n \rightarrow {}^{1}_{1}H + {}^{0}_{-1}e}$

6. $^{4}_{2}He \rightarrow 4^{1}_{1}H + \mathbf{x}\,{}^{0}_{-1}e$

$\mathbf{^{4}_{2}He \rightarrow 4^{1}_{1}H + 2\,{}^{0}_{-1}e}$

7. $^{228}_{88}Ra \rightarrow {}^{228}_{90}Th + \mathbf{x}\,\beta$

$\mathbf{^{228}_{88}Ra \rightarrow {}^{228}_{92}Th + 4\,\beta}$

8. $^{232}_{90}Th \rightarrow {}^{212}_{82}Pb + 2\,\beta + \mathbf{x}\,\alpha$

$\mathbf{^{232}_{90}Th \rightarrow {}^{212}_{82}Pb + 2\,\beta + 5\,\alpha}$

9. $^{238}_{92}U \rightarrow {}^{226}_{88}Ra + \mathbf{x}\,\beta + 3\,\alpha$

$\mathbf{^{238}_{92}U \rightarrow {}^{226}_{88}Ra + 2\,\beta + 3\,\alpha}$

10. $^{218}_{84}Po \rightarrow {}^{206}_{82}Pb + \mathbf{x}\,\beta + 3\,\alpha$

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$^{218}_{84}Po$ -> $^{206}_{82}Pb$ + 4β + 3 α

(iii)Gamma (y) particle decay

The gamma (y) particle has the following main characteristic:

i)is **neither** negatively charged(like electrons/beta) nor positively charged(like protons/alpha) therefore **neutral**.

ii)has **no** mass number and atomic number therefore equal to **electromagnetic waves**.

iii) have very **high** penetrating power and thus can be stopped /blocked/shielded by a thick block of lead..

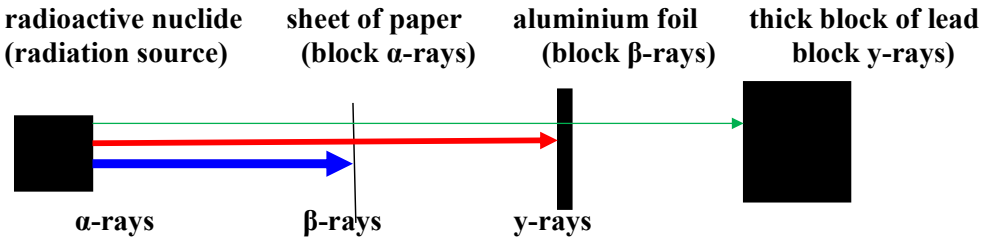
iv) have very **low** ionizing power thus cause less damage to living cells unless on prolonged exposure..

v) a nuclide undergoing y -decay has its mass number and its atomic number **remain** the **same**.

Examples of gamma (y) decay

- $^{37}_{17}Cl$ -> $^{37}_{17}Cl$ + y
- $^{14}_{6}C$ -> $^{14}_{6}C$ + y

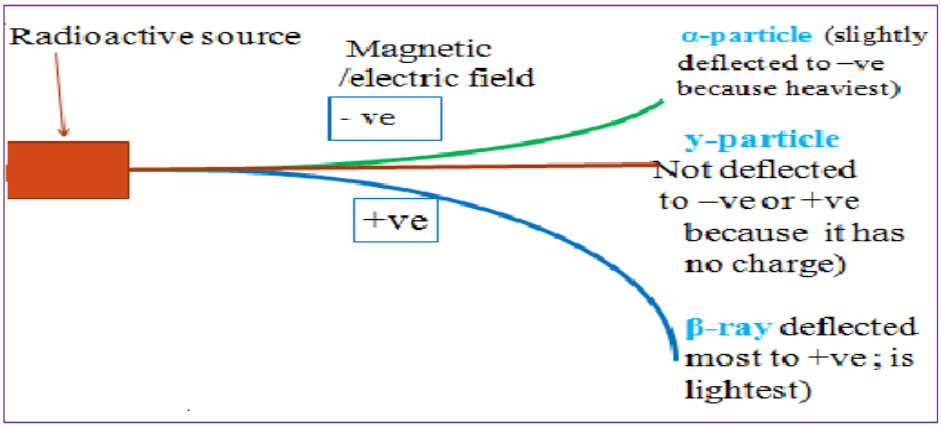
The sketch diagram below shows the **penetrating power** of the radiations from a radioactive nuclide.



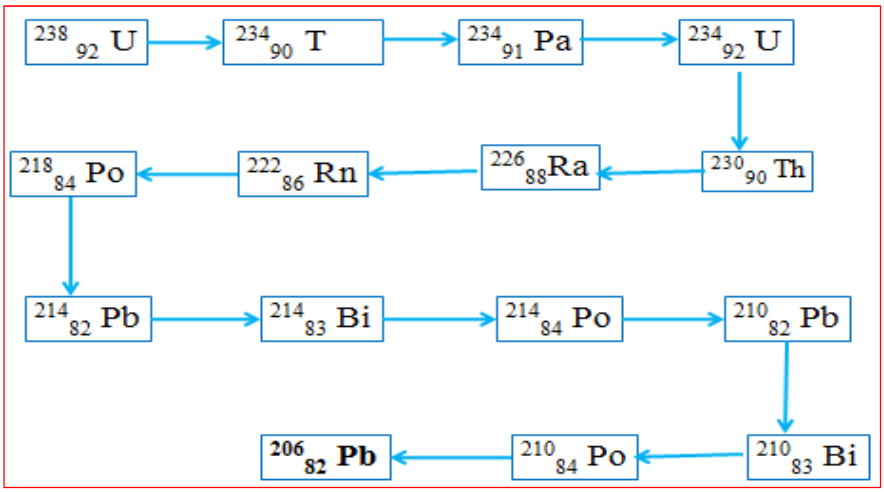
The sketch diagram below illustrates the effect of **electric /magnetic field** on the three radiations from a radioactive nuclide

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Radioactive disintegration/decay **naturally** produces the stable $^{206}_{82}Pb$ nuclide /isotope of lead.Below is the $^{238}_{92}U$ natural decay series. Identify the particle emitted in each case



Write the nuclear equation for the disintegration from :

(i) $^{238}_{92}U$ to $^{234}_{90}T$

$^{238}_{92}\mathbf{U}$ -> $^{234}_{90}\mathbf{T}$ + $^{4}_{2}\mathbf{He}^{2+}$

$^{238}_{92}\mathbf{U}$ -> $^{234}_{90}\mathbf{T}$ + $\boldsymbol{\alpha}$

(ii) $^{238}_{92}U$ to $^{222}_{84}Rn$

$^{238}_{92}\mathbf{U}$ -> $^{222}_{84}\mathbf{Rn}$ + $\mathbf{4}\ ^{4}_{2}\mathbf{He}^{2+}$

$^{238}_{92}\mathbf{U}$ -> $^{222}_{84}\mathbf{Rn}$ + $\mathbf{4\alpha}$

$^{230}_{90}Th$ undergoes alpha decay to $^{222}_{86}Rn$. Find the number of α particles emitted. Write the nuclear equation for the disintegration.

Working

$^{230}_{90}\mathbf{Th}$ -> $^{222}_{86}Rn$ + $\mathbf{x}\ ^{4}_{2}\mathbf{He}$

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Method 1

Using mass numbers

230 = 222 + 4 x => 4 x = 230 - 222 = 8

x = 8 / 4 = 2 α

Using atomic numbers

90 = 86 + 2 x => 2 x = 90 - 86 = 4

x = 4 / 2 = 2 α

Nuclear equation

$^{230}_{90}$ Th -> $^{222}_{86}$ Rn + 2 $^{4}_{2}$ He

$^{214}_{82}$ Pb undergoes beta decay to $^{214}_{84}$ Rn. Find the number of β particles emitted. Write the nuclear equation for the disintegration.

Working

$^{214}_{82}$ Pb -> $^{214}_{84}$ Rn + **x $^{0}_{-1}$ e**

Using atomic numbers only

82 = 84 - x => -x = 82 - 84 = -2

x = 2 β

Nuclear equation

$^{214}_{82}$ Pb -> $^{214}_{84}$ Rn + **2 $^{0}_{-1}$ e**

$^{238}_{92}$ U undergoes beta and alpha decay to $^{206}_{82}$ Pb. Find the number of β and α particles emitted. Write the nuclear equation for the disintegration.

Working

$^{238}_{92}$ U -> $^{206}_{82}$ Pb + **x $^{0}_{-1}$ e +** **y $^{4}_{2}$ He**

Using Mass numbers only

238 = 206 + 4y => 4y = 238 - 206 = 32

y = 32 = 8 α

4

Using atomic numbers only and substituting the 8 α(above)

$^{238}_{92}$ U -> $^{206}_{82}$ Pb + 8 $^{4}_{2}$ He + x $^{0}_{-1}$ e

92 = 82 + 16 + - x

=> 92 – (82 + 16) = - x

x = 6 β

Nuclear equation

$^{238}_{92}$ U -> $^{206}_{82}$ Pb + **6 $^{0}_{-1}$ e +** **8 $^{4}_{2}$ He**

$^{298}_{92}$ U undergoes alpha and beta decay to $^{214}_{83}$ Bi. Find the number of α and β particles emitted. Write the nuclear equation for the disintegration.

Working

$^{298}_{92}$ U -> $^{210}_{83}$ Bi + **x $^{4}_{2}$ He + y $^{0}_{-1}$ e**

Using Mass numbers only

298 = 214 + 4x => 4x = 298 - 214 = 84

y = 84 = 21 α

4

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Using atomic numbers only and substituting the 21 α (above)

$^{238}_{92}U \rightarrow {}^{214}_{83}Bi + 21\,{}^{4}_{2}He + y\,{}^{0}_{-1}e$

$92 = 83 + 42 + -y$

$\Rightarrow 92 - (83 + 42) = -x$

$x = 33\ \beta$

Nuclear equation

$$^{298}_{92}U \rightarrow {}^{210}_{83}Bi + 21\,{}^{4}_{2}He + 33\,{}^{0}_{-1}e$$

B:NUCLEAR FISSION AND NUCLEAR FUSION

Radioactive disintegration/decay can be initiated in an industrial laboratory through two chemical methods:

a) nuclear **fission**

b) nuclear **fusion.**

a)Nuclear fission

Nuclear fission is the process which a fast moving neutron bombards /hits /knocks a heavy **unstable** nuclide releasing l**ighter** nuclide, **three** daughter neutrons and a large quantity of **energy.**

Nuclear fission is the basic chemistry behind **nuclear bombs** made in the nuclear reactors.

The three daughter neutrons becomes again fast moving neutron bombarding / hitting /knocking a heavy unstable nuclide releasing lighter nuclides, three more daughter neutrons each and a larger quantity of energy setting of a **chain reaction**

Examples of nuclear equations showing nuclear fission

$^{1}_{0}n + {}^{235}_{b}U \rightarrow {}^{90}_{38}Sr + {}^{c}_{54}Xe + 3\,{}^{1}_{0}n + a$

$^{1}_{0}n + {}^{27}_{13}Al \rightarrow {}^{28}_{13}Al + y + a$

$^{1}_{0}n + {}^{28}_{a}Al \rightarrow {}^{b}_{11}Na + {}^{4}_{2}He$

$^{a}_{0}n + {}^{14}_{7}N \rightarrow {}^{14}_{b}C + {}^{1}_{1}H$

$^{1}_{0}n + {}^{1}_{1}H \rightarrow {}^{2}_{1}H + a$

$^{1}_{0}n + {}^{235}_{92}U \rightarrow {}^{95}_{42}Mo + {}^{139}_{57}La + 2\,{}^{1}_{0}n + 7\,a$

b) Nuclear fusion

Nuclear fusion is the process which **smaller** nuclides join together to form **larger** / heavier nuclides and releasing a large quantity of **energy**.

Very high temperatures and pressure is required to overcome the repulsion between the atoms.

Nuclear fusion is the basic chemistry behind solar/sun radiation.

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Two daughter atoms/nuclides of Hydrogen fuse/join to form Helium atom/nuclide on the surface of the sun releasing large quantity of energy in form of heat and light.

$^{2}_{1}H + {}^{2}_{1}H \rightarrow {}^{a}_{b}He + {}^{1}_{0}n$

$^{2}_{1}H + a \rightarrow {}^{3}_{2}He$

$^{2}_{1}H + {}^{2}_{1}H \rightarrow a + {}^{1}_{1}H$

$4\ {}^{1}_{1}H \rightarrow {}^{4}_{2}He + a$

$^{14}_{7}H + a \rightarrow {}^{17}_{8}O + {}^{1}_{1}H$

C: HALF LIFE PERIOD ($t^{1}/_{2}$)

The half-life period is the **time** taken for a radioactive nuclide to spontaneously decay/ disintegrate to **half** its **original** mass/ amount.

It is usually denoted **$t\ ^{1}/_{2}$**.

The rate of radioactive nuclide disintegration/decay is **constant** for each nuclide.

The table below shows the half-life period of some elements.

Element/Nuclide	**Half-life period($t\ ^{1}/_{2}$)**
$^{238}_{92}U$	4.5×10^{9} years
$^{14}_{6}C$	5600 years
$^{229}_{88}Ra$	1620 years
$^{35}_{15}P$	14 days
$^{210}_{84}Po$	0.0002 seconds

The **less** the half life the **more unstable** the nuclide /element.

The half-life period is determined by using a Geiger-Muller counter (**GM tube**) .A GM tube is connected to ratemeter that records the **count-rates per unit time**.

This is the rate of decay/ disintegration of the nuclide.

If the count-rates per unit time **fall** by **half**, then the **time** taken for this **fall** is the half-life period.

Examples

a)A radioactive substance gave a count of 240 counts per minute but after 6 hours the count rate were 30 counts per minute. Calculate the half-life period of the substance.

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If $t^{1}/_{2} = x$
then 240 --x-->120 –x-->60 –x--->30
From 240 to 30 =3x =6 hours
=>x $= t^{1}/_{2} = (6 / 3)$
= **2 hours**

b) The count rate of a nuclide fell from 200 counts per second to 12.5 counts per second in 120 minutes.
Calculate the half-life period of the nuclide.

If $t^{1}/_{2} = x$
then
200 --x-->100 –x-->50 –x--->25 –x--->12.5
From 200 to 12.5 =4x =120 minutes
=>x $= t^{1}/_{2} = (120 / 4)$
= **30 minutes**

c) After 6 hours the count rate of a nuclide fell from 240 counts per second to 15 counts per second on the GM tube. Calculate the half-life period of the nuclide.

If $t^{1}/_{2} = x$
then 240 --x-->120 –x-->60 –x--->30 –x--->15
From 240 to 15 =4x =6 hours
=>x $= t^{1}/_{2} = (6 / 4)$= **1.5 hours**

d) Calculate the mass of nitrogen-13 that remain from 2 grams after 6 half-lifes if the half-life period of nitrogen-13 is 10 minutes.

If $t^{1}/_{2} = x$ then:
2 --**x**-->1 –**2x**-->0.5 –**3x**--->0.25 –**4x**-->0.125–**5x**--->0.0625–**6x**--->0.03125
After the 6th half life **0.03125 g** of nitrogen-13 remain.

e) What fraction of a gas remains after 1hour if its half-life period is 20 minutes?

If $t^{1}/_{2} = x$ then:
then 60 /20 = 3x
1 --x--> $^{1}/_{2}$ –2x--> $^{1}/_{4}$ –3x---> $^{1}/_{8}$
After the 3rd half-life $^{1}/_{8}$ of the gas remain

f) 348 grams of a nuclide A was reduced to 43.5 grams after 270days.Determine the half-life period of the nuclide.

If $t^{1}/_{2} = x$ then:
348 --x-->174 –2x-->87 –3x--->43.5
From 348 to 43.5=3x =270days
=>x $= t^{1}/_{2} = (270 / 3)$
= **90 days**

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g) How old is an Egyptian Pharaoh in a tomb with 2grams of ^{14}C if the normal ^{14}C in a present tomb is 16grams.The half-life period of ^{14}C is 5600years.

If $t^{1}/_{2}$ = x = 5600 years then:

16 --x-->8 –2x-->4 –3x--->2

3x = (3 x 5600)

= 16800years

h) 100 grams of a radioactive isotope was reduced 12.5 grams after 81days.Determine the half-life period of the isotope.

If $t^{1}/_{2}$ = x then:

100 --x-->50 –2x-->25 –3x--->12.5

From 100 to 12.5=3x =81days

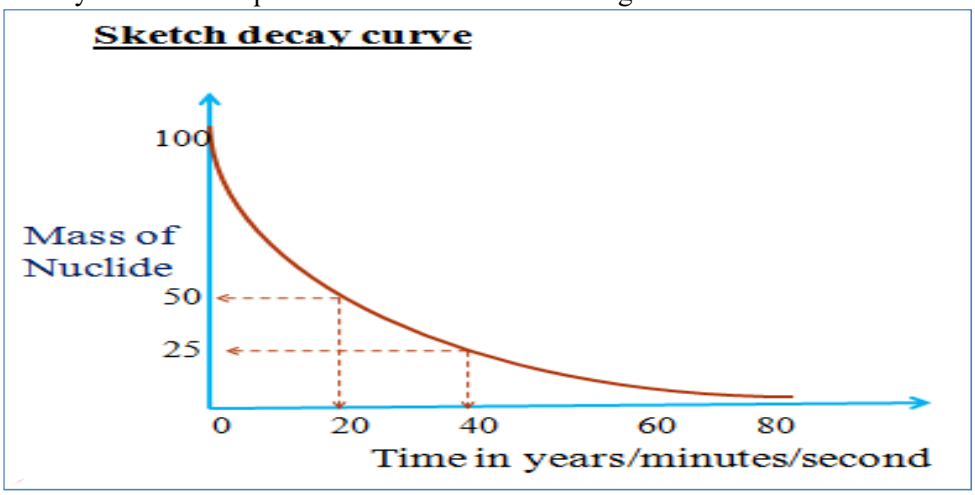
=>x = $t^{1}/_{2}$

= (81 / 3)

= **27 days**

A graph of activity against time is called **decay curve.**

A decay curve can be used to determine the half-life period of an isotope since activity decrease at equal time interval to half the original



(i)From the graph show and determine the half-life period of the isotope.

From the graph $t^{1}/_{2}$ changes in activity from:

(100 – 50) => (20 – 0) = **20** minutes

(50 – 25) => (40 – 20) = **20** minutes

Thus $t^{½}$ = **20 minutes**

(ii)Why does the graph tend to 'O'?

Smaller particle/s will disintegrate /decay to half its original.

There can never be 'O'/zero particles

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D: CHEMICAL vs NUCLEAR REACTIONS

Nuclear and chemical reaction has the following **similarities**:

(i)-both involve the **subatomic** particles; electrons, protons and neutrons in an atom

(ii)-both involve the subatomic particles trying to make the atom more **stable**.

(iii)-Some for of **energy** transfer/release/absorb from/to the environment take place.

Nuclear and chemical reaction has the following **differences**:

(i) Nuclear reactions mainly involve **protons** and **neutrons** in the **nucleus** of an atom.

Chemical reactions mainly involve outer **electrons** in the **energy levels** an atom.

(ii) Nuclear reactions form a **new element**.

Chemical reactions **do not** form new elements

(iii) Nuclear reactions mainly involve evolution/production of **large** quantity of **heat/energy**.

Chemical reactions produce or absorb **small** quantity of heat/energy.

(iv)Nuclear reactions are accompanied by a **loss** in **mass**/mass defect.Do not obey the **law of conservation of matter**.

Chemical reactions are not accompanied by a loss in mass/ mass defect hence obey the law of conservation of matter.

(v)The **rate** of decay/ disintegration of the nuclide is **independent** of **physical** conditions (temperature/pressure /purityp/article size)

The rate of a chemical reaction is **dependent** on physical **conditions** (temperature/pressure/purity/particle size/ surface area)

E: APPLICATION AND USES OF RADIOCTIVITY.

The following are some of the fields that apply and use radioisotopes;

a)**Medicine: -**Treatment of cancer to **kill** malignant tumors through **radiotherapy.**

-**Sterilizing** hospital /surgical instruments /equipments by exposing them to gamma radiation.

b) Agriculture:

If a plant or animal is fed with radioisotope, the metabolic processes of the plant/animal is better understood by **tracing** the route of the radioisotope.

c) Food preservation:

X-rays are used to kill bacteria in **tinned** food to last for a long time.

d) Chemistry:

To study **mechanisms** of a chemical reaction, one reactant is **replaced** in its structure by a radioisotope e.g.

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During esterification the **'O'** joining the ester was discovered comes from the **alkanol** and not alkanoic acid.
During photosynthesis the **'O'** released was discovered comes from **water**.

e) Dating rocks/fossils:

The quantity of ^{14}C in living things (plants/animals) is **constant.**
When they die the fixed mass of ^{14}C is t**rapped** in the cells and **continues** to decay/disintegrate.
The half-life period of ^{14}C is 5600 years .
Comparing the mass of ^{14}C in l**iving** and **dead** cells, the age of the dead can be determined.

F: DANGERS OF RADIOCTIVITY.

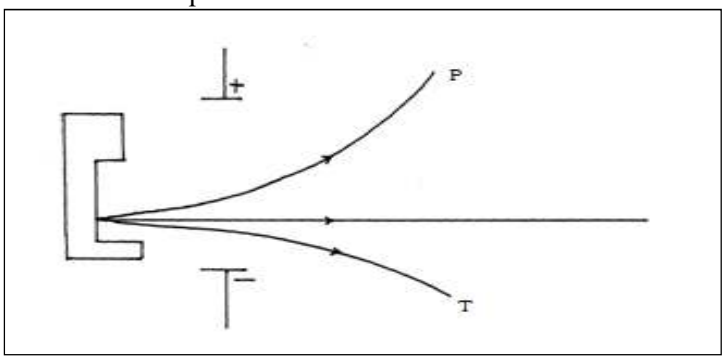
All rays emitted by radioactive isotopes have **ionizing effect** of changing the **genetic** make up of living cells.
Exposure to theses radiations causes **chromosomal** and /or **genetic** mutation in living cells.
Living things should therefore **not** be exposed for a long time to radioactive substances.
One of the main uses of radioactive isotopes is in generation of large cheap **electricity** in nuclear reactors.
Those who work in these reactors must wear protective devises made of **thick** glass or l**ead** sheet.
Accidental leakages of radiations usually occur
In 1986 the Nuclear reactor at **Chernobyl** in Russia had a major explosion that emitted poisonous nuclear material that caused immediate environmental disaster
In 2011, an **earthquake** in Japan caused a nuclear reactor to leak and release poisonous radioactive waste into the Indian Ocean.
The immediate and long term effects of exposure to these **poisonous** radioactive waste on human being is of major concern to all environmentalists.

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G: SAMPLE REVISION QUESTIONS

The figure below shows the behaviour of emissions by a radioactive isotope x. Use it to answer the question follow



(a) Explain why isotope **X** emits radiations. (1mk)

-is unstable //has n/p ratio greater/less than one

(b) Name the radiation labeled **T** (1mk)

alpha particle

(c) Arrange the radiations labeled **P** and **T** in the increasing order of ability to be deflected by an electric filed. (1mk)

T -> P

a) Calculate the mass and atomic numbers of element B formed after $^{212}_{80}X$ has emitted three beta particles, one gamma ray and two alpha particles.

Mass number

= 212 – (0 beta+ o gamma + (2 x 4) alpha = 204

Atomic number

= 80 – (-1 x3) beta + 0 gamma + (2 x 2)) alpha =79

b)Write a balanced nuclear equations for the decay of $^{212}_{80}X$ to B using the information in (a) above.

$^{212}_{80}X$ -> $^{204}_{79}B$ + $2\,^{4}_{2}He$ + 3 $^{0}_{-1}e$ + y

Identify the type of radiation emitted from the following nuclear equations.

(i) $^{14}_{6}C$ -> $^{14}_{7}N$ + ……

β - Beta

(ii) $^{1}_{1}H$ + $^{1}_{0}n$ -> $^{2}_{1}H$ + ……

y -gamma

(iii) $^{235}_{92}U$ -> $^{95}_{42}Mo$ + $^{139}_{57}La$ + $^{1}_{0}n$ +……

7 β – seven beta particles

(iv) $^{238}_{92}U$ -> $^{234}_{90}Th$ + … …

α-alpha

(v) $^{14}_{6}C$ + $^{1}_{1}H$ -> $^{15}_{7}N$ + ……

y-gamma

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X grams of a radioactive isotope takes 100 days to disintegrate to 20 grams. If the half-life period isotope is 25 days, calculate the initial mass X of the radio isotope.

Number of half-lifes = (100 / 25) = 4

20g -----> 40g ----> 80g-----> 160g -----> 320g

Original mass X = 320g

Radium has a half-life of 1620 years.

(i)What is half-life?

The half-life period is the time taken for a radioactive nuclide to spontaneously decay/ disintegrate to half its original mass/ amount

b)If one milligram of radium contains 2.68 x 10 18 atoms ,how many atoms disintegrate during 3240 years.

Number of half-lifes = (3240 / 1620) = 2

1 mg ---1620---> 0.5mg ---1620----> 0.25mg

If 1mg -> 2.68×10^{18} atoms

Then 0.25 mg -> (0.25 x 2.68×10^{18}) = 6.7×10^{17}

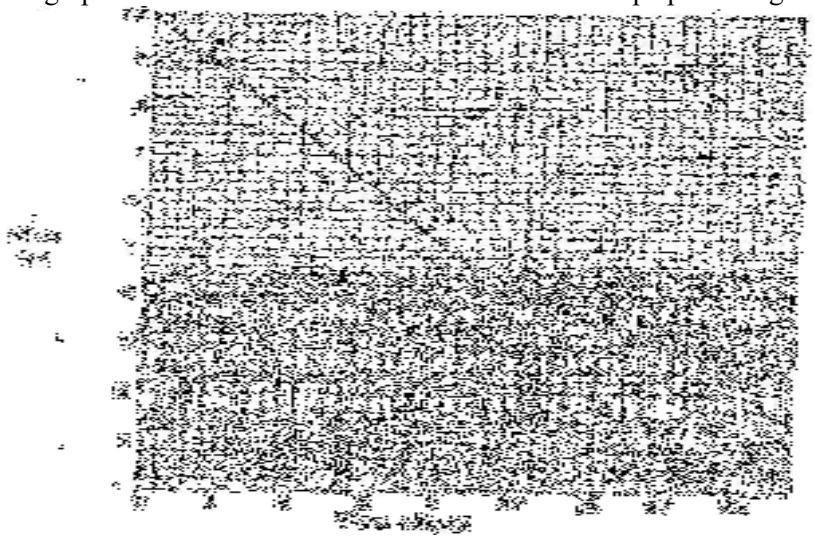
Number of atoms remaining = 6.7×10^{17}

Number of atoms disintegrated =

(2.68×10^{18} - 6.7×10^{17})

= 2.01×10^{18}

The graph below shows the mass of a radioactive isotope plotted against time



Using the graph, determine the half – life of the isotope

From graph 10 g to 5 g takes 8 days

From graph 5 g to 2.5 g takes 16 – 8 = 8 days

Calculate the mass of the isotope dacayed after 32 days

Number of half lifes= 32/8 = 4

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Original mass = 10g

10g—1st -->5g—2nd-->2.5—3rd –>1.25—4th -->0.625 g

Mass remaining = 0.625 g

Mass decayed after 32 days = 10g - 0.625 g = 9.375g

A radioactive isotope X_2 decays by emitting two alpha (a) particles and one beta (β) to form $^{214}_{83}Bi$

(a)Write the nuclear equation for the radioactive decay

$^{212}_{86}X \rightarrow {}^{214}_{83}Bi + 2{}^{4}_{2}He + {}^{0}_{-1}e$

(b)What is the atomic number of X_2?

86

(c) After 112 days, 1/16 of the mass of X_2 remained. Determine the half life of X_2

1—x-> $^1/_2$ –x-> $^1/_4$ –x-> $^1/_8$–x-> $^1/_{16}$

Number of $t^1/_2$ in 112 days = 4

$t^1/_2$ = 112 / 4 = 28 days

1.Study the nuclear reaction given below and answer the questions that follow.

$^{12}_{6}C$ --step 1--> $^{12}_{7}N$ --step 2--> $^{12}_{11}Na$

(a)$^{12}_{6}C$ and $^{14}_{6}C$ are isotopes. What does the term isotope mean?

Atoms of the same element with different mass number /number of neutrons.

(b)Write an equation for the nuclear reaction in step II

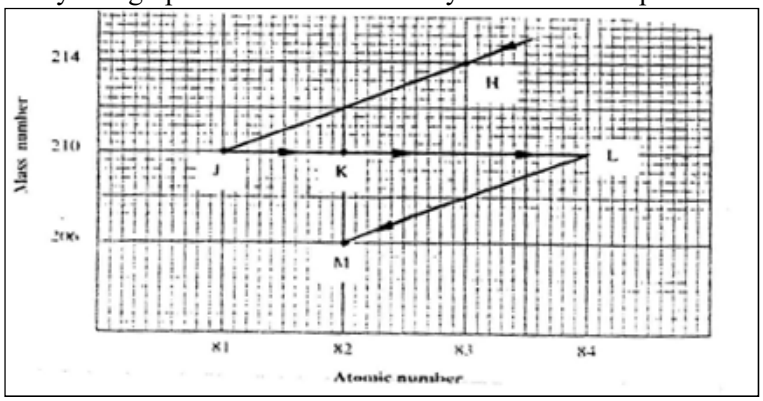
$^{12}_{7}N \rightarrow {}^{12}_{11}Na + {}^{0}_{-1}e$

(c)Give one use of $^{14}_{6}C$

Dating rocks/fossils:

Study of metabolic pathways/mechanisms on plants/animals

Study the graph of a radioactive decay series for isotope H below.



(a)Name the type of radiation emitted when isotope

(i) H changes to isotope J.

Alpha-Mass number decrease by 4 from 214 to 210(y-axis)

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atomic number decrease by 2 from 83 to 81(x-axis)

(ii) J changes to isotope K

Beta-Mass number remains 210(y-axis)

atomic number increase by 1 from 81 to 82(x-axis).

(b) Write an equation for the nuclear reaction that occur when isotope

(i)J changes to isotope L

$^{210}_{81}$**J** -> $^{210}_{84}$**L** + **3** $^{0}_{-1}$**e**

(i)H changes to isotope M

$^{214}_{83}$**H** -> $^{206}_{82}$**M** + **3** $^{0}_{-1}$**e** + **2** $^{4}_{2}$**He**

Identify a pair of isotope of an element in the decay series

K and M

Have same atomic number 82 but different mass number K-210 and M-206

a)A radioactive substance emits three different particles.

Identify the particle:

(i)with the highest mass.

Alpha/ α

(ii) almost equal to an electron

Beta/ β

1.a)State two differences between chemical and nuclear reactions(2mks)

(i) Nuclear reactions mainly involve **protons** and **neutrons** in the **nucleus** of an atom.Chemical reactions mainly involve outer **electrons** in the **energy levels** an atom.

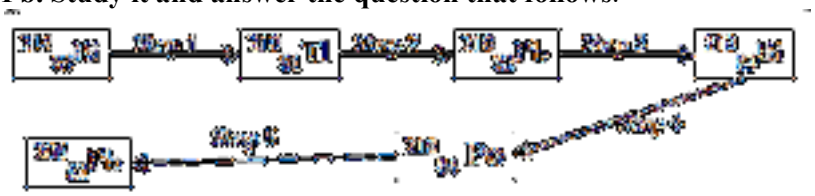
(ii) Nuclear reactions form a **new element**. Chemical reactions **do not** form new elements

(iii) Nuclear reactions mainly involve evolution/production of **large** quantity of **heat/energy.**Chemical reactions produce or absorb **smaller** quantity of heat/energy.

(iv)Nuclear reactions are accompanied by a **loss** in **mass** /mass defect. Chemical reactions are not accompanied by a loss in mass.

(v)R**ate** of decay/ disintegration of nuclide is **independent** of **physical** conditionsThe rate of a chemical reaction is **dependent** on physical **conditions** of temperature/pressure/purity/particle size/ surface area

b)Below is a radioactive decay series starting from $^{214}_{83}$ Bi and ending at $^{206}_{82}$ Pb. Study it and answer the question that follows.



Identify the particles emitted in steps I and III (2mks)

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I - **α-particle**
III - **β-ray**

ii)Write the nuclear equation for the reaction which takes place in (a) step I

$^{214}_{83}Bi \rightarrow {}^{210}_{81}Bi + {}^{4}_{2}He$

(b) step 1 to 3

$^{214}_{83}Bi \rightarrow {}^{210}_{81}Bi + {}^{4}_{2}He + 2\ {}^{0}_{-1}e$

(c) step 3 to 5

$^{210}_{82}Pb \rightarrow {}^{206}_{82}Pb + {}^{4}_{2}He + 2\ {}^{0}_{-1}e$

(c) step 1 to 5

$^{214}_{83}Bi \rightarrow {}^{206}_{82}Pb + 2\,{}^{4}_{2}He + 3\ {}^{0}_{-1}e$

The table below give the percentages of a radioactive isotope of Bismuth that remains after decaying at different times.

Time (min)	0	6	12	22	38	62	100
Percentage of Bismuth	100	81	65	46	29	12	3

i)On the grid below , plot a graph of the percentage of Bismuth remaining(Vertical axis) against time.

ii)Using the graph, determine the:

I. Half – life of the Bismuth isotope

II. Original mass of the Bismuth isotope given that the mass that remained after 70 minutes was 0.16g (2mks)

d) Give one use of radioactive isotopes in medicine (1mk)

14.a)Distinguish between nuclear fission and nuclear fusion. (2mks)

Describe how solid wastes containing radioactive substances should be disposed of. (1mk)

b)(i)Find the values of Z_1 and Z_2 in the nuclear equation below

$$^{Z1}_{92}U + {}^{1}_{0}n \rightarrow {}^{94}_{38}Sr + {}^{140}_{Z2}Xe + 2\ {}^{1}_{0}n$$

iii)What type of nuclear reaction is represented in b (i) above?

A radioactive cobalt $^{61}_{28}Co$ undergoes decay by emitting a beta particle and forming Nickel atom,

Write a balanced decay equation for the above change

1 mark

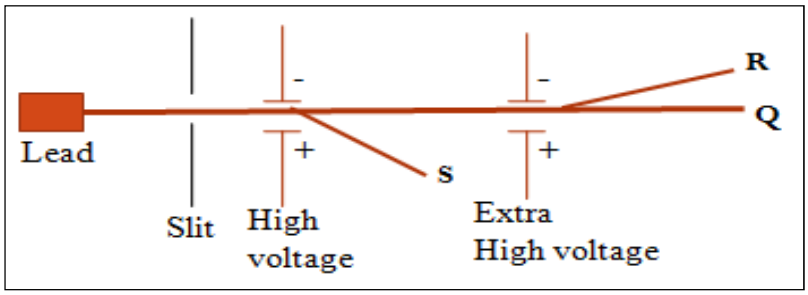
If a sample of the cobalt has an activity of 1000 counts per minute, determine the time it would take for its activity to decrease to 62.50 if the half-life of the element is 30years 2 marks

Define the term half-life.

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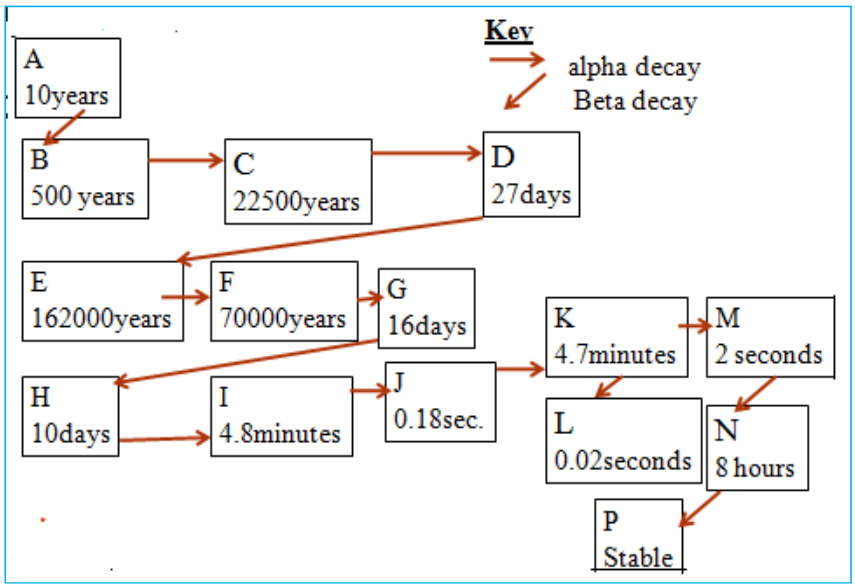
The diagram below shows the rays emitted by a radioactive sample



a) Identify the rays S,R and Q
S- Beta (**β)particle/ray**
R- Alpha (**α)particle/ray**
Q- Gamma (**y)particle/ray**

b) State what would happen if an aluminium plate is placed in the path of ray R,S and Q:
R-is blocked/stopped/do not pass through
Q-is not blocked/pass through
S-is blocked/stopped/do not pass through

(c) The diagram bellow is the radioactive decay series of nuclide A which is $^{241}_{94}Pu$.Use it to answer the questions that follow. The letters are not the actual symbols of the elements.



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(a)Which letter represent the : Explain.

(i)shortest lived nuclide

L-has the shortest half life

(ii)longest lived nuclide

P-Is stable

(iii) nuclide with highest n/p ratio

L-has the shortest half life thus most unstable thus easily/quickly decay/disintegrate

(iv) nuclide with lowest n/p ratio

P-is stable thus do not decay/disintegrate

(b)How long would it take for the following:

(i)Nuclide A to change to B

10 years (half life of A)

(ii) Nuclide D to change to H

27days +162000years+70000years+16days

232000 years and 43 days

(iii) Nuclide A to change to P

27days +162000years+70000years+16days

232000 years and 43 days

Study

A.THE RATE OF CHEMICAL REACTION

(CHEMICAL KINETICS)

1.Introduction

The rate of a chemical reaction is the time taken for a given mass/amount of products to be formed. The rate of a chemical reaction is also the time taken for a given mass/amount of reactant to be consumed /used up.

Some reactions are too slow to be determined. e.g rusting ,decomposition of hydrogen peroxide and weathering.

Some reactions are too fast and instantaneous e.g. neutralization of acid and bases/alkalis in aqueous solution and double decomposition/precipitation.

Other reactions are explosive and very risky to carry out safely e.g. reaction of potassium with water and sodium with dilute acids.

The study of the rate of chemical reaction is useful in knowing the factors that influence the reaction so that efficiency and profitability is maximized in industries.

Theories of rates of reaction.

The rate of a chemical reaction is defined as the rate of change of concentration/amount of reactants in unit time. It is also the rate of formation of given concentration of products in unit time. i.e.

Rate of reaction = Change in concentration/amount of reactants

Time taken for the change to occur

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Rate of reaction = Change in concentration/amount of products formed
Time taken for the products to form

For the above, therefore the rate of a chemical reaction is rate of decreasing reactants to form an increasing product.

The SI unit of **time** is **second**(s) but minutes and hours are also used.

(a)The collision theory

The collision theory is an application of the Kinetic Theory of matter which assumes matter is made up of small/tiny/minute particles like ions atoms and molecules.

The collision theory proposes that

(i)for a reaction to occur, reacting particles must collide.

(ii)not all collisions between reacting particles are successful in a reaction. Collisions that initiate a chemical reaction are called successful / fruitful/ effective collisions

(iii)the speed at which particles collide is called **collision frequency**.

The higher the collision frequency the higher the **chances** of successful / fruitful/ effective collisions to form products.

(iv)the higher the chances of successful collisions, the faster the reaction.

(v)the average distance between solid particles from one another is too big for them to meet and collide successfully.

(vi)dissolving substances in a solvent ,make the solvent a medium for the reaction to take place.

The solute particle distance is reduced as the particle ions are free to move in the solvent medium.

(vii)successful collisions take place if the particles colliding have the required **energy** and right **orientation** which increases their **vibration** and **intensity** of successful / fruitful/ effective collisions to form products.

(b)The Activation Energy(Ea) theory

The **Enthalpy of activation(ΔH_a) /Activation Energy(Ea)** is the minimum amount of energy which the reactants must overcome before they react. Activation Energy(Ea) is usually required /needed in bond breaking of the reacting particles. Bond breaking is an endothermic process that require an energy input.

The higher the bond energy the slower the reaction to **start of**.

Activation energy does not influence whether a reaction is exothermic or endothermic.

The energy level diagrams below shows the activation energy for exothermic and endothermic processes/reactions.

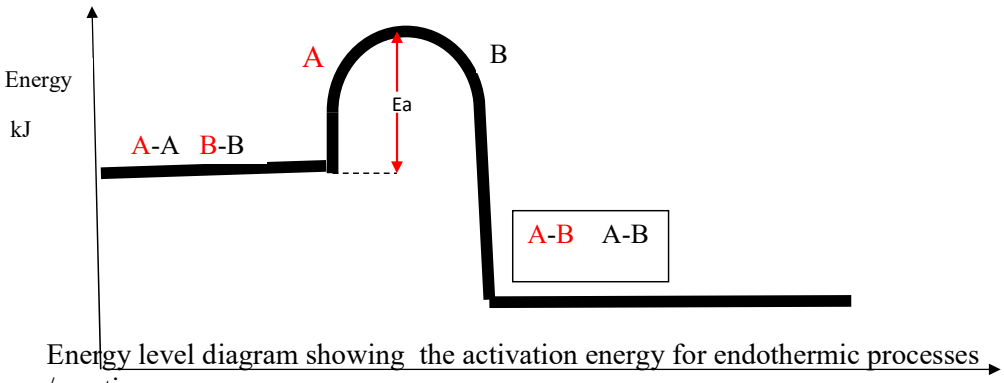
Energy level diagram showing the activation energy for exothermic processes /reactions.

Activated complex

A B

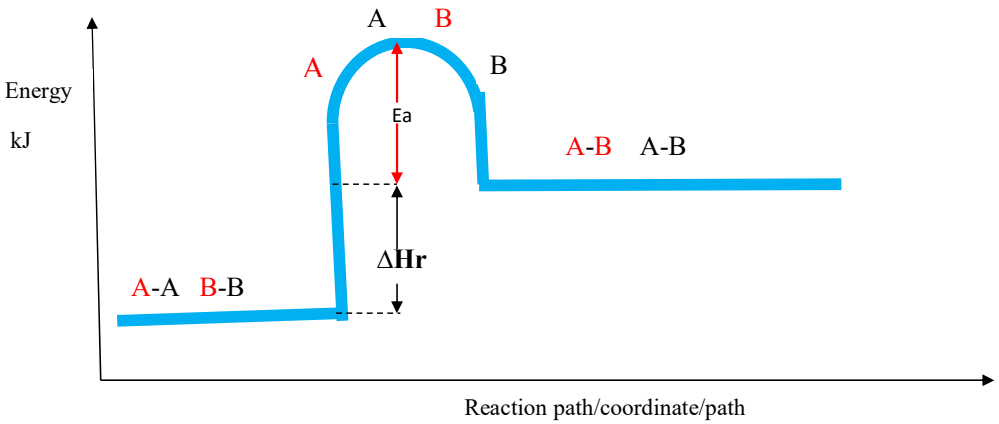
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Energy level diagram showing the activation energy for endothermic processes /reactions.

Activated complex



The activated complex is a mixture of many intermediate possible products which may not exist under normal physical conditions ,but can theoretically exist.
Exothermic reaction proceeds without further heating /external energy because it generates its own energy/heat to overcome activation energy.
Endothermic reaction cannot proceed without further heating /external energy because it does not generates its own energy/heat to overcome activation energy. It generally therefore requires continuous supply of more energy/heat to sustain it to completion.

3. Measuring the rate of a chemical reaction.

The rate of a chemical reaction can be measure as:

(i)Volume of a gas in unit time;

- if reaction is producing a gas as one of the products.
- if reaction is using a gas as one reactants

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(ii)Change in mass of reactants/products for solid products/reactants in unit time.

(iii)formation of a given mass of precipitate in unit time

(iv)a certain mass of reactants to completely form products/diminish.

Reactants may be homogenous or heterogenous.

-Homogenous reactions involve reactants in the **same phase/state** e.g. solid-solid,gas-gas,liquid-liquid.

-Heterogenous reactions involve reactants in the **different phase/state** e.g. solid-liquid,gas-liquid,solid-gas.

4. Factors influencing/altering/affecting/determining rate of reaction

The following factors alter/influence/affect/determine the rate of a chemical reaction:

(a)Concentration
(b)Pressure
(c) Temperature
(d)Surface area
(e)Catalyst

a) Influence of concentration on rate of reaction

The higher the concentration, the higher the rate of a chemical reaction. An increase in concentration of the reactants reduces the distance between the reacting particles increasing their collision frequency to form products.

Practically an increase in concentration **reduces** the time taken for the reaction to take place.

Practical determination of effect of concentration on reaction rate

Method 1(a)

Reaction of sodium thisulphate with dilute hydrochloric acid

Procedure:

Measure 20cm3 of 0.05M sodium thisulphate into a 50cm3 glass beaker. Place the beaker on a white piece of filter paper with ink **mark** 'X' on it. Measure 20cm3 of 0.1M hydrochloric acid solution using a 50cm3 measuring cylinder. Put the acid into the beaker containing sodium thisulphate. Immediately start off the stop watch/clock. Determine the time taken for the ink **mark** 'X' to become invisible /obscured when viewed from above. Repeat the procedure by measuring different volumes of the acid and adding the volumes of the distilled water to complete table 1. Sample results:Table 1.

Volume of acid(cm3)	**Volume of water(cm3)**	**Volume of sodium thiosulphate(cm3)**	**Time taken for mark 'X' to be invisible/obscured(seconds)**	**Reciprocal of time $\frac{1}{t}$**

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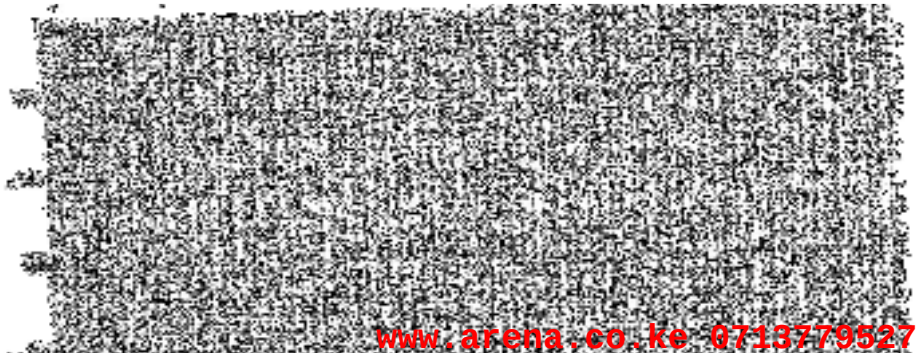
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20.0	0.0	20.0	20.0	5.0×10^{-2}
18.0	2.0	20.0	23.0	4.35×10^{-2}
16.0	4.0	20.0	27.0	3.7×10^{-2}
14.0	6.0	20.0	32.0	3.13×10^{-2}
12.0	8.0	20.0	42.0	2.38×10^{-2}
10.0	10.0	20.0	56.0	1.78×10^{-2}

For most examining bodies/councils/boards the above results score for:
(a) **complete table** as evidence for all the practical work done and completed.
(b) (i)Consistent use of **a decimal point** on time as evidence of understanding/knowledge of the degree of accuracy of stop watches/clock.
(ii)Consistent use of a minimum of **four decimal points** on inverse/reciprocal of time as evidence of understanding/knowledge of the degree of accuracy of scientific calculator.
(c) **accuracy** against a school value based on **candidate's teachers-results** submitted.
(d) correct **trend** (time increase as more water is added/acid is diluted) in conformity with expected **theoretical** results.
Sample questions

1. On separate graph papers plot a graph of:
(i)volume of acid used(x-axis) against time. Label this graph I
(ii) volume of acid used(x-axis) against 1/t. Label this graph II
2. Explain the shape of graph I
Diluting/adding water is causes a decrease in concentration.
Decrease in concentration reduces the rate of reaction by increasing the time taken for reacting particle to collide to form products.

Sketch sample Graph I



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Time

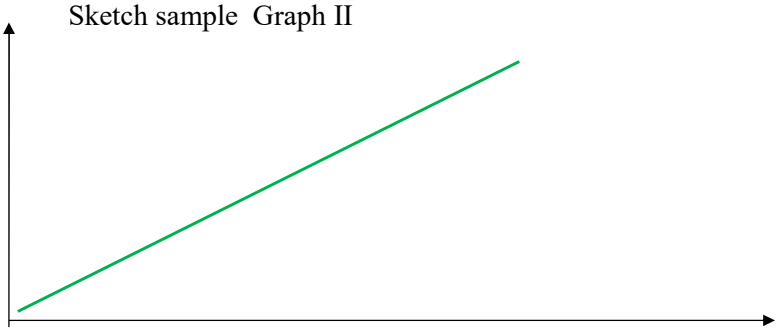
(seconds)

Volume of acid(cm3)

Sketch sample Graph II

1/t

$Sec^{-1} \times 10^{-2}$



3.From graph II ,determine the time taken for the cross to be obscured/invisible when the volume of the acid is:

(i) 13cm3

From a correctly plotted graph

1/t at 13cm3 on the graph =

t = 1 / 2.75 x 10^{-2} = **36.3**

Volume of acid(cm3)

(ii) 15cm3

From a correctly plotted graph

Volume of acid (cm3)

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1/t at 15cm3 on the graph => **3.35×10^{-2}**
t = 1 / 3.35×10^{-2} = **29.8507 seconds**

(iii) 15cm3
From a correctly plotted graph
1/t at 17cm3 on the graph => **4.0×10^{-2}**
t = 1 / 4.0×10^{-2} = **25.0 seconds**

(iv) 19cm3
From a correctly plotted graph
1/t at 19cm3 on the graph => **4.65×10^{-2}**
t = 1 / 4.65×10^{-2} = **21.5054 seconds**

4.From graph II ,determine the volume of the acid used if the time taken for the cross to be obscured/invisible is:

(i)25 seconds
1/t => 1/25 = **4.0×10^{-2}**
Reading from a correctly plotted graph;
4.0×10^{-2} correspond to **17.0 cm3**

(ii)30 seconds
1/t => 1/30 = **3.33×10^{-2}**
Reading from a correctly plotted graph;
3.33×10^{-2} correspond to **14.7 cm3**

(iii)40 seconds
1/t => 1/40 = **2.5×10^{-2}**
Reading from a correctly plotted graph;
2.5×10^{-2} correspond to **12.3 cm3**

4. Write the equation for the reaction taking place
$Na_2S_2O_3$ (aq) + 2HCl(aq) -> 2NaCl (aq)+ SO_2 (g) + S(s) + H_2O(l)
Ionically:
$S_2O_3^{2-}$ (aq) + $2H^+$ (aq) -> SO_2 (g) + S(s) + H_2O(l)

5.Name the yellow precipitate
Colloidal sulphur

Method 1(b)
Reaction of sodium thisulphate with dilute hydrochloric acid
You are provided with
2.0M Hydrochloric acid
0.4M sodium thiosulphate solution
Procedure:
Measure 10cm3 of sodium thisulphate into a 50cm3 glass beaker. Place the beaker on a white piece of filter paper with ink **mark** 'X' on it.
Add 5.0cm3 of hydrochloric acid solution using a 10cm3 measuring cylinder into the beaker containing sodium thisulphate.

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Immediately start off the stop watch/clock. Determine the time taken for the ink **mark** 'X' to become invisible /obscured when viewed from above.
Repeat the procedure by measuring different volumes of the thiosulphate and adding the volumes of the distilled water to complete table 1.

Sample results:Table 1.

Volume of acid(cm3)	**Volume of water (cm3)**	**Volume of sodium thiosulphate (cm3)**	**Concentation of sodium thisulphate in molesdm^{-3}**	**Time(T) taken for mark 'X' to be invisible/ obscured(seconds)**	**T^{-1}**
5.0	0.0	25.0	0.4	20.0	5.0 x 10^{-2}
5.0	5.0	20.0	0.32	23.0	4.35 x 10^{-2}
5.0	10.0	15.0	0.24	27.0	3.7 x 10^{-2}
5.0	15.0	10.0	0.16	32.0	3.13 x 10^{-2}

Note concentration of diluted solution is got:
$C_1V_1=C_2V_2$ => 0.4 x 25 = C_2x 25 =0.4M
$C_1V_1=C_2V_2$ => 0.4 x 20 = C_2x 25 =0.32M
$C_1V_1=C_2V_2$ => 0.4 x 15 = C_2x 25 =0.24M
$C_1V_1=C_2V_2$ => 0.4 x 10 = C_2x 25 =0.16M

Sample questions

1. On separate graph papers plot a graph of:
   (i)Concentration of sodium thiosulphate against time. Label this graph I
   (ii)Concentration of sodium thiosulphate against against **T^{-1}**.Label this graph II
2. Explain the shape of graph I

Diluting/adding water causes a decrease in concentration.
Decrease in concentration reduces the rate of reaction by increasing the time taken for reacting particle to collide to form products.
From graph II
Determine the time taken if
(i)12cm3 of sodium thisulphate is diluted with 13cm3 of water.
At 12cm3 concentration of sodium thisulphate
= $C_1V_1=C_2V_2$ => 0.4 x 1 2 = C_2x 25 =0.192M
From correct graph at concentration 0.192M => 2.4 x10^{-2}
I/t = 2.4 x10^{-2} **t = 41.6667seconds**
(ii)22cm3 of sodium thisulphate is diluted with 3cm3 of water.
At 22cm3 concentration of sodium thisulphate
= $C_1V_1=C_2V_2$ => 0.4 x 22 = C_2x 25 =0.352M

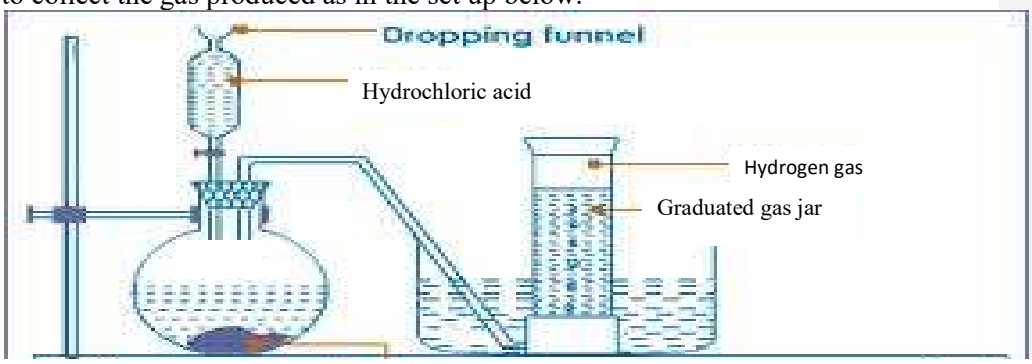
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From correct graph at concentration 0.352M => 3.6 x10^{-2}
I/t = 3.6 x10^{-2} **t = 27.7778seconds**
Determine the volume of water and sodium thiosulphate if T^{-1} is 3.0 x10^{-1}
From correct graph at T^{-1} = 3.0 x10^{-1} => concentration = 0.65 M
= $C_1V_1=C_2V_2$ => 0.4 x 25 = 0.65 M x V_2 = **15.3846cm3**
Volume of water = 25 - 15.3846cm3 = **9.6154cm3**
Determine the concentration of hydrochloric acid if 12cm3 of sodium thiosulphate and 13cm3 of water was used.
At 12cm3 concentration of sodium thisulphate
= $C_1V_1=C_2V_2$ => 0.4 x 1 2 = C_2x 25 =0.192M
Mole ratio $Na_2S_2O_3$:HCl =1:2

Moles of $Na_2S_2O_3$ = $\frac{0.192M \times 12}{1000}$ => 2.304 x 10^{-3} moles
Mole ratio HCl =$\frac{2.304 \times 10^{-1} \text{ moles}}{2}$ = 1.152 x 10^{-3} moles
Molarity o f HCl = $\frac{1.152 \times 10^{-3} \text{ moles} \times 1000}{5.0}$ = **0.2304M**

Method 2
Reaction of Magnesium with dilute hydrochloric acid
Procedure
Scub 10centimeter length of magnesium ribbon with sand paper/steel wool.
Measure 40cm3 of 0.5M dilute hydrochloric acid into a flask .Fill a graduated gas jar with water and invert it into a trough. Stopper the flask and set up the apparatus to collect the gas produced as in the set up below:



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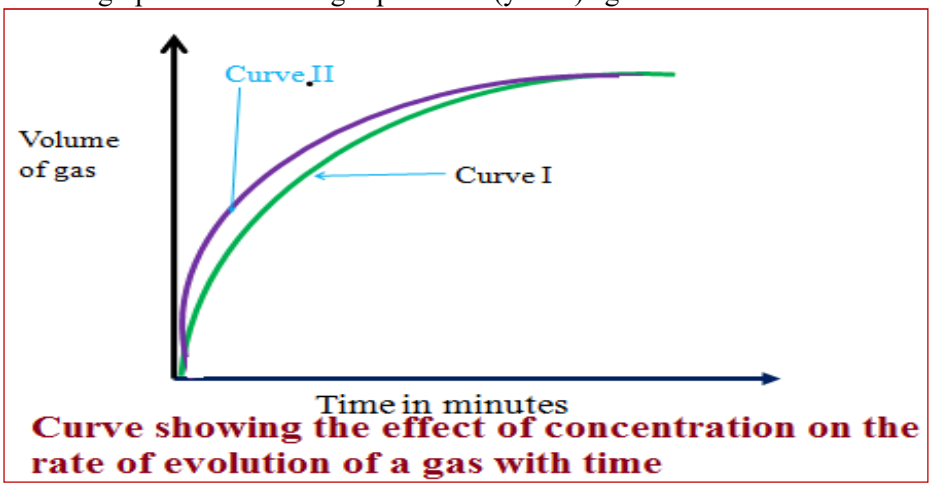
Carefully remove the stopper, carefully put the magnesium ribbon into the flask . cork tightly. Add the acid into the flask. Connect the delivery tube into the gas jar. Immediately start off the stop watch and determine the volume of the gas produced after every 30 seconds to complete table II below.

Sample results: Table II

Time(seconds)	0	30	60	90	120	150	180	210	240
Volume of gas produced(cm3)	**0.0**	**20.0**	**40.0**	**60.0**	**80.0**	**90.0**	**95.0**	**96.0**	**96.0**

Sample practice questions

1.Plot a graph of volume of gas produced (y-axis) against time



Curve showing the effect of concentration on the rate of evolution of a gas with time

2.Explain the shape of the graph.

The rate of reaction is faster when the concentration of the acid is high .

As time goes on, the concentration of the acid decreases and therefore less gas is produced.

When all the acid has reacted, no more gas is produced after 210 seconds and the graph flattens.

3.Calculate the rate of reaction at 120 seconds

From a tangent at 120 seconds rate of reaction = Change in volume of gas / Change in time

=> **From the tangent** at 120seconds $\frac{V_2 - V_1}{T_2 - T_1} = \frac{96-84}{150-90} = \frac{12}{60} =$ **$0.2cm^3sec^{-1}$**

4. Write an ionic equation for the reaction taking place.

$Mg^{2+}(s) + 2H^{+}(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

5. On the same axis sketch then explain the curve that would be obtained if:

(i) 0.1 M hydrochloric acid is used –Label this curve I

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(ii)1.0 M hydrochloric acid is used –Label this curve II

Observation:

Curve I is to the right

Curve II is to the left

Explanation

A decrease in concentration shift the rate of reaction graph to the right as more time is taken for completion of the reaction.

An increase in concentration shift the rate of reaction graph to the left as less time is taken for completion of the reaction.

Both graphs **flatten** after some time indicating the **completion** of the reaction.

b)Influence of pressure on rate of reaction

Pressure affects only gaseous reactants.

An increase in pressure reduces the volume(Boyles law) in which the particles are contained.

Decrease in volume of the container bring the reacting particles closer to each other which increases their chances of effective/successful/fruitful collision to form products.

An increase in pressure therefore increases the rate of reaction by reducing the time for reacting particles of gases to react.

At industrial level, the following are some reactions that are affected by pressure:

(a)Haber process for manufacture of ammonia

$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

(b)Contact process for manufacture of sulphuric(VI)acid

$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

(c)Ostwalds process for the manufacture of nitric(V)acid

$4NH_3(g) + 5O_2(g) \rightarrow 4NO\ (g) + 6H_2O\ (l)$

The influence of pressure on reaction rate is not felt in solids and liquids.

This is because the solid and liquid particles have fixed positions in their strong bonds and therefore no degree of freedom (Kinetic Theory of matter)

c)Influence of temperature on rate of reaction

An increase in temperature increases the kinetic energy of the reacting particles by increasing their collision frequency.

Increase in temperature increases the particles which can overcome the activation energy (Ea).

A **10^oC rise** in temperature doubles the rate of reaction by reducing the time taken for the reaction to complete by **a half.**

Practical determination of effect of Temperature on reaction rate

Method 1

Reaction of sodium thisulphate with dilute hydrochloric acid

Procedure:

Measure 20cm3 of 0.05M sodium thisulphate into a 50cm3 glass beaker.

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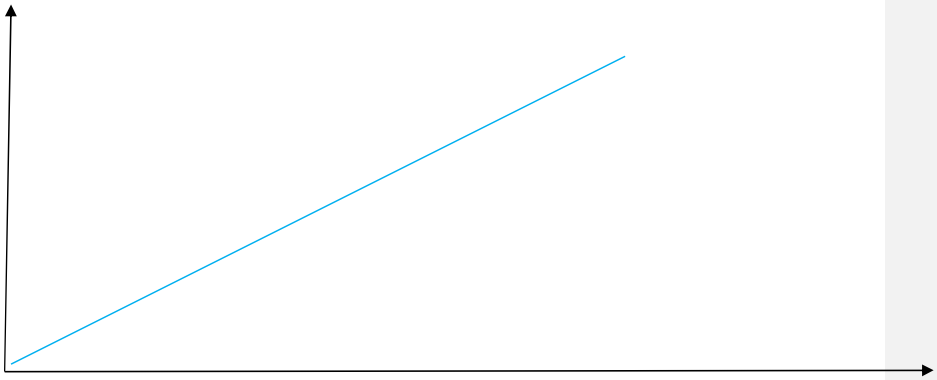
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Place the beaker on a white piece of filter paper with ink **mark** 'X' on it.
Determine and record its temperature as room temperature in table 2 below.
Measure 20cm3 of 0.1M hydrochloric acid solution using a 50cm3 measuring cylinder.
Put the acid into the beaker containing sodium thisulphate.
Immediately start off the stop watch/clock.
Determine the time taken for the ink **mark** 'X' to become invisible /obscured when viewed from above.
Measure another 20cm3 separate portion of the thisulphate into a beaker, heat the solution to 30°C.
Add the acid into the beaker and repeat the procedure above. Complete table 2 below using different temperatures of the thiosulphate.
Sample results:Table 2.

Temperature of $Na_2S_2O_3$	Room temperature	30	40	50	60
Time taken for mark X to be obscured /invisible (seconds)	50.0	40.0	20.0	15.0	10.0
Reciprocal of time($^1/_t$)	0.02	0.025	0.05	0.0667	0.1

Sample practice questions

1. **Plot a graph of temperature(x-axis) against $^1/_t$**



2(a)From your graph determine the temperature at which:

(i)1/t is ;

I. 0.03

Reading directly from a correctly plotted graph = **32.25 °C**

II. 0.07

Reading directly from a correctly plotted graph = **48.0 °C**

(ii) t is;

I. 30 seconds

30 seconds => 1/t =1/30 =**0.033**

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Reading directly from a correctly plotted graph 0.033 => **33.5 ºC**

II. 45 seconds

45 seconds => 1/t =1/45 =**0.022**

Reading directly from a correctly plotted graph 0.022 => **29.0 ºC**

III. 25 seconds

25 seconds => 1/t =1/25 =**0.04**

Reading directly from a correctly plotted graph 0.04 => **36.0 ºC**

(b) From your graph determine the time taken for the cross to become invisible at:

(i) 57.5 ºC

Reading directly from a correctly plotted graph at 57.5 ºC= **0.094**

=>1/t = **0.094**

t= 1/0.094 => **10.6383 seconds**

(ii) 45 ºC

Reading directly from a correctly plotted graph at 45 ºC = **0.062**

=>1/t = **0.062**

t= 1/0.094 => **16.1290 seconds**

(iii) 35 ºC

Reading directly from a correctly plotted graph at 35 ºC = **0.047**

=>1/t = **0.047**

t= 1/0.047 => 2**1.2766 seconds**

Method 2

Reaction of Magnesium with dilute hydrochloric acid

Procedure

Scub 5centimeter length of magnesium ribbon with sand paper/steel wool.

Cut the piece into five equal one centimeter smaller pieces.

Measure 20cm3 of 1.0M dilute hydrochloric acid into a glass beaker .

Put one piece of the magnesium ribbon into the acid, swirl.

Immediately start off the stop watch/clock.

Determine the time taken for the effervescence/fizzing/bubbling to stop when viewed from above.

Record the time in table 2 at room temperature.

Measure another 20cm3 portions of 1.0M dilute hydrochloric acid into a clean beaker.

Heat separately one portion to 30ºC, 40ºC , 50ºC and 60ºC and adding 1cm length of the ribbon and determine the time taken for effervescence /fizzing /bubbling to stop when viewed from above .

Record each time to complete table 2 below using different temperatures of the acid.

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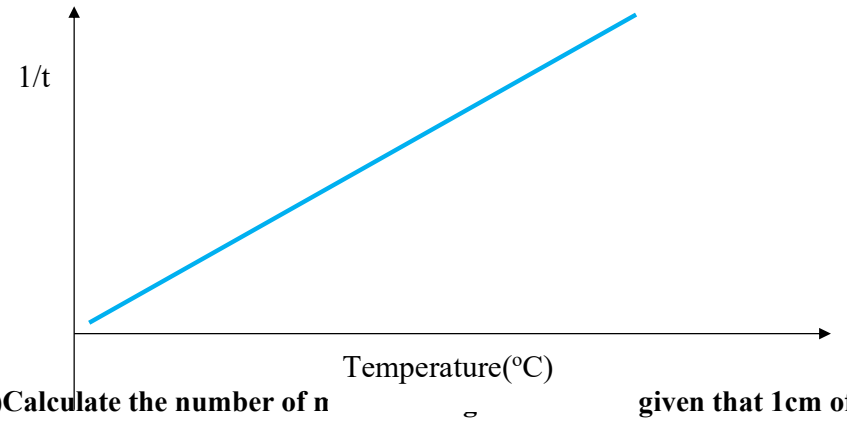
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Sample results:Table 1.

Temperature of acid(°C)	Room temperature	30	40	50	60
Time taken effervescence to stop (seconds)	80.0	50.0	21.0	13.5	10.0
Reciprocal of time($^1/_t$)	0.0125	0.02	0.0476	0.0741	0.1

Sample practice questions

1. **Plot a graph of temperature(x-axis) against $^1/_t$**



2.(a)Calculate the number of n given that 1cm of magnesium has a mass of 1g.(Mg= 24.0)

Moles = Mass of magnesium => 1.0 = **4.167 x 10 $^{-2}$ moles**
Molar mass of Mg 24

(b)Calculate the number of moles of hydrochloric acid used

Moles of acid = molarity x volume of acid

$$\Rightarrow 1.0 \times \frac{20}{1000} = \mathbf{2.0 \times 10^{-2}\ moles}$$

(c)Calculate the mass of magnesium that remain unreacted

Mole ratio Mg: HCl = 1:2
Moles Mg = ½ moles HCl
=> ½ x 2.0 x 10 $^{-2}$ moles = **1.0 x 10 $^{-2}$ moles**
Mass of reacted Mg = moles x molar mass
=> 1.0 x 10 $^{-2}$ moles x 24 = **0.24 g**
Mass of unreacted Mg = Original total mass - Mass of reacted Mg
=> 1.0 g – 0.24 = **0.76 g**

(b)Calculate the total volume of hydrogen gas produced during the above reactions.

Mole ratio Mg : H_2 = 1:1
Moles of Mg that reacted per experiment = moles H_2 =**1.0 x 10 $^{-2}$ moles**

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Volume of Hydrogen at s.t.p produced per experiment = moles x 24 dm3
=> 1.0 x 10^{-2} moles x 24 dm3 = **0.24dm3**
Volume of Hydrogen at s.t.p produced in 5 experiments =0.24 dm3 x 5
= **1.2 dm3**

3.(a)At what temperature was the time taken for magnesium to react equal to:

(i)70seconds

70 seconds => 1/t =1/70 =**0.01429**

Reading directly from a correctly plotted graph **0.01429** => **28.0 ºC**

(ii)40seconds

40 seconds => 1/t =1/40 =**0.025**

Reading directly from a correctly plotted graph **0.025** => **32.0 ºC**

(b)What is the time taken for magnesium to react if the reaction was done at:

(i) 55.0 ºC

Reading directly from a correctly plotted graph at 55.0 ºC=> 1/t = **8.0 x 10^{-2}**
=> t = 1/8.0 x 10^{-2} = **12.5 seconds**

(ii) 47.0 ºC

Reading directly from a correctly plotted graph at 47.0 ºC=> 1/t = **6.0 x 10^{-2}**
=> t = 1/6.0 x 10^{-2} = **16.6667 seconds**

(iii) 33.0 ºC

Reading directly from a correctly plotted graph at 33.0 ºC=> 1/t = **2.7 x 10^{-2}**
=> t = 1/2.7 x 10^{-2} = **37.037 seconds**

4. Explain the shape of the graph.

Increase in temperature increases the rate of reaction as particles gain kinetic energy increasing their frequency and intensity of collision to form products.

d)Influence of surface area on rate of reaction

Surface area is the area of contact. An increase in surface area is a decrease in particle size. Practically an increase in surface area involves chopping /cutting solid lumps into smaller pieces/chips then crushing the chips into powder. Chips thus have a higher surface area than solid lumps but powder has a highest surface area. An increase in surface area of solids increases the area of contact with a liquid solution increasing the chances of successful/effective/fruitful collision to form products. The influence of surface area on rate of reaction is mainly in heterogeneous reactions.

Reaction of chalk/calcium carbonate on dilute hydrochloric acid

Procedure

Measure 20cm3 of 1.0 M hydrochloric acid into three separate conical flasks labeled C_1 C_2 and C_3 .

Using a watch glass weigh three separate 2.5g a piece of white chalk. Place the conical flask C_1 on an electronic balance.

Reset the balance scale to 0.0.

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Put one weighed sample of the chalk into the acid in the conical flask. Determine the scale reading and record it at time =0.0.
Simultaneously start of the stop watch.
Determine and record the scale reading after every 30 seconds to complete Table I .Repeat all the above procedure separately with C_2 and C_3 to complete Table II and Table III by cutting the chalk into small pieces/chips for C_2 and crushing the chalk to powder for C_3
Sample results:Table 1.

Time(seconds)	0.0	30.0	60.0	90.0	120.0	150.0	180.0	210.0	240.0
Mass of $CaCO_3$	**2.5**	**2.0**	**1.8**	**1.4**	**1.2**	**1.0**	**0.8**	**0.5**	**0.5**
Loss in mass	**0.0**	**0.5**	**0.7**	**1.1**	**1.3**	**1.5**	**1.7**	**2.0**	**2.0**

Sample results:Table 1I.

Time(seconds)	0.0	30.0	60.0	90.0	120.0	150.0	180.0	210.0	240.0
Mass of $CaCO_3$	**2.5**	**1.9**	**1.5**	**1.3**	**1.0**	**0.8**	**0.5**	**0.5**	**0.5**
Loss in mass	**0.0**	**0.6**	**1.0**	**1.2**	**1.5**	**1.7**	**2.0**	**2.0**	**2.0**

Sample results:Table III.

Time(seconds)	0.0	30.0	60.0	90.0	120.0	150.0	180.0	210.0	240.0
Mass of $CaCO_3$	**2.5**	**1.8**	**1.4**	**1.0**	**0.8**	**0.5**	**0.5**	**0.5**	**0.5**
Loss in mass	**0.0**	**0.7**	**1.1**	**1.5**	**1.7**	**2.0**	**2.0**	**2.0**	**2.0**

Sample questions:
1.Calculate the loss in mass made at the end of each time from the original to complete table I,II and III
2.On the same axes plot a graph of total loss in mass against time (x-axes) and label them curve I, II, and III from Table I, II, and III.
3.Explain why there is a loss in mass in all experiments.
Calcium carbonate react with the acid to form carbon(IV)oxide gas that escape to the atmosphere.
4.Write an ionic equation for the reaction that take place
$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$
5.Sulphuric(VI)acid cannot be used in the above reaction. On the same axes sketch the curve which would be obtained if the reaction was attempted by reacting a piece of a lump of chalk with 0.5M sulphuric(VI)acid. Label it curve IV. Explain the shape of curve IV.
Calcium carbonate would react with dilute 0.5M sulphuric(VI)acid to form insoluble calcium sulphate(VI) that coat /cover unreacted Calcium carbonate stopping the reaction from reaching completion.

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6.Calculate the volume of carbon(IV)oxide evolved(molar gas volume at room temperature = 24 dm3, C= 12.0, O= 16.O Ca=40.0)

Method I

Mole ratio $CaCO_3(s) : CO_2(g) = 1:1$

Moles $CaCO_3(s)$ used = Mass $CaCO_3(s)$ / Molar mass $CaCO_3(s)$ = **0.025 moles**

Moles $CO_2(g) = 0.025$ moles

Volume of $CO_2(g)$ = moles x molar gas volume

=>0.025 moles x 24 dm3 = **0.600 dm3/600cm3**

Method II

Molar mass of $CaCO_3(s) = 100g$ produce 24 dm3 of $CO_2(g)$

Mass of $CaCO_3(s)$ =2.5 g produce $\frac{2.5 \times 24}{100}$ = **0.600dm3**

7.From curve I ,determine the rate of reaction (loss in mass per second)at time 180 seconds on the curve.

From tangent at 180 seconds on curve I

Rate = $\frac{M_2-M_1}{T_2-T_1}$ => $\frac{2.08 - 1.375}{222-132} = \frac{0.625}{90}$ = **$0.006944g\ sec^{-1}$**

8.What is the effect of particle size on the rate of reaction?

A larger surface area is a reduction in particle size which increases the area of contact between reacting particles increasing their collision frequency.

Theoretical examples

1. Excess marble chips were put in a beaker containing 100cm3 of 0.2M hydrochloric acid. The beaker was then placed on a balance and total loss in mass recorded after every two minutes as in the table below.

Time(minutes)	0.0	2.0	4.0	6.0	8.0	10.0	12.0
Loss in mass(g)	0.0	1.80	2.45	2.95	3.20	3.25	3.25

(a)Why was there a loss in mass?

Carbon (IV) oxide gas was produced that escape to the surrounding

(b)Calculate the average rate of loss in mass between:

(i) 0 to 2 minutes

Average rate = $\frac{M_2-M_1}{T_2-T_1}$ => $\frac{1.80 - 0.0}{2.0 - 0.0} = \frac{1.8}{2}$ = **$9.00g\ min^{-1}$**

(i) 6 to 8 minutes

Average rate = $\frac{M_2-M_1}{T_2-T_1}$ => $\frac{3.20 - 2.95}{8.0 - 6.0} = \frac{0.25}{2}$ = **$0.125g\ min^{-1}$**

(iii) Explain the difference between the average rates of reaction in (i) and(ii) above.

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Between 0 and 2 minutes , the concentration of marble chips and hydrochloric acid is high therefore there is a higher collision frequency between the reacting particles leading to high successful rate of formation of products.

Between 6 and 8 minutes , the concentration of marble chips and hydrochloric acid is low therefore there is low collision frequency between the reacting particles leading to less successful rate of formation of products.

(c)Write the equation for the reaction that takes place.

$CaCO_3(s) + 2HCl\ (aq) \rightarrow CaCO_3\ (aq) + H_2O(l) + CO_2(g)$

(d)State and explain three ways in which the rate of reaction could be increased.

(i) Heating the acid- increasing the temperature of the reacting particles increases their kinetic energy and thus collision frequency.

(ii) Increasing the concentration of the acid-increasing in concentration reduces the distances between the reacting particles increasing their chances of effective/fruitful/successful collision to form products faster.

(iii) Crushing the marble chips to powder-this reduces the particle size/increase surface area increasing the area of contact between reacting particles.

(e)If the solution in the beaker was evaporated to dryness then left overnight in the open, explain what would happen.

It becomes wet because calcium (II) chloride absorbs water from the atmosphere and form solution/is deliquescent.

(f)When sodium sulphate (VI) was added to a portion of the contents in the beaker after the reaction , a white precipitate was formed .

(i)Name the white precipitate.

Calcium(II)sulphate(VI)

(ii)Write an ionic equation for the formation of the white precipitate

$Ca^{2+}(aq) + SO_4^{2-}(aq) \rightarrow CaSO_4(s)$

(iii)State one use of the white precipitate

-Making plaster for building
-Manufacture of plaster of Paris
-Making sulphuric(VI)acid

(g)(i) Plot a graph of total loss in mass(y-axes) against time

(ii) From the graph, determine the rate of reaction at time 2 minutes.

From a tangent/slope at 2 minutes;

Rate of reaction = Average rate $= \frac{M_2 - M_1}{T_2 - T_1} \Rightarrow \frac{2.25 - 1.30}{3.20 - 0.8} = \frac{0.95}{2.4} =$ **$0.3958 g\ min^{-1}$**

(iii) Sketch on the same axes the graph that would be obtained if 0.02M hydrochloric acid was used. Label it curve II

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e) Influence of catalyst on rate of reaction

Catalyst is a substance that alter the rate /speed of a chemical reaction but remain chemically unchanged at the end of a reaction. Biological catalysts are called **enzymes**. A catalyst does not alter the amount of products formed but itself may be altered **physically** e.g. from solid to powder to fine powder. Like biological enzymes, a catalyst only catalyse specific type of reactions

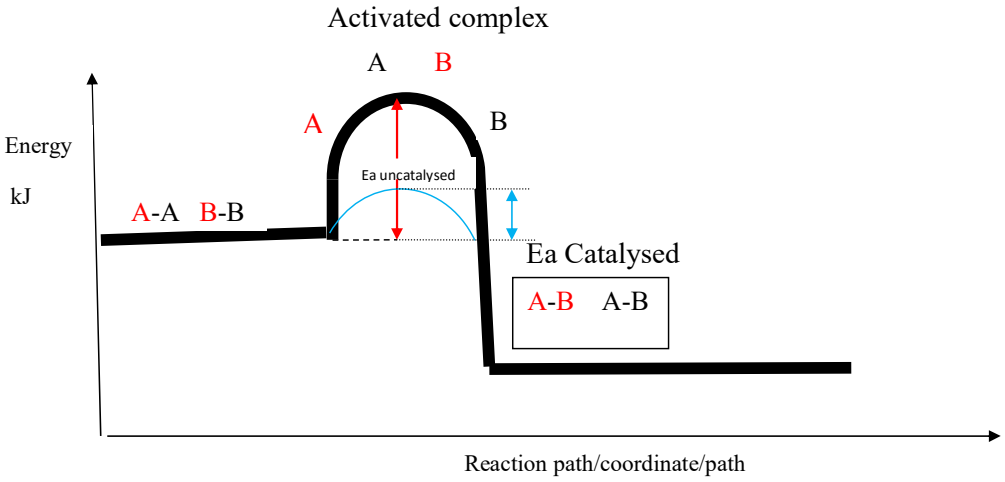
Most industrial catalysts are **transition metals** or their compounds. Catalyst works by lowering the Enthalpy of activation(ΔH_a)/activation energy (Ea) of the reactants .The catalyst lowers the Enthalpy of activation(ΔH_a)/activation energy (Ea) by:

(i) forming short lived intermediate compounds called activated complex that break up to form the final product/s

(ii) being absorbed by the reactants thus providing the surface area on which reaction occurs.

A catalyst has no effect on the enthalpy of reaction ΔH_r but only lowers the Enthalpy of activation(ΔH_a)/activation energy (Ea)It thus do not affect/influence whether the reaction is exothermic or endothermic as shown in the energy level diagrams below.

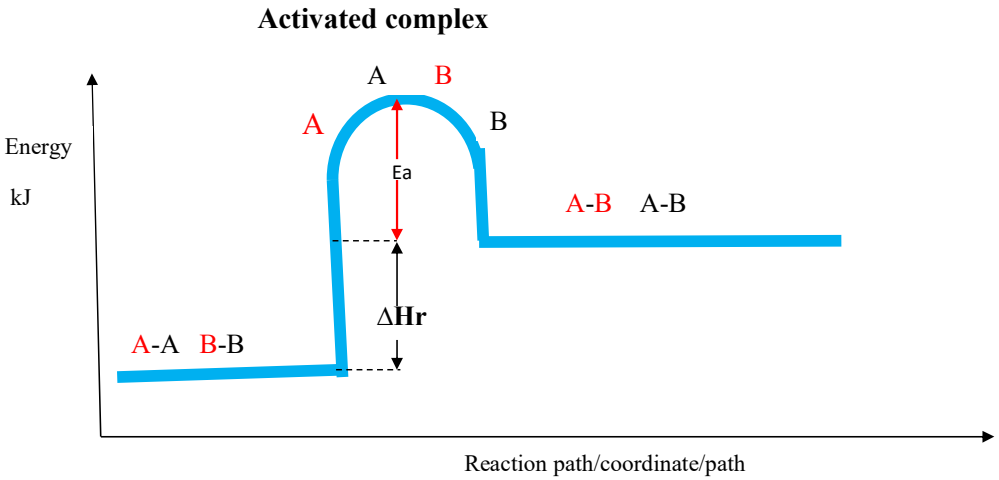
Energy level diagram showing the activation energy for exothermic processes /reactions.



Energy level diagram showing the activation energy for endothermic processes /reactions.

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The following are some catalysed reaction processes.

(a)The contact process

Vanadium(V) Oxide(V_2O_5) or platinum(Pt) catalyses the oxidation of sulphur(IV)oxide during the manufacture of sulphuric(VI) acid from contact process.

$SO_2(g) + O_2(g)$ ----**V_2O_5**--> $SO_3(g)$

To **reduce** industrial cost of manufacture of sulphuric (VI) acid from contact process Vanadium(V) Oxide(V_2O_5) is used because it is **cheaper** though it is **easily poisoned** by impurities.

(b)Ostwalds process

Platinum promoted with Rhodium catalyses the oxidation of ammonia to nitrogen(II)oxide and water during the manufacture of nitric(V)acid

$4NH_3(g) + 5O_2(g)$ ----**Pt/Rh**--> $4NO(g) + 6H_2O(l)$

(c)Haber process

Platinum or iron catalyses the combination of nitrogen and hydrogen to form ammonia gas

$N_2(g) + 3H_2(g)$ ---**Pt or Fe**---> $2NH_3(g)$

(d)Hydrogenation/Hardening of oil to fat

Nickel (Ni) catalyses the hydrogenation of unsaturated compound containing - C=C- or –C=C- to saturated compounds without double or triple bond

This process is used is used in hardening oil to fat.

(e)Decomposition of hydrogen peroxide

Manganese(IV)oxide speeds up the rate of decomposition of hydrogen peroxide to water and oxygen gas.

This process/reaction is used in the school laboratory preparation of Oxygen.

$2H_2O_2(g)$ ----**MnO_2**--> $O_2(g) + 2H_2O(l)$

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(f)Reaction of metals with dilute sulphuric(VI)acid

Copper(II)sulphate(VI) speeds up the rate of production of hydrogen gas from the reaction of Zinc and dilute sulphuric(VI)acid.

This process/reaction is used in the school laboratory preparation of Hydrogen.

$H_2SO_4(aq) + Zn(s) \xrightarrow{CuSO_4} ZnSO_4(aq) + H_2(g)$

(g) Substitution reactions

When placed in bright **sunlight** or **U.V /ultraviolet** light , a mixture of a halogen and an alkane undergo substitution reactions **explosively** to form halogenoalkanes. When paced in **diffused** sunlight the reaction is very **slow**.

e.g. $CH_4(g) + Cl_2(g) \xrightarrow{\text{u.v. light}} CH_3Cl(g) + HCl(g)$

(h)Photosynthesis

Plants convert carbon(IV)oxide gas from the atmosphere and water from the soil to form glucose and oxygen as a byproduct using sunlight / ultraviolet light.

$6CO_2(g) + 6H_2O(l) \xrightarrow{\text{u.v. light}} C_6H_{12}O_6(g) + O_2(g)$

(i)Photography

Photographic film contains silver bromide emulsion which decomposes to silver and bromine on exposure to **sunlight**.

$2AgBr(s) \xrightarrow{\text{u.v/sun light}} 2Ag(s) + Br_2(l)$

When developed, the silver deposits give the picture of the object whose photograph was taken depending on **intensity** of light. A picture photographed in **diffused** light is therefore **blurred**.

Practical determination of effect of catalyst on decomposition of hydrogen peroxide

Measure 5cm3 of 20 volume hydrogen peroxide and then dilute to make 40cm3 in a measuring cylinder by adding distilled water.

Divide it into two equal portions.

(**i**)Transfer one 20cm3volume hydrogen peroxide into a conical/round bottomed/flat bottomed flask. Cork and swirl for 2 minutes. Remove the cork. Test the gas produced using a glowing splint. Clean the conical/round bottomed/flat bottomed flask.

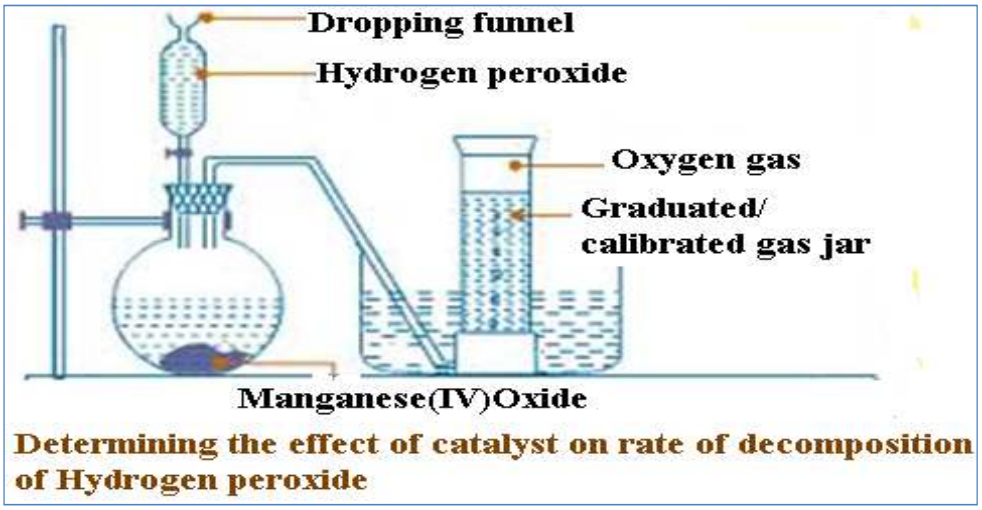
(**ii**)Put 2.0g of Manganese (IV) oxide into the clean conical/round bottomed/flat bottomed flask. Stopper the flask.

Transfer the second portion of the 20cm3volume hydrogen peroxide into a conical/round bottomed/flat bottomed flask through the dropping/thistle funnel. Connect the delivery tube to a calibrated/graduated gas jar as in the set up below. Start off the stop watch and determine the volume of gas in the calibrated/graduated gas jar after every 30 seconds to complete Table 1.

(**iii**)Weigh a filter paper .Use the filter paper to filter the contents of the conical conical/round bottomed/flat bottomed flask. Put the residue on a sand bath to dry. Weigh the dry filter paper again .Determine the new mass Manganese (IV) oxide.

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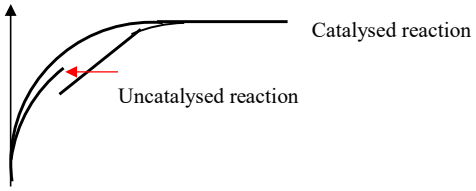
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Time(seconds)	0.0	30.0	60.0	90.0	120.0	150.0	180.0	210.0	240.0	270.0
Volume of gas (cm3)	0.0	20.0	40.0	60.0	80.0	90.0	95.0	96.0	96.0	96.0

Mass of MnO_2 **before** reaction(g)	Mass of MnO_2 **after** reaction(g)
2.0	2.0

Plot a graph of volume of gas produced against time(x-axes)



b) On the same axes, plot a graph of the uncatalysed reaction.

(c) Explain the changes in mass of manganese(IV)oxide before and after the reaction.

The mass of MnO_2 before and after the reaction is the same but a more fine powder after the experiment. A catalyst therefore remains unchanged chemically but may physically change.

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B.EQUILIBRIA (CHEMICAL CYBERNETICS)

Equilibrium is a state of balance.

Chemical equilibrium is state of balance between the reactants and products.

As reactants form products, some products form back the reactants.

Reactions in which the reactants form products to completion are said to be reversible i.e.

A + B -> C + D

Reactions in which the reactants form products and the products can reform the reactants are said to be reversible.

A + B C + D

Reversible reactions may be:

(**a**)Reversible physical changes

(**b**)Reversible chemical changes

(**c**)Dynamic equilibrium

(a)Reversible physical changes

Reversible physical change is one which involves:

(i) change of state/phase from solid, liquid, gas or aqueous solutions. States of matter are interconvertible and a reaction involving a change from one state/phase can be reversed back to the original.

(ii) colour changes. Some substances/compounds change their colours without change in chemical substance.

Examples of reversible physical changes

(i) colour change on heating and cooling:

I. Zinc(II)Oxide changes from white when cool/cold to yellow when hot/heated and back.

$ZnO(s) \rightleftharpoons ZnO(s)$

(white when cold) (yellow when hot)

II. Lead(II)Oxide changes from yellow when cold/cool to brown when hot/heated and back.

$PbO(s) \rightleftharpoons PbO(s)$

(brown when hot) (yellow when cold)

(ii)Sublimation

I. Iodine **sublimes** from a grey crystalline solid on heating to purple vapour. Purple vapour undergoes **deposition** back to the grey crystalline solid.

$I_2(s) \rightleftharpoons I_2(g)$

(grey crystalline solid undergo sublimation) (purple vapour undergo deposition)

II. Carbon (IV)oxide gas undergoes **deposition** from a colourless gas to a white solid at very high pressures in a cylinder. It **sublimes** back to the colourless gas if pressure is reduced

$CO_2(s) \rightleftharpoons CO_2(g)$

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(white powdery solid undergo sublimation) (colourless/odourless gas undergo deposition)

(iii)Melting/ freezing and boiling/condensation

Ice on heating undergo **melting** to form a liquid/water. Liquid/water on further heating **boil**/vaporizes to form gas/water vapour. Gas/water vapour on cooling, **condenses**/liquidifies to water/liquid. On further cooling, liquid water **freezes** to ice/solid.

$H_2O(s)$ ⇌ (Melting / Freezing) $H_2O(l)$ ⇌ (boiling / condensing) $H_2O(s)$

(iv)Dissolving/ crystallization/distillation

Solid crystals of soluble substances (solutes) dissolve in water /solvents to form a uniform mixture of the solute and solvent/solution. On crystallization /distillation /evaporation the solvent evaporate leaving a solute back. e.g.

$NaCl(s) + aq \rightleftharpoons NaCl(aq)$

(b)Reversible chemical changes

These are reactions that involve a chemical change of the reactants which can be reversed back by recombining the new substance formed/products.

Examples of Reversible chemical changes

(i)Heating Hydrated salts/adding water to anhydrous salts.

When **hydrated** salts are heated they **lose** some/all their **water** of crystallization and become **anhydrous**.Heating an unknown substance /compound that forms a **colourless liquid droplets** on the **cooler** parts of a dry test/boiling tube is in fact a **confirmation** inference that the substance/compound being heated is **hydrated**. When anhydrous salts are added (back) some water they form hydrated compound/salts.

Heating Copper(II)sulphate(VI)pentahydrate and cobalt(II)chloride hexahydrate

(i)Heat about 5.0g of Copper(II)sulphate(VI) pentahydrate in a clean dry test tube until there is no further colour change on a small Bunsen flame. Observe any changes on the side of the test/boiling tube. Allow the boiling tube to cool.Add about 10 drops of distilled water. Observe any changes.

(ii)Dip a filter paper in a solution of cobalt(II)chloride hexahydrate. Pass one end the filter paper to a small Bunsen flame repeatedly. Observe any changes on the filter paper. Dip the paper in a beaker containing distilled water. Observe any changes.

Sample observations

Hydrated compound	**Observation before heating**	**Observation after heating**	**Observation on adding water**
Copper(II)sulphate (VI) pentahydrate	**Blue** crystalline solid	(i)colour changes from **blue** to **white**. (ii)colourless liquid forms on the cooler parts of	(i)colour changes from **white** to **blue** (ii)boiling tube becomes warm

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		boiling / test tube	/hot.
Cobalt(II)chloride hexahydrate	**Pink** crystalline solid/solution	(i)colour changes from **pink** to **blue**. (ii) colourless liquid forms on the cooler parts of boiling / test tube (if crystal are used)	(i)colour changes from **blue** to **pink** (ii)boiling tube becomes warm/hot.

When blue Copper(II)sulphate (VI) pentahydrate is heated, it loses the five molecules of water of crystallization to form white anhydrous Copper(II)sulphate (VI).Water of crystallization form and condenses as colourless droplets on the cooler parts of a dry boiling/test tube.

This is a chemical change that produces a new substance. On adding drops of water to an anhydrous white copper(II)sulphate(VI) the hydrated compound is formed back. The change from **hydrated** to **anhydrous** and **back** is therefore **reversible chemical change.**Both anhydrous white copper(II)sulphate(VI) and blue cobalt(II)chloride hexahydrate are therefore used to test for the presence of water when they turn to blue and pink respectively.

$CuSO_4$(**s**) + $5H_2O$(**l**) $\rightleftharpoons$ $CuSO_4 \cdot 5H_2O$(**s/aq**)
(white/anhydrous) (blue/hydrated)

$CoCl_2$(**s**) + $6H_2O$(**l**) $\rightleftharpoons$ $CoCl_2 \cdot 6H_2O$(**s/aq**)
(blue/anhydrous) (pink/hydrated)

(ii)Chemical sublimation

Some compounds sublime from solid to gas by dissociating into new different compounds. e.g.

Heating ammonium chloride

(**i**)Dip a glass rod containing concentrated hydrochloric acid. Bring it near the mouth of a bottle containing concentrated ammonia solution. Explain the observations made.

When a glass rod containing hydrogen chloride gas is placed near ammonia gas, they react to form ammonium chloride solid that appear as **white fumes**.

This experiment is used interchangeably to test for the presence of hydrogen chloride gas (and hence Cl^- ions) and ammonia gas (and hence NH_4^+ ions)

(**ii**)Put 2.0 g of ammonium chloride in a long dry boiling tube. Place wet / moist /damp blue and red litmus papers separately on the sides of the mouth of the boiling tube. Heat the boiling tube gently then strongly. Explain the observations made.

When ammonium chloride is heated it dissociates into ammonia and hydrogen chloride gases. Since ammonia is less dense, it diffuses faster to turn both litmus papers blue before hydrogen chloride turn red because it is denser. The heating and cooling of ammonium chloride is therefore a **reversible chemical change**.

$NH_3(g)$ + $HCl(g)$ $\rightleftharpoons$ $NH_4Cl(s)$

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(Turns moist litmus paper blue) (Turns moist litmus paper red) (forms white fumes)

(c)Dynamic equilibria

For reversible reactions in a closed system:

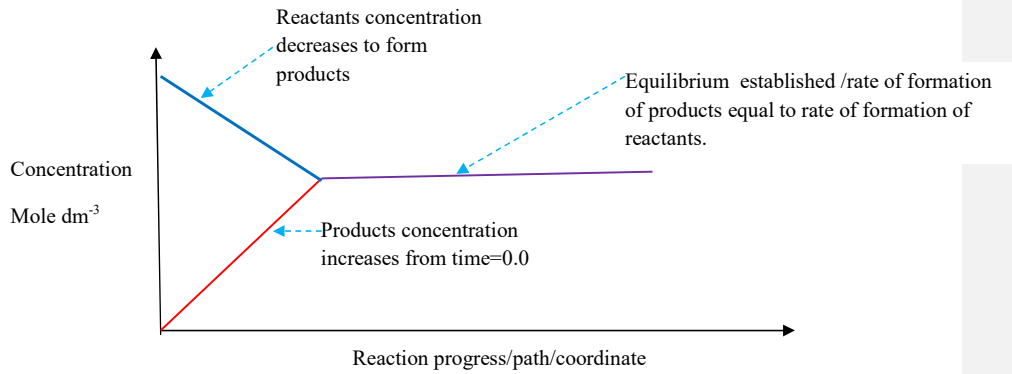
(i) at the beginning;

-the reactants are decreasing in concentration with time

-the products are increasing in concentration with time

(ii) after some time a point is reached when as the reactants are forming products the products are forming reactants. This is called equilibrium.

Sketch showing the changes in concentration of reactants and products in a closed system



For a system in equilibrium:

(i) a reaction from left to right (reactants to products) is called forward reaction.

(ii) a reaction from right to left (products to reactants) is called backward reaction.

(iii)a reaction in which the rate of forward reaction is equal to the rate of backward reaction is called a **dynamic equilibrium.**

A dynamic equilibrium is therefore a balance of the rate of formation of products and reactants. This balance continues until the reactants or products are disturbed/changed/ altered.

The influence of different factors on a dynamic equilibrium was first investigated from 1850-1936 by the French Chemist Louis Henry Le Chatellier. His findings were called Le Chatelliers Principle which states that:

"**if a stress/change is applied to a system in dynamic equilibrium, the system readjust/shift/move/behave so as to remove/ reduce/ counteract/ oppose the stress/change**"

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Le Chatelliers Principle is applied in determining the effect/influence of several factors on systems in dynamic equilibrium. The following are the main factors that influence /alter/ affect systems in dynamic equilibrium:

(a)Concentration
(b)Pressure
(c)Temperature
(d)Catalyst

(a)Influence of concentration on dynamic equilibrium

An increase/decrease in concentration of reactants/products at equilibrium is a stress. From Le Chatelliers principle the system redjust so as to remove/add the excessreduced concentration.

Examples of influence of concentration on dynamic equilibrium

(i)Chromate(VI)/CrO_4^{2-} ions in solution are **yellow**. Dichromate(VI)/$Cr_2O_7^{2-}$ ions in solution are **orange**. The two solutions exist in equilibrium as in the equation:

$2H^+(aq) + 2CrO_4^{2-}(aq) \rightleftharpoons Cr_2O_7^{2-}(aq) + H_2O(l)$

(Yellow) (Orange)

I. If an acid is/H^+ (aq) is added to the equilibrium mixture a stress is created on the reactant side where there is already H^+ ions. The equilibrium shift forward to the right to remove/reduce the **excess** H^+ ions added. Solution mixture becomes More $Cr_2O_7^{2-}$ ions formed in the solution mixture make it to be more **orange** in colour.

II. If a base/OH^- (aq) is added to the equilibrium mixture a stress is created on the reactant side on the H^+ ions. H^+ ions react with OH^- (aq) to form water.

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

The equilibrium shift backward to the left to add/replace the H^+ ions that have reacted with the OH^- (aq) ions . More of the CrO_4^{2-} ions formed in the solution mixture makes it to be more **yellow** in colour.

$2OH^-(aq) + 2Cr_2O_7^{2-}(aq) \rightleftharpoons CrO_4^{2-}(aq) + H_2O(l)$

(Orange) (Yellow)

I. If an acid/ H^+ (aq) is added to the equilibrium mixture a stress is created on the reactant side on the OH^- (aq). H^+ ions react with OH^- (aq) to form water.

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

The equilibrium shift backward to the left to add/replace the $2OH^-$ (aq) that have reacted with the H^+ (aq) ions . More $Cr_2O_7^{2-}$ (aq)ions formed in the solution mixture makes it to be more **Orange** in colour.

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II. If a base /OH^- (aq) is added to the equilibrium mixture a stress is created on the reactant side where there is already OH^- (aq) ions. The equilibrium shift forward to the right to remove/reduce the **excess** OH^- (aq) ions added. More of the $Cr_2O_7^{2-}$ ions are formed in the solution mixture making it to be more **orange** in colour.

(i)Practical determination of the influence of alkali/acid on $Cr_2O_7^{2-}$ / CrO_4^{2-} equilibrium mixture

Measure about 2 cm3 of Potassium dichromate (VI) solution into a test tube.

Note that the solution mixture is orange.

Add three drops of 2M sulphuric(VI) acid. Shake the mixture carefully.

Note that the solution mixture is remains orange.

Add about six drops of 2M sodium hydroxide solution. Shake carefully.

Note that the solution mixture is turns yellow.

Explanation

The above observations can be explained from the fact that both the dichromate(VI)and chromate(VI) exist in equilibrium. Dichromate(VI) ions are stable in acidic solutions while chromate(VI)ions are stable in basic solutions. An equilibrium exist thus:

$$Cr_2O_7^{2-} \underset{H+}{\overset{OH-}{\rightleftharpoons}} CrO_4^{2-}$$

When an **acid** is added, the equilibrium shift **forward** to the right and the mixture become more **orange** as more $Cr_2O_7^{2-}$ ions exist.

When a **base** is added, the equilibrium shift **backward** to the left and the mixture become more **yellow** as more CrO_4^{2-} ions exist.

(ii)Practical determination of the influence of alkali/acid on bromine water in an equilibrium mixture

Measure 2cm3 of bromine water into a boiling tube. Note its colour.

Bromine water is yellow

Add three drops of 2M sulphuric(VI)acid. Note any colour change

Colour becomes more yellow

Add seven drops of 2M sodium hydroxide solution. Note any colour change.

Solution mixture becomes colourless/Bromine water is decolourized.

Explanation

When added distilled water,an equilibrium exist between bromine liquid (Br_2(aq)) and the bromide ion(Br^-), hydrobromite ion(OBr^-) and hydrogen ion(H^+) as in the equation:

$$H_2O(l) + Br_2(aq) \rightleftharpoons OBr^-(aq) + H^+(aq) + Br^-(aq)$$

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If an acid (H^+)ions is added to the equilibrium mixture, it increases the concentration of the ions on the product side which shift backwards to the left to remove the excess H^+ ions on the product side making the colour of the solution mixture more yellow.

If a base/alkali OH^- is added to the equilibrium mixture, it reacts with H^+ ions on the product side to form water.

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

This decreases the concentration of the H^+ ions on the product side which shift the equilibrium forward to the right to replace H^+ ions making the solution mixture colourless/less yellow (Bromine water is decolorized)

(iii)Practical determination of the influence of alkali/acid on common acid-base indicators.

Place 2cm3 of phenolphthalein ,methyl orange and litmus solutions each in three separate test tubes.

To each test tube add two drops of water. Record your observations in Table 1 below.

To the same test tubes, add three drops of 2M sulphuric(VI)acid. Record your observations in Table 1 below.

To the same test tubes, add seven drops of 2M sodium hydroxide solution. Record your observations in Table 1 below.

To the same test tubes, repeat adding four drops of 2M sulphuric(VI)acid. **Table 1**Indicator	Colour of indicator in		
	Water	Acid(2M sulphuric (VI) acid)	Base(2M sodium hydroxide)
Phenolphthalein	**Colourless**	**Colourless**	**Pink**
Methyl orange	**Yellow**	**Red**	**Orange**
Litmus solution	**Colourless**	**Red**	**Blue**

Explanation

An indicator is a substance which shows whether **another** substance is an **acid** , **base** or **neutral**.

Most indicators can be regarded as very weak acids that are partially dissociated into ions.An equilibrium exist between the undissociated molecules and the dissociated anions. Both the molecules and anions are **coloured**. i.e.

$HIn(aq) \rightleftharpoons H^+(aq) + In^-(aq)$

(undissociated indicator molecule(**coloured**)) (dissociated indicator molecule(**coloured**))

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When an acid H^+ is added to an indicator, the H^+ ions increase and equilibrium shift backward to remove excess H^+ ions and therefore the **colour** of the undissociated (**HIn**) molecule **shows/appears.**

When a base/alkali OH^- is added to the indicator, the OH^- reacts with H^+ ions from the dissociated indicator to form water.

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

(from indicator) (from alkali/base)

The equilibrium shift forward to the right to replace the H^+ ion and therefore the colour of dissociated (**In^-**) molecule **shows/appears.**

The following examples illustrate the above.

(i)Phenolphthalein indicator exist as:

$HPh \rightleftharpoons H^+(aq) + Ph^-(aq)$

(colourless molecule) (Pink anion)

On adding an acid ,equilibrium shift backward to the left to remove excess H^+ ions and the solution mixture is therefore **colourless.**

When a base/alkali OH^- is added to the indicator, the OH^- reacts with H^+ ions from the dissociated indicator to form water.

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

(from indicator) (from alkali/base)

The equilibrium shift forward to the right to replace the removed/reduced H^+ ions. The **pink** colour of dissociated (**Ph^-**) molecule **shows/appears.**

(ii)Methyl Orange indicator exists as:

$HMe \rightleftharpoons H^+(aq) + Me^-(aq)$

(Red molecule) (Yellow/Orange anion)

On adding an acid ,equilibrium shift backward to the left to remove excess H^+ ions and the solution mixture is therefore **red.**

When a base/alkali OH^- is added to the indicator, the OH^- reacts with H^+ ions from the dissociated indicator to form water.

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

(from indicator) (from alkali/base)

The equilibrium shift forward to the right to replace the removed/reduced H^+ ions. The **Orange** colour of dissociated (**Me^-**) molecule **shows/appears.**

(b)Influence of Pressure on dynamic equilibrium

Pressure affects gaseous reactants/products. Increase in **pressure** shift/**favours** the equilibrium towards the side with **less volume/molecules**. Decrease in pressure shift the equilibrium towards the side with more volume/molecules. More yield of products is obtained if high pressures produce less molecules / volume of products are formed.

If the products and reactants have **equal** volume/molecules then pressure has **no effect** on the position of equilibrium

The following examples show the influence of pressure on dynamic equilibrium:

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(i)Nitrogen(IV)oxide /Dinitrogen tetroxide mixture

Nitrogen(IV)oxide and dinitrogen tetraoxide can exist in dynamic equilibrium in a closed test tube. Nitrogen(IV)oxide is a brown gas. Dinitrogen tetraoxide is a yellow gas.

Chemical equation : $2NO_2(g) \rightleftharpoons N_2O_4(g)$

Gay Lussacs law	**2**Volume	**1**Volume
Avogadros law	**2**molecule	**1**molecule

2 volumes/molecules of Nitrogen(IV)oxide form **1** volumes/molecules of dinitrogen tetraoxide

Increase in pressure shift the equilibrium forward to the left where there is less volume/molecules.The equilibrium mixture become more **yellow**.

Decrease in pressure shift the equilibrium backward to the right where there is more volume/molecules. The equilibrium mixture become more **brown.**

(ii)Iodine vapour-Hydrogen gas/Hydrogen Iodide mixture.

Pure hydrogen gas reacts with Iodine vapour to form Hydrogen Iodide gas.

Chemical equation : $I_2(g) + H_2(g) \rightleftharpoons 2HI(g)$

Gay Lussacs law	**1**Volume	**1**Volume	**2**Volume
Avogadros law	**1**molecule	**1**molecule	**2**molecule

(1+1) **2** volumes/molecules of Iodine and Hydrogen gasform **2** volumes/molecules of Hydrogen Iodide gas.

Change in pressure thus has no effect on position of equilibrium.

(iii)Haber process.

Increase in pressure of the Nitrogen/Hydrogen mixture favours the formation of **more** molecules of Ammonia gas in Haber process**.**

The yield of ammonia is thus favoured by high pressures

Chemical equation : $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

Gay Lussacs law	**1**Volume	**3**Volume	**2**Volume
Avogadros law	**1**molecule	**3**molecule	**2**molecule

(1 + 3) **4** volumes/molecules of Nitrogen and Hydrogen react to form **2** volumes/molecules of ammonia.

Increase in pressure shift the equilibrium forward to the left where there is less volume/molecules.

The yield of ammonia increase.

Decrease in pressure shift the equilibrium backward to the right where there is more volume/molecules.

The yield of ammonia decrease.

(iv)Contact process.

Increase in pressure of the Sulphur(IV)oxide/Oxygen mixture favours the formation of **more** molecules of Sulphur(VI)oxide gas in Contact process**.** The yield of Sulphur(VI)oxide gas is thus favoured by high pressures.

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Chemical equation : $2SO_2(g)$ + $O_2(g)$ -> $2SO_3(g)$

Gay Lussacs law	**2**Volume	**1**Volume	**2**Volume
Avogadros law	**2**molecule	**1**molecule	**2**molecule

(2 + 1) **3** volumes/molecules of Sulphur(IV)oxide/Oxygen mixture react to form **2** volumes/molecules of Sulphur(VI)oxide gas.

Increase in pressure shift the equilibrium forward to the left where there is less volume/molecules. The yield of Sulphur(VI)oxide gas increase.

Decrease in pressure shift the equilibrium backward to the right where there is more volume/molecules. The yield of Sulphur(VI)oxide gas decrease.

(v)Ostwalds process.

Increase in pressure of the Ammonia/Oxygen mixture favours the formation of **more** molecules of Nitrogen(II)oxide gas and water **vapour** in Ostwalds process**.** The yield of Nitrogen(II)oxide gas and water **vapour** is thus favoured by low pressures.

Chemical equation : $4NH_3(g)$ + $5O_2(g)$ -> $4NO(g)$ + $6H_2O(g)$

Gay Lussacs law	**4**Volume	**5**Volume	**4**Volume	**6**Volume
Avogadros law	**4**molecule	**5**molecule	**4**molecule	**6**Molecule

(4 + 5) **9** volumes/molecules of Ammonia/Oxygen mixture react to form **10** volumes/molecules of Nitrogen(II)oxide gas and water **vapour**.

Increase in pressure shift the equilibrium backward to the left where there is less volume/molecules. The yield of Nitrogen(II)oxide gas and water vapour **decrease**.

Decrease in pressure shift the equilibrium forward to the right where there is more volume/molecules. The yield of Nitrogen(II)oxide gas and water vapour **increase**.

Note

If the water vapour is **condensed** on cooling, then:

Chemical equation : $4NH_3(g)$ + $5O_2(g)$ -> $4NO(g)$ + $6H_2O(\mathbf{l})$

Gay Lussacs law	**4**Volume	**5**Volume	**4**Volume	**0**Volume
Avogadros law	**4**molecule	**5**molecule	**4**molecule	**0**Molecule

(4 + 5) **9** volumes/molecules of Ammonia/Oxygen mixture react to form **4** volumes/molecules of Nitrogen(II)oxide gas and no **vapour**.

Increase in pressure shift the equilibrium forward to the right where there is less volume/molecules. The yield of Nitrogen(II)oxide gas **increase**.

Decrease in pressure shift the equilibrium backward to the left where there is more volume/molecules. The yield of Nitrogen(II)oxide gas **decrease**.

(c)Influence of Temperature on dynamic equilibrium

A **decrease** in temperature favours the reaction that liberate/generate more heat thus **exothermic** reaction(**-ΔH**).

An **increase** in temperature favours the reaction that do not liberate /generate more heat thus **endothermic** reaction(**+ΔH**).

Endothermic reaction are thus favoured by high temperature/heating

Exothermic reaction are favoured by low temperature/cooling.

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If a reaction/equilibrium mixture is neither exothermic or endothermic, then a change in temperature/cooling/heating has no effect on the equilibrium position.

(i)Nitrogen(IV)oxide /Dinitrogen tetroxide mixture

Nitrogen(IV)oxide and dinitrogen tetraoxide can exist in dynamic equilibrium in a closed test tube. Nitrogen(IV)oxide is a brown gas. Dinitrogen tetraoxide is a yellow gas.

Chemical equation : $2NO_2(g) \rightleftharpoons N_2O_4(g)$

On heating /increasing temperature, the mixture becomes more brown. On cooling the mixture become more yellow.

This show that

(i)the forward reaction to the right is exothermic(**-ΔH**).

On heating an exothermic process the equilibrium shifts to the side that generate /liberate less heat.

(ii)the backward reaction to the right is endothermic(**+ΔH**).

On cooling an endothermic process the equilibrium shifts to the side that do not generate /liberate heat.

(c)Influence of Catalyst on dynamic equilibrium

A catalyst has no effect on the position of equilibrium. It only speeds up the rate of attainment. e.g.

Esterification of alkanols and alkanoic acids naturally take place in fruits.In the laboratory concentrated sulphuric(VI)acid catalyse the reaction.The equilibrium mixture forms the ester faster but the yield does not increase.

$CH_3CH_2OH(l)+CH_3COOH(l) \xrightleftharpoons{Conc.H_2SO_4} CH_3COOCH_2CH_3(aq) + H_2O(l)$

(d)Influence of rate of reaction and dynamic equilibrium (Optimum conditions) on industrial processes

Industrial processes are commercial profit oriented. All industrial processes take place in closed systems and thus in dynamic equilibrium.

For manufacturers, obtaining the highest yield at minimum cost and shortest time is paramount.

The conditions required to obtain the highest yield of products within the shortest time at minimum cost are called **optimum conditions**

Optimum condition thus require understanding the effect of various factors on:

(i)rate of reaction(**Chemical kinetics**)

(ii)dynamic equilibrium(**Chemical cybernetics**)

1.Optimum condition in Haber process

Chemical equation

$N_2(g) + 3H_2(g) \xrightarrow{Fe/Pt} 2NH_3(g) \quad \Delta H = -92kJ$

Equilibrium/Reaction rate considerations

(i)**Removing ammonia** gas once formed shift the equilibrium forward to the right to replace the ammonia. More/higher yield of ammonia is attained.

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(ii)**Increase in pressure** shift the equilibrium forward to the right where there is less volume/molecules . **More/higher yield** of ammonia is attained. Very **high** pressures raises the **cost** of production because they are **expensive** to produce and maintain. An optimum pressure of about 500atmospheres is normally used.
(iii)**Increase in temperature** shift the equilibrium backward to the left because the reaction is exothermic(ΔH = -92kJ) . Ammonia formed **decomposes** back to Nitrogen and Hydrogen to remove excess heat therefore a **less** yield of ammonia is attained. Very **low** temperature decrease the collision frequency of Nitrogen and Hydrogen and thus the rate of reaction **too slow** and **uneconomical**.
An optimum temperature of about 450ºC is normally used.
(iv)Iron and platinum can be used as catalyst. **Platinum** is a **better** catalyst but more **expensive** and easily **poisoned** by impurities than Iron. Iron is promoted /impregnated with AluminiumOxide(Al_2O_3) to increase its surface area/area of contact with reactants and thus efficiency.The catalyst does not increase the yield of ammonia but it speed up its rate of formation.

2.Optimum condition in Contact process

Chemical equation

$$2SO_2\,(g) + O_2\,(g) \xrightarrow{V_2O_5/Pt} 2SO_3\,(g) \qquad \Delta H = -197kJ$$

Equilibrium/Reaction rate considerations

(i)**Removing sulphur(VI)oxide** gas once formed shift the equilibrium forward to the right to replace the sulphur(VI)oxide. More/higher yield of sulphur(VI) oxide is attained.

(ii)**Increase in pressure** shift the equilibrium forward to the right where there is less volume/molecules . **More/higher yield** of sulphur(VI)oxide is attained. Very **high** pressures raises the **cost** of production because they are **expensive** to produce and maintain. An optimum pressure of about 1-2 atmospheres is normally used to attain about 96% yield of SO_3.
(iii)**Increase in temperature** shift the equilibrium backward to the left because the reaction is exothermic(ΔH = -197kJ) . Sulphur(VI)oxide formed **decomposes** back to Sulphur(IV)oxide and Oxygen to remove excess heat therefore a **less** yield of Sulphur(VI)oxide is attained. Very **low** temperature decrease the collision frequency of Sulphur(IV)oxide and Oxygen and thus the rate of reaction **too slow** and **uneconomical**.
An optimum temperature of about 450ºC is normally used.
(iv)Vanadium(V)Oxide and platinum can be used as catalyst. **Platinum** is a **better** catalyst and **less** easily **poisoned** by impurities but more **expensive.**
Vanadium(V)Oxide is very cheap even if it is easily poisoned by impurities. The catalyst does not increase the yield of Sulphur (VI)Oxide but it speed up its rate of formation.

3.Optimum condition in Ostwalds process

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Chemical equation

$4NH_3\ (g) + 5O_2\ (g) \xrightarrow{Pt/Rh} 4NO\ (g) + 6H_2O\ (g)\ \Delta H = -950kJ$

Equilibrium/Reaction rate considerations

(i)**Removing Nitrogen(II)oxide** gas once formed shift the equilibrium forward to the right to replace the Nitrogen(II)oxide. More/higher yield of Nitrogen(II) oxide is attained.

(ii)**Increase in pressure** shift the equilibrium backward to the left where there is less volume/molecules . **Less/lower yield** of Nitrogen(II)oxide is attained. Very **low** pressures increases the distance between reacting $NH_3$and O_2 molecules.
An optimum pressure of about **9** atmospheres is normally used.

(iii)**Increase in temperature** shift the equilibrium backward to the left because the reaction is exothermic($\Delta H = -950kJ$) . Nitrogen(II)oxide and water vapour formed **decomposes** back to Ammonia and Oxygen to remove excess heat therefore a **less** yield of Nitrogen(II)oxide is attained. Very **low** temperature decrease the collision frequency of Ammonia and Oxygen and thus the rate of reaction **too slow** and **uneconomical**.
An optimum temperature of about 900°C is normally used.

(iv)Platinum can be used as catalyst. **Platinum** is very **expensive.**It is:

-promoted with Rhodium to increase the surface area/area of contact.

-added/coated on the surface of asbestos to form platinized –asbestos to reduce the amount/quantity used.

The catalyst does not increase the yield of Nitrogen (II)Oxide but it speed up its rate of formation.

C.SAMPLE REVISION QUESTIONS

1.State two distinctive features of a dynamic equilibrium.

(i)the rate of forward reaction is equal to the rate of forward reaction

(ii)at equilibrium the concentrations of reactants and products do not change.

2. Explain the effect of increase in pressure on the following:

(i) $N_2(g) + O_2(g) \xrightarrow{} 2NO(g)$

Gay Lussacs law	1Volume	1Volume	2 Volume
Avogadros law	1 molecule	1 molecule	2 molecule

2 volume on reactant side produce 2 volume on product side.
Increase in pressure thus have no effect on position of equilibrium.

(ii) $2H_2(g) + CO(g) \xrightarrow{} CH_3OH\ (g)$

Gay Lussacs law	2Volume	1Volume	1 Volume
Avogadros law	2 molecule	1 molecule	1 molecule

3 volume on reactant side produce 1 volume on product side.

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Increase in pressure shift the equilibrium forward to the left. More yield of CH_3OH is formed.

4. Explain the effect of increasing temperature on the following:

$2SO_2(g) + O_2(g) \text{ ===== } 2SO_3(g)$ $\Delta H = -189kJ$ Forward reaction is exothermic. Increase in temperature shift the equilibrium backward to reduce the excess heat.

5.120g of brass an alloy of copper and Zinc was put it a flask containing dilute hydrochloric acid. The flask was placed on an electric balance. The readings on the balance were recorded as in the table below

Time(Seconds)	Mass of flask(grams)	Loss in mass(grams)
0	600	
20	599.50	
40	599.12	
60	598.84	
80	598.66	
100	598.54	
120	598.50	
140	598.50	
160	598.50	

(a)Complete the table by calculating the loss in mass

(b)What does the "600" gram reading on the balance represent

The initial mass of brass and the acid before any reaction take place.

(c)Plot a graph of Time (x-axes) against loss in mass.

(d)Explain the shape of your graph

The reaction produce hydrogen gas as one of the products that escape to the atmosphere. This decreases the mass of flask.After 120 seconds,the react is complete. No more hydrogen is evolved.The mass of flask remain constant.

(d)At what time was the loss in mass equal to:

(i)1.20g

Reading from a correctly plotted graph =

(ii)1.30g

Readng from a correctly plotted graph =

(iii)1.40g

Reading from a correctly plotted graph =

(e)What was the loss in mass at:

(i)50°C

Reading from a correctly plotted graph =

(ii) 70°C

Reading from a correctly plotted graph =

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(iii) 90°C g
Reading from a correctly plotted graph =

THERMOCHEMISTRY

1.Introduction to Energy changes

Energy is the capacity to do **work**. There are many/various **forms** of energy like heat, electric, mechanical, and/ or chemical energy.There are two types of energy:

(i)Kinetic Energy(KE) ;the energy in motion.
(ii)Potential Energy(PE); the stored/internal energy.

Energy like matter , is **neither created nor destroyed** but can be transformed /changed from one form to the other/ is interconvertible. This is the **principle of conservation** of energy. e.g. Electrical energy into heat through a filament in bulb.
Chemical and physical processes take place with **absorption** or **evolution**/production of energy mainly in form of **heat**
The study of energy changes that accompany physical/chemical reaction/changes is called **Thermochemistry.** Physical/chemical reaction/changes that involve energy changes are called **thermochemical reactions.** The SI unit of energy is the **Joule**(**J**).Kilo Joules(**kJ**)and megaJoules(**MJ**) are also used. The Joule(J) is defined as the:

(i) quantity of energy transferred when a force of one **newton** acts through a distance of one **metre**.
(ii) quantity of energy transferred when one **coulomb** of electric charge is passed through a potential difference of one **volt**.

All thermochemical reactions should be carried out at standard conditions of:

(i) **298K /25°C** temperature
(ii)**101300Pa/101300N/m^2 /760mmHg/1** atmosphere pressure.

2.Exothermic and endothermic processes/reactions

Some reactions / processes take place with evolution/production of energy. They are said to be **exothermic** while others take place with **absorption** of energy. They are said to be **endothermic.**
Practically exothermic reactions / processes cause a rise in temperature (by a rise in thermometer reading/mercury or alcohol level rise)
Practically endothermic reactions / processes cause a fall in temperature (by a fall in thermometer reading/mercury or alcohol level decrease)

To demonstrate/illustrate exothermic and endothermic processes/reactions

a) Dissolving Potassium nitrate(V)/ammonium chloride crystals
Procedure:

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Measure 20cm3 of water in a beaker. Determine and record its temperature T_1.Put about 1.0g of Potassium nitrate(V) crystals into the beaker. Stir the mixture carefully and note the highest temperature rise /fall T_2.Repeat the whole procedure by using ammonium chloride in place of Potassium nitrate (V) crystals.

Sample results

Temperture (°C)	**Using Potassium nitrate(V) crystals**	**Using Ammonium chloride crystals**
T_2(Final temperature)	21.0	23.0
T_1 (Initial temperature)	25.0	26.0
Change in temperature($T_2 - T_1$)	**4.0**	**3.0**

Note:

(i)Initial($\mathbf{T_1}$) temperature of dissolution of both potassium nitrate(V) crystals and ammonium chloride crystals is **higher** than the final temperature($\mathbf{T_2}$)

(ii) Change in temperature($T_2 - T_1$) is **not** a mathematical "**-4.0"** or "**-3.0"**.

(iii)Dissolution of both potassium nitrate(V) and ammonium chloride crystals is an **endothermic** process because initial($\mathbf{T_1}$) temperature is **higher** than the final temperature($\mathbf{T_2}$) thus causes a **fall/drop** in temperature.

b) Dissolving concentrated sulphuric(VI) acid/sodium hydroxide crystals

Procedure:

Measure 20cm3 of water in a beaker. Determine and record its temperature T_1.**Carefully** put about 1.0g/four pellets of sodium hydroxide crystals into the beaker. Stir the mixture carefully and note the highest temperature rise /fall T_2.Repeat the whole procedure by using 2cm3 of concentrated sulphuric(VI) acid in place of sodium hydroxide crystals.

CAUTION:

(i)Sodium hydroxide crystals are **caustic** and cause painful blisters on contact with skin.

(ii) Concentrated sulphuric (VI) acid is **corrosive** and cause painful wounds on contact with skin.

Sample results

Temperture (°C)	**Using Sodium hydroxide pellets**	**Using Concentrated sulphuric(VI) acid**
T_2(Final temperature)	30.0	32.0
T_1 (Initial temperature)	24.0	25.0
Change in temperature($T_2 - T_1$)	**6.0**	**7.0**

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Note:
(i)Initial ($\mathbf{T_1}$) temperature of dissolution of both concentrated sulphuric (VI) acid and sodium hydroxide pellets is **lower** than the final temperature ($\mathbf{T_2}$).
(ii)Dissolution of both Sodium hydroxide pellets and concentrated sulphuric (VI) acid is an **exothermic** process because final ($\mathbf{T_2}$) temperature is **higher** than the initial temperature ($\mathbf{T_1}$) thus causes a **rise** in temperature.
The above reactions show heat loss **to** and heat gain **from** the surrounding as illustrated by a **rise** and **fall** in temperature/thermometer readings.
Dissolving both potassium nitrate(V) and ammonium chloride crystals causes heat **gain from** the **surrounding** that causes **fall** in thermometer reading.
Dissolving both Sodium hydroxide pellets and concentrated sulphuric (VI) acid causes heat **loss to** the **surrounding** that causes **rise** in thermometer reading.
At the same temperature and pressure ,heat absorbed and released is called enthalpy/ heat content denoted **H.**
Energy change is measured from the heat content/enthalpy of the **final** and **initial** products. It is denoted **ΔH(**delta H**)**.i.e.

Enthalpy/energy/ change in heat content $\Delta H = H_{\mathbf{final}} - H_{\mathbf{initial}}$

For chemical reactions:

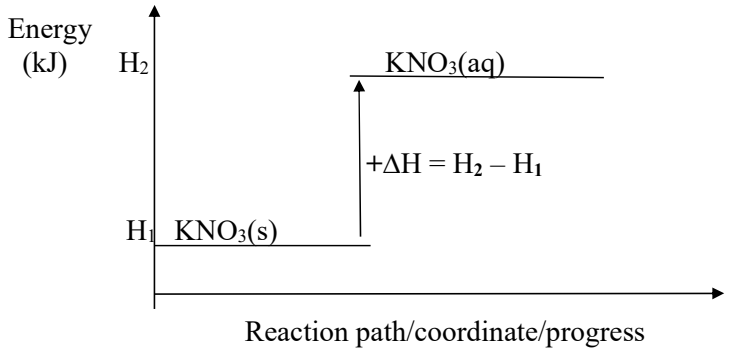
$\Delta H = H_{\mathbf{products}} - H_{\mathbf{reactants}}$

For exothermic reactions, the heat contents of the **reactants** is **more** than/**higher** than the heat contents of **products**, therefore the ΔH is negative (**-ΔH**)
For endothermic reactions, the heat contents of the **reactants** is **less** than/**lower** than the heat contents of **products**, therefore the ΔH is negative (**+ΔH**)
Graphically, in a **sketch** energy level diagram:

(i)For endothermic reactions the heat content of the reactants should be relatively/slightly **lower** than the heat content of the products

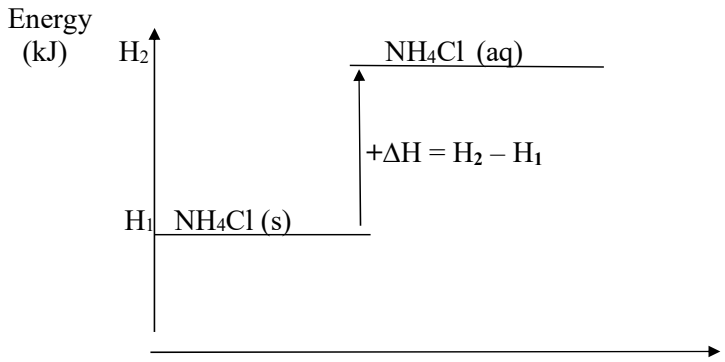
(ii)For exothermic reactions the heat content of the reactants should be relatively/slightly **higher** than the heat content of the products

Sketch energy level diagrams for endothermic dissolution

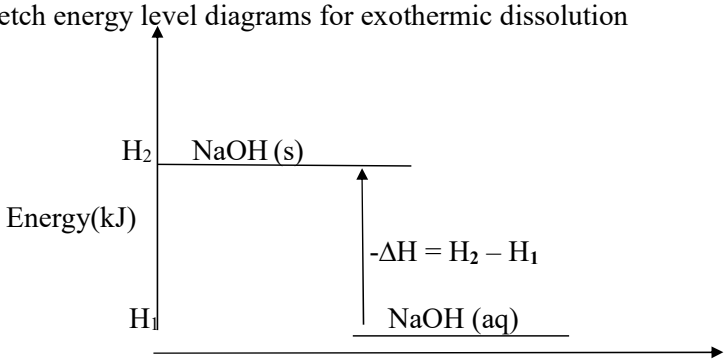


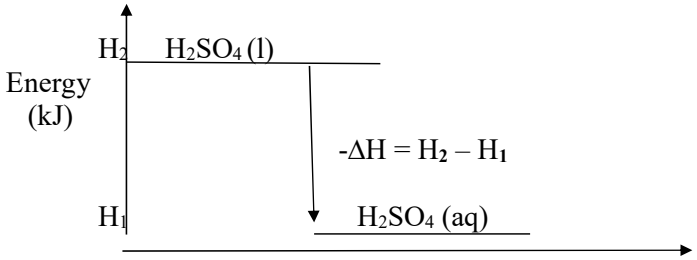
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Sketch energy level diagrams for exothermic dissolution





3.Energy changes in physical processes

Melting/freezing/fusion/solidification and **boiling/vaporization/evaporation** are the two physical processes. Melting /freezing point of pure substances is fixed /constant. The boiling point of pure substance depend on **external** atmospheric **pressure**.

Melting/fusion is the physical change of a **solid** to **liquid**. Freezing is the physical change of a **liquid** to **solid**.

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Melting/freezing/fusion/solidification are therefore two **opposite** but **same** reversible physical processes. i.e

A (**s**) ========A(**l**)

Boiling/vaporization/evaporation is the physical change of a **liquid** to **gas/vapour**.
Condensation/liquidification is the physical change of **gas/vapour** to **liquid**.
Boiling/vaporization/evaporation and condensation/liquidification are therefore two **opposite** but **same** reversible physical processes. i.e

B (**l**) ========B(**g**)

Practically

(i) Melting/liquidification/fusion involves **heating** a solid to **weaken** the strong bonds holding the solid particles together. Solids are made up of very strong bonds holding the particles **very close** to each other (**Kinetic Theory of matter**).On heating these particles gain energy/heat from the surrounding heat source to form a liquid with **weaker** bonds holding the particles close together but with some degree of **freedom**. Melting/freezing/fusion is an **endothermic** (+ΔH)process that require/absorb energy from the surrounding.

(ii)Freezing/fusion/solidification involves cooling a a liquid to reform /rejoin the very strong bonds to hold the particles **very close** to each other as solid and thus lose their degree of **freedom** (**Kinetic Theory of matter**). Freezing /fusion / solidification is an **exothermic** (-ΔH)process that require particles holding the liquid together to lose energy to the surrounding.

(iii)Boiling/vaporization/evaporation involves **heating** a liquid to completely **break/free** the bonds holding the liquid particles together. Gaseous particles have high degree of **freedom** (**Kinetic Theory of matter**). Boiling /vaporization / evaporation is an **endothermic** (+ΔH) process that require/absorb energy from the surrounding.

(iv)Condensation/liquidification is **reverse** process of boiling /vaporization / evaporation.It involves gaseous particles losing energy to the surrounding to form a liquid.It is an **exothermic**(+ΔH) process.

The quantity of energy required to **change** one mole of a solid **to** liquid or to **form** one mole of a solid **from** liquid at constant temperature is called **molar enthalpy/latent heat of fusion**. e.g.

H_2O(**s**) -> H_2O(**l**) ΔH = **+6.0kJ mole^{-1}** (endothermic process)
H_2O(**l**) -> H_2O(**s**) ΔH = **-6.0kJ mole^{-1}** (exothermic process)

The quantity of energy required to **change** one mole of a liquid **to** gas/vapour or to **form** one mole of a liquid **from** gas/vapour at constant temperature is called **molar enthalpy/latent heat of vapourization**. e.g.

H_2O(**l**) -> H_2O(**g**) ΔH = **+44.0kJ mole^{-1}** (endothermic process)
H_2O(**g**) -> H_2O(**l**) ΔH = **-44.0kJ mole^{-1}** (exothermic process)

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The following experiments illustrate/demonstrate practical determination of melting and boiling

a) To determine the boiling point of water

Procedure:

Measure 20cm3 of tap water into a 50cm3 glass beaker. Determine and record its temperature.Heat the water on a strong Bunsen burner flame and record its temperature after every thirty seconds for four minutes.

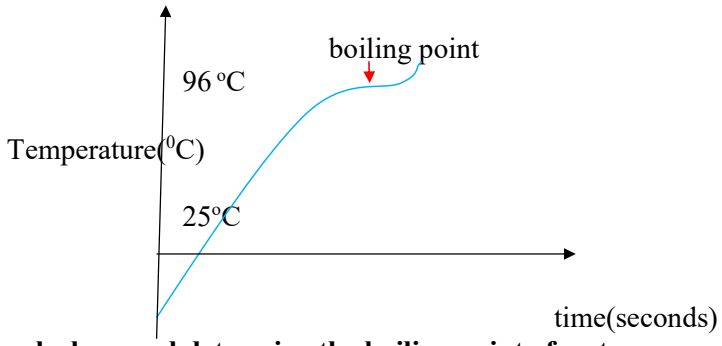
Sample results

Time(seconds)	0	30	60	90	120	150	180	210	240
Temperature(°C)	25.0	45.0	85.0	95.0	96.0	96.0	96.0	97.0	98.0

Questions

1.Plot a graph of temperature against time(y-axis)

Sketch graph of temperature against time



2.From the graph show and determine the boiling point of water

Note:

Water boils at 100°C at sea level/one atmosphere pressure/101300Pa **but** boils at **below** 100°C at **higher** altitudes. The sample results above are from Kiriari Girls High School-Embu County on the slopes of Mt Kenya in Kenya. Water here boils at 96°C.

3.Calculate the molar heat of vaporization of water.(H= 1.0,O= 16.O)

Working:

Mass of water = density x volume => (20 x 1) /1000 = **0.02kg**

Quantity of heat produced

= mass of water x specific heat capacity of water x temperature change

=>0.02kg x 4.2 x (96 – 25) = **5.964kJ**

Heat of vaporization of one mole H_2O = $\frac{\text{Quantity of heat}}{\text{Molar mass of } H_2O}$

=>$\frac{5.964kJ}{18}$ = **0.3313 kJ mole $^{-1}$**

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To determine the melting point of candle wax

Procedure

Weigh exactly 5.0 g of candle wax into a boiling tube. Heat it on a strongly Bunsen burner flame until it completely melts. Insert a thermometer and remove the boiling tube from the flame. Stir continuously. Determine and record the temperature after every 30seconds for four minutes.

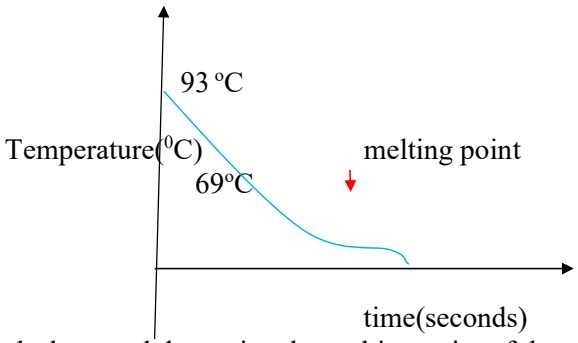
Sample results

Time(seconds)	0	30	60	90	120	150	180	210	240
Temperature(°C)	93.0	85.0	78.0	70.0	69.0	69.0	69.0	67.0	65.0

Questions

1.Plot a graph of temperature against time(y-axis)

Sketch graph of temperature against time



2.From the graph show and determine the melting point of the candle wax

4.Energy changes in chemical processes

Thermochemical reactions measured at **standard** conditions of 298K(25°C) and 101300Pa/101300Nm^2/ 1 atmospheres/760mmHg/76cmHg produce standard **enthalpies** denoted ΔH^{θ}.

Thermochemical reactions are named from the type of reaction producing the energy change. Below are some thermochemical reactions:

(a) Standard enthalpy/heat of reaction ΔH^{θ}_r
(b) Standard enthalpy/heat of combustion ΔH^{θ}_c
(c) Standard enthalpy/heat of displacement ΔH^{θ}_d
(d) Standard enthalpy/heat of neutralization ΔH^{θ}_n
(e) Standard enthalpy/heat of solution/dissolution ΔH^{θ}_s
(f) Standard enthalpy/heat of formation ΔH^{θ}_f

(a)Standard enthalpy/heat of reaction ΔH^{θ}_r

The molar standard enthalpy/heat of reaction may be defined as the energy/heat change when one mole of products is formed at standard conditions

A chemical reaction involves the reactants forming products. For the reaction to take place the bonds holding the reactants must be broken so that new bonds of the products are formed. i.e.

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A-B + C-D -> **A**-C + **B**-D

Old Bonds **broken** A-B and C-D on reactants

New Bonds **formed** A-C and B-D on products

The energy required to break one mole of a (covalent) bond is called **bond dissociation energy**. The SI unit of bond dissociation energy is **kJmole^{-1}**

The higher the bond dissociation energy the stronger the (covalent)bond

Bond dissociation energies of some (covalent)bonds

Bond	Bond dissociation energy (kJmole^{-1})		Bond dissociation energy (kJmole^{-1})
H-H	431	I-I	151
C-C	436	C-H	413
C=C	612	O-H	463
C≡C	836	C-O	358
N≡N	945	H-Cl	428
N-H	391	H-Br	366
F-F	158	C-Cl	346
Cl-Cl	239	C-Br	276
Br-Br	193	C-I	338
H-I	299	O=O	497
Si-Si	226	C-F	494

The molar enthalpy of reaction can be calculated from the bond dissociation energy by:

(i)**adding** the total bond dissociation energy of the **reactants**(endothermic process/+Δ**H**) and total bond dissociation energy of the **products**(exothermic process/-Δ**H**).

(ii)**subtracting** total bond dissociation energy of the **reactants from** the total bond dissociation energy of the **products**(exothermic process/**-ΔH less/minus** endothermic process/+Δ**H**).

Practice examples/Calculating ΔH_r

1.Calculate ΔH_r from the following reaction:

a) $H_2(g)$ + $Cl_2(g)$ -> 2HCl(g)

Working

Old bonds broken (endothermic process/+Δ**H**)

= (H-H + Cl-Cl) => (+431 + (+ 239)) = **+ 670kJ**

New bonds broken (exothermic process/**-ΔH**)

= (2(H-Cl) => (- 428 x 2)) = **-856kJ**

$\Delta H_r = \frac{(+ 670kJ + -856kJ)}{2} = \frac{186\ kJ}{2} =$ **-93kJ mole^{-1}**

The above reaction has negative **-ΔH** enthalpy change and is therefore practically exothermic.

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The thermochemical reaction is thus:

½ $H_2(g)$ + ½ $Cl_2(g)$ -> HCl(g) ΔH_r = -93kJ

b) $CH_4(g)$ + $Cl_2(g)$ -> CH_3Cl + HCl(g)

Working

Old bonds broken (endothermic process/+Δ**H**)

= (4(C-H) + Cl-Cl)

=> ((4 x +413) + (+ 239)) = **+ 1891kJ**

New bonds broken (exothermic process/-Δ**H**)

= (3(C-H + H-Cl + C-Cl)

=> ((3 x - 413) + 428 + 346) = **-2013 kJ**

ΔH_r =(+ 1891kJ + -2013 kJ) = **-122 kJ mole^{-1}**

The above reaction has negative **-ΔH** enthalpy change and is therefore practically exothermic.

The thermochemical reaction is thus:

$CH_4(g)$ + $Cl_2(g)$ -> $CH_3Cl(g)$ + HCl(g) ΔH = -122 kJ

c) $CH_2CH_2(g)$ + $Cl_2(g)$ -> CH_3Cl CH_3Cl (g)

Working

Old bonds broken (endothermic process/+Δ**H**)

= (4(C-H) + Cl-Cl + C=C)

=> ((4 x +413) + (+ 239) +(612)) = **+ 2503kJ**

New bonds broken (exothermic process/-Δ**H**)

= (4(C-H + C-C + 2(C-Cl))

=> ((3 x - 413) + -436 +2 x 346 = **-2367 kJ**

ΔH_r =(+ 2503kJ + -2367 kJ) = **+136 kJ mole^{-1}**

The above reaction has negative **+ΔH** enthalpy change and is therefore practically endothermic.

The thermochemical reaction is thus:

$CH_2CH_2(g)$ + $Cl_2(g)$ -> CH_3Cl CH_3Cl (g) ΔH = +136 kJ

Note that:

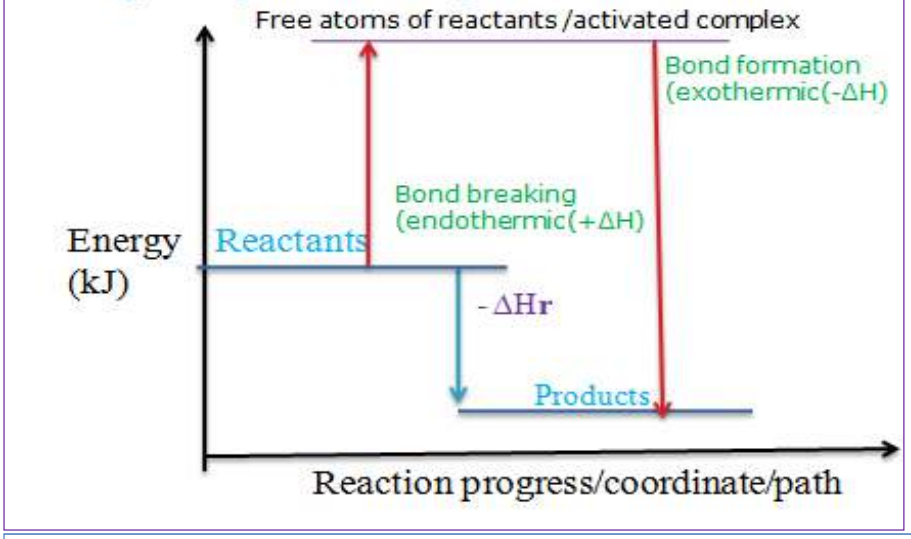
(i)a reaction is **exothermic** if the bond dissociation energy of **reactants** is **more** than bond dissociation energy of products.

(ii)a reaction is **endothermic** if the bond dissociation energy of **reactants** is **less** than bond dissociation energy of products.

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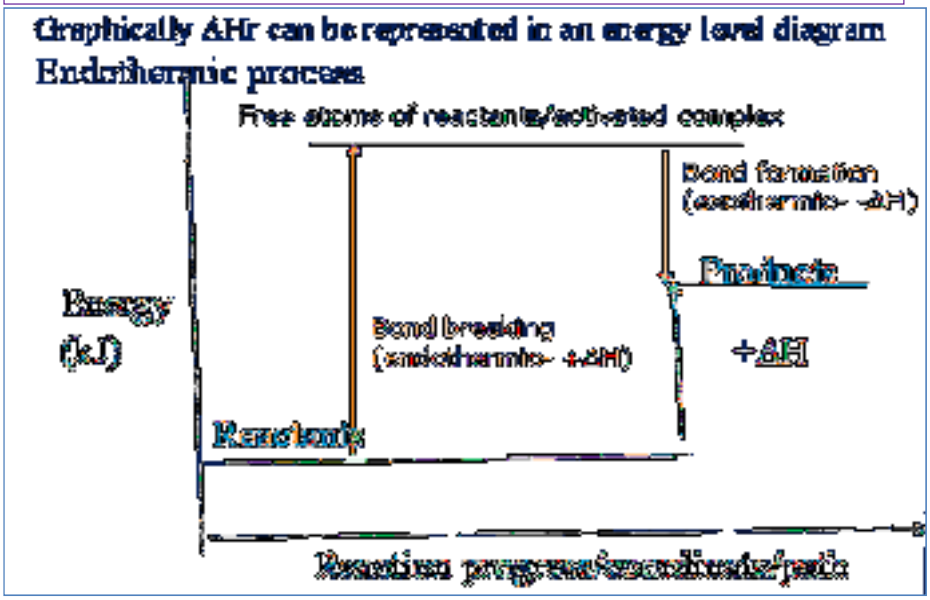
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Sketch energy level diagrams for exothermic process



Graphically ΔHr can be represented in an energy level diagram

Endothermic process



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(b)Standard enthalpy/heat of combustion ΔH^{θ}_{c}

The molar standard enthalpy/heat of **combustion(ΔH^{θ}_{c})** is defined as the energy/heat change when one mole of a substance is burnt in oxygen/excess air at standard conditions.

Burning is the reaction of a substance with oxygen/air. It is an exothermic process producing a lot of energy in form of heat.

A substance that undergoes burning is called a **fuel.** A fuel is defined as the combustible substance which burns in air to give heat energy for domestic or industrial use. A fuel may be **solid** (e.g coal, wood, charcoal) **liquid** (e.g petrol, paraffin, ethanol, kerosene) or **gas** (e.g liquefied petroleum gas/LPG, Water gas-CO_2/H_2, biogas-methane, Natural gas-mixture of hydrocarbons)

To determine the molar standard enthalpy/heat of **combustion(ΔH^{θ}_{c})** of ethanol

Procedure

Put 20cm3 of distilled water into a 50cm3 beaker. Clamp the beaker. Determine the temperature of the water T_1.Weigh an empty burner(empty tin with wick). Record its mass M_1.Put some ethanol into the burner. Weigh again the burner with the ethanol and record its mass M_2. Ignite the burner and place it below the clamped 50cm3 beaker. Heat the water in the beaker for about one minute. Put off the burner. Record the highest temperature rise of the water, T_2. Weigh the burner again and record its mass M_3

Sample results:

Volume of water used	20cm3
Temperature of the water before heating T_1	25.0°C
Temperature of the water after heating T_2	35.0°C
Mass of empty burner M_1	28.3g
Mass of empty burner + ethanol before igniting M_2	29.1g
Mass of empty burner + ethanol after igniting M_3	28.7g

Sample calculations:

1.Calculate:

(a) ΔT the change in temperature

$\Delta T = T_2 - T_1$ => (35.0°C – 25.0°C) = **10.0°C**

(b) the mass of ethanol used in burning

mass of ethanol used = $M_2 - M_1$ => 29.1g – 28.7g = **0.4g**

(c) the number of moles of ethanol used in burning

moles of ethanol = $\frac{\text{mass used}}{\text{molar mass of ethanol}}$ => $\frac{0.4}{46}$ = **$0.0087 / 8.7 \times 10^{-3}$** moles

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2. Given that the specific heat capacity of water is 4.2 $kJ^{-1}kg^{-1}K^{-1}$,determine the heat produced during the burning.

Heat produced Δ**H** = mass of water(**m**) x specific heat capacity (**c**)x Δ**T**

=> 20 x 4.2 x 10 = **840 J**oules = **0.84 kJ**

1000

3.Calculate the molar heat of combustion of ethanol

Molar heat of combustion ΔH_c = Heat produced ΔH

Number of moles of fuel

=> 0.84 kJ = **96.5517 $kJmole^{-1}$**

0.0087 /8.7 x 10^{-3} moles

4.List two sources of error in the above experiment.

(i)Heat loss to the surrounding **lowers** the practical value of the molar heat of combustion of ethanol.

A draught shield tries to minimize the loss by protecting wind from wobbling the flame.

(ii) Heat gain by reaction vessels/beaker **lowers** ΔT and hence ΔH_c

5.Calculate the heating value of the fuel.

Heating value = molar heat of combustion => 96.5517 $kJmole^{-1}$ = **2.0989 kJg^{-1}**

Molar mass of fuel 46 g

Heating value is the enrgy produced when a unit mass/gram of a fuel is completely burnt

6.Explain other factors used to determine the choice of fuel for domestic and industrial use.

(i) **availability and affordability**-some fuels are more available cheaply in rural than in urban areas at a lower cost.

(ii)**cost of storage and transmission**-a fuel should be easy to transport and store safely. e.g LPG is very convenient to store and use. Charcoal and wood are bulky.

(iii)**environmental effects –**Most fuels after burning produce carbon(IV) oxide gas as a byproduct. Carbon(IV) oxide gas is green house gas that causes global warming. Some other fuel produce acidic gases like sulphur(IV) oxide ,and nitrogen(IV) oxide. These gases cause acid rain. Internal combustion engines exhaust produce lead vapour from leaded petrol and diesel. Lead is carcinogenic.

(iv)**ignition point-**The temperature at which a fuel must be heated before it burns in air is the ignition point. Fuels like petrol have very low ignition point, making it highly flammable. Charcoal and wood have very high ignition point.

7.Explain the methods used to reduce pollution from common fuels.

(i)Planting trees-Plants absorb excess carbon(IV)oxide for photosynthesis and release oxygen gas to the atmosphere.

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(ii)using catalytic converters in internal combustion engines that convert harmful/toxic/poisonous gases like carbon(II)oxide and nitrogen(IV)oxide to harmless non-poisonous carbon(IV)oxide, water and nitrogen gas by using platinum-rhodium catalyst along the engine exhaust pipes.

Further practice calculations

1.Calculate the heating value of methanol CH_3OH given that 0.87g of the fuel burn in air to raise the temperature of 500g of water from 20°C to 27°C.(C-12.0,H=1.0 O=16.0).

Moles of methanol used = Mass of methanol used => 0.87 g = **0.02718 moles**
Molar mass of methanol 32

Heat produced **ΔH** = mass of water(**m**) x specific heat capacity (**c**)x **ΔT**
=> 500 x 4.2 x 7 = 14700 **J**oules = **14.7 kJ**
1000

Molar heat of combustion ΔH_c = Heat produced ΔH
Number of moles of fuel
=> 14.7 kJ = **540.8389 kJmole^{-1}**
0.02718 moles

Heating value = molar heat of combustion => 540.8389 kJmole^{-1} = **16.9012 kJg^{-1}**
Molar mass of fuel 32 g

2. 1.0 g of carbon burn in excess air to raise the temperature of 400g of water by 18°C.Determine the molar heat of combustion and hence the heating value of carbon(C-12.0,).

Moles of carbon used = Mass of carbon used => 1.0 g = **0.0833 moles**
Molar mass of carbon 12

Heat produced **ΔH** = mass of water(**m**) x specific heat capacity (**c**)x **ΔT**
=> 400 x 4.2 x 18 = 30240 **J**oules = **30.24 kJ**
1000

Molar heat of combustion ΔH_c = Heat produced ΔH
Number of moles of fuel
=> 30.24 kJ = **363.0252 kJmole^{-1}**
0.0833 moles

Heating value = molar heat of combustion => 363.0252 kJmole^{-1}= **30.2521 kJg^{-1}**
Molar mass of fuel 12 g

(c)Standard enthalpy/heat of displacement ΔH^{θ}_{d}

The molar standard enthalpy/heat of **displacement** ΔH^{θ}_{d} is defined as the energy/heat change when one mole of a substance is displaced from its solution. A displacement reaction takes place when a more reactive element/with less electrode potential **E^{θ}** / negative E^{θ} /higher in the reactivity/electrochemical series remove/displace another with less reactive element/with higher electrode potential **E^{θ}** / positive E^{θ} /lower in the reactivity/electrochemical series **from** its solution.e.g.

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(i)$Zn(s) + CuSO_4(aq) \rightarrow Cu(s) + ZnSO_4(aq)$
Ionically: $Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}(aq)$
(ii)$Fe(s) + CuSO_4(aq) \rightarrow Cu(s) + FeSO_4(aq)$
Ionically: $Fe(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Fe^{2+}(aq)$
(iii)$Pb(s) + CuSO_4(aq) \rightarrow Cu(s) + PbSO_4(\mathbf{s})$
This reaction stops after some time as insoluble $PbSO_4(\mathbf{s})$ coat/cover unreacted lead.
(iv)$Cl_2(g) + 2NaBr(aq) \rightarrow Br_2(aq) + 2NaCl(aq)$
Ionically: $Cl_2(g) + 2Br^-(aq) \rightarrow Br_2(aq) + 2Cl^-(aq)$
Practically, a displacement reaction takes place when a known amount /volume of a solution is added excess of a more reactive metal.
To determine the molar standard enthalpy/heat of **displacement(ΔH^{θ}_d)** of copper
Procedure
Place 20cm3 of 0.2M copper(II)sulphate(VI)solution into a 50cm3 plastic beaker/calorimeter. Determine and record the temperature of the solution T_1.Put all the Zinc powder provided into the plastic beaker. Stir the mixture using the thermometer. Determine and record the highest temperature change to the nearest 0.5°C- T_2 . Repeat the experiment to complete table 1 below
Table 1

Experiment	I	II
Final temperature of solution(T_2)	30.0°C	31.0°C
Final temperature of solution(T_1)	25.0°C	24.0°C
Change in temperature(ΔT)	5.0	6.0

Questions
1.(a) Calculate:
(i)average ΔT
AverageΔT = change in temperature in experiment I and II
$\Rightarrow \frac{5.0 + 6.0}{2} =$ **5.5°C**
(ii)the number of moles of solution used
Moles used = molarity x volume of solution $= \frac{0.2 \times 20}{1000} =$ **0.004 moles**
(iii)the enthalpy change ΔH for the reaction
Heat produced $\Delta \mathbf{H}$ = mass of solution(**m**) x specific heat capacity (**c**)x $\Delta \mathbf{T}$
$\Rightarrow \frac{20 \times 4.2 \times 5.5}{1000} = 462$ **J**oules = **0.462 kJ**
(iv)State two assumptions made in the above calculations.
Density of solution = density of water = $1 gcm^{-3}$
Specific heat capacity of solution=Specific heat capacity of solution=$4.2 kJ^{-1}kg^{-1}K$
This is because the solution is assumed to be **infinite dilute.**

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2. Calculate the enthalpy change for one mole of displacement of Cu^{2+} (aq) ions.

Molar heat of displacement ΔH_d = Heat produced ΔH
Number of moles of fuel

$$\Rightarrow \quad \frac{0.462 \text{ kJ}}{0.004} = \mathbf{115.5\ kJmole^{-1}}$$

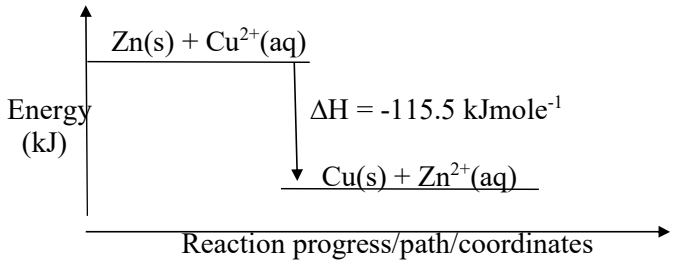
3.Write an ionic equation for the reaction taking place.

$Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}(aq)$

4.State the observation made during the reaction.

Blue colour of copper(II)sulphate(VI) fades/becomes less blue/colourless.
Brown solid deposits are formed at the bottom of reaction vessel/ beaker.

5.Illustrate the above reaction using an energy level diagram.



6. Iron is less reactive than Zinc. Explain the effect of using iron instead of Zinc on the standard molar heat of displacement ΔH_d of copper(II)sulphate (VI) solution.

No effect. Cu^{2+} (aq) are displaced from their solution.The element used to displace it does not matter.The reaction however faster if a more reactive metal is used.

7.(a)If the standard molar heat of displacement ΔH_d of copper(II)sulphate (VI) solution is 209kJmole^{-1} calculate the temperature change if 50cm3 of 0.2M solution was displaced by excess magnesium.

Moles used = molarity x volume of solution $= \frac{0.2 \times 50}{1000} = \frac{}{1000} = \mathbf{0.01\ moles}$

Heat produced ΔH = Molar heat of displacement ΔH_d x Number of moles
$\Rightarrow$ 209kJmole^{-1} x 0.01 moles = **2.09 kJ**

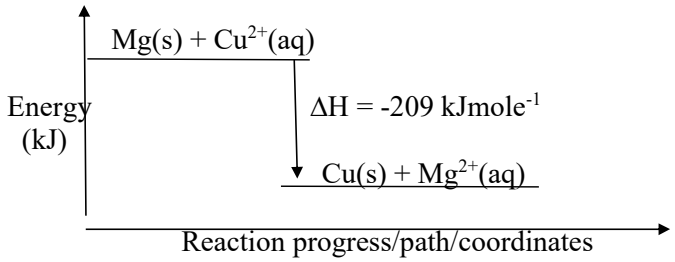
ΔT (change in temperature) = Heat produced ΔH
Molar heat of displacement ΔH_d x Number of moles

$$\Rightarrow \frac{2.09 \text{ kJ}}{0.01 \text{ moles}} = \mathbf{9.9524}\text{Kelvin}$$

(b)Draw an energy level diagram to show the above energy changes

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8. The enthalpy of displacement ΔH_d of copper(II)sulphate (VI) solution is 12k6kJmole^{-1}.Calculate the molarity of the solution given that 40cm3 of this solution produces 2.204kJ of energy during a displacement reaction with excess iron filings.

Number of moles = $\frac{\text{Heat produced } \Delta H}{\text{Molar heat of displacement } \Delta H_d}$

$= \frac{2.204 \text{ kJ}}{126 \text{ moles}} =$ **0.0206moles**

Molarity of the solution = $\frac{\text{moles} \times 1000}{\text{Volume of solution used}}$

$= \frac{0.0206\text{moles} \times 1000}{40} =$ **0.5167 M**

9. If the molar heat of displacement of Zinc(II)nitrate(V)by magnesium powder is 25.05kJmole^{-1} ,calculate the volume of solution which must be added 0.5 moles solution if there was a 3.0K rise in temperature.

Heat produced ΔH = Molar heat of displacement ΔH_d x Number of moles

=>25.08kJmole^{-1}x 0.5 moles = **1.254 kJ** x 1000 **=1254J**

Mass of solution (**m**) = $\frac{\text{Heat produced } \Delta \mathbf{H}}{\text{specific heat capacity } (\mathbf{c}) \times \Delta \mathbf{T}}$

$\Rightarrow \frac{1254\text{J}}{4.2 \times 3} =$ **99.5238 g**

Volume = mass x density = 99.5238 g x 1 = **99.5238cm3**

Note: The solution assumes to be too dilute /infinite dilute such that the density and specific heat capacity is assumed to be that of water.

Graphical determination of the molar enthalpy of displacement of copper

Procedure:

Place 20cm3 of 0.2M copper(II)sulphate (VI) solution into a calorimeter/50cm3 of plastic beaker wrapped in cotton wool/tissue paper.

Record its temperature at time T= 0.

Stir the solution with the thermometer carefully and continue recording the temperature after every 30 seconds .

Place all the (1.5g) Zinc powder provided.

Stir the solution with the thermometer carefully and continue recording the temperature after every 30 seconds for five minutes.

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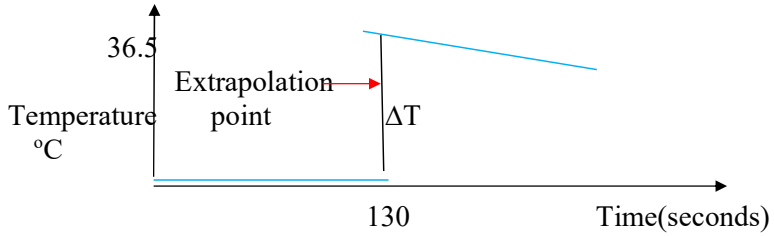
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Determine the highest temperature change to the nearest 0.5°C.

Sample results

Time °C	0.0	30.0	60.0	90.0	120.0	150.0	180.0	210.0	240.0	270.0
Temperature	25.0	25.0	25.0	25.0	25.0	xxx	36.0	35.5	35.0	34.5

Sketch graph of temperature against time



Questions

1. **Show and determine the change in temperature ΔT**

From a well constructed graph ΔT= T2 –T1 at **150** second by **extrapolation**

ΔT = 36.5 – 25.0 = 11.5°C

2. **Calculate the number of moles of copper(II) sulphate(VI)used given the molar heat of displacement of Cu^{2+} (aq)ions is 125kJmole^{-1}**

Heat produced **ΔH** = mass of solution(**m**) x specific heat capacity (**c**)x **ΔT**

$$\Rightarrow \frac{20 \times 4.2 \times 11.5}{1000} = 966 \text{ Joules} = \mathbf{0.966\ kJ}$$

$$\text{Number of moles} = \frac{\text{Heat produced } \Delta H}{\text{Molar heat of displacement } \Delta H_d}$$

$$\Rightarrow \frac{.966 \text{ kJ}}{125 \text{ moles}} = \mathbf{0.007728 moles} = \mathbf{7.728 \times 10^{-3} moles}$$

2. **What was the concentration of copper(II)sulphate(VI) in moles per litre.**

$$\text{Molarity} = \frac{\text{moles} \times 1000}{\text{Volume used}} \Rightarrow \frac{7.728 \times 10^{-3} \text{moles} \times 1000}{20} = \mathbf{0.3864M}$$

4. **The actual concentration of copper(II)sulphate(VI) solution was 0.4M.Explain the differences between the two.**

Practical value is **lower** than theoretical. Heat/energy **loss** to the surrounding and that absorbed by the reaction vessel **decreases** ΔT hence **lowering** the practical number of moles and molarity against the theoretical value

5.a) In an experiment to determine the molar heat of reaction when magnesium displaces copper ,0.15g of magnesium powder were added to 25.0cm^3 of 2.0M copper (II) chloride solution. The temperature of copper (II) chloride solution was 25°C.While that of the mixture was 43°C.

i)Other than increase in temperature, state and explain the observations which were made during the reaction.(3mks)

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ii)Calculate the heat change during the reaction (specific heat capacity of the solution = 4.2jg^{-1}k^{-1}and the density of the solution = 1g/cm^3(2mks)
iii)Determine the molar heat of displacement of copper by magnesium.(Mg=24.0).
iv)Write the ionic equation for the reaction.(1mk)
v)Sketch an energy level diagram for the reaction.(2mks)

(c)Standard enthalpy/heat of neutralization ΔH^{θ}_{n}

The molar standard enthalpy/heat of **neutralization** ΔH^{θ}_{n} is defined as the energy/heat change when one mole of a H^+ (H_3O^+)ions react completely with one mole of OH^- ions to form one mole of H_2O/water.
Neutralization is thus a reaction of an acid /H^+ (H_3O^+)ions with a base/alkali/ OH^- ions to form salt and water only.
Strong acids/bases/alkalis are completely dissociated to **many** free ions(H^+ /H_3O^+ and OH^- ions).
Weak acids/bases/alkalis are partially dissociated to **few** free ions(H^+ (H_3O^+ and OH^- ions) and exist **more** as molecules.
Neutralization is an exothermic(-ΔH) process.The enrgy produced during neutralization depend on the amount of free ions (H^+ H_3O^+ and OH^-)ions existing in the acid/base/alkali reactant:

(i)for weak acid-base/alkali neutralization,some of the **energy** is used to dissociate /ionize the **molecule** into free H^+ H_3O^+ and OH^- ions therefore the overall energy evolved is comparatively **lower/lesser/smaller** than strong acid / base/ alkali neutralizations.

(ii) (i)for strong acid/base/alkali neutralization, no **energy** is used to dissociate /ionize since **molecule** is wholly/fully dissociated/ionized into free H^+ H_3O^+ and OH^- ions.The overall energy evolved is comparatively **higher/more** than weak acid-base/ alkali neutralizations. For strong acid-base/alkali neutralization, the enthalpy of neutralization is **constant** at about **57.3**kJmole^{-1} irrespective of the acid-base used. This is because ionically:

$OH^-(aq)$+ $H^+(aq)$ -> $H_2O(l)$ for any wholly dissociated acid/base/alkali
Practically ΔH^{θ}_{n} can be determined as in the examples below:
To determine the molar enthalpy of neutralization ΔH_n of Hydrochloric acid

Procedure

Place 50cm3 of 2M hydrochloric acid into a calorimeter/200cm3 plastic beaker wrapped in cotton wool/tissue paper. Record its temperature T_1.Using a clean measuring cylinder, measure another 50cm3 of 2M sodium hydroxide. Rinse the bulb of the thermometer in distilled water. Determine the temperature of the sodium hydroxide T_2.Average T_2 andT_1 to get the initial temperature of the mixture T_3.
Carefully add all the alkali into the calorimeter/200cm3 plastic beaker wrapped in cotton wool/tissue paper containing the acid. Stir vigorously the mixture with the thermometer.

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Determine the highest temperature change to the nearest 0.5°C T_4 as the final temperature of the mixture. Repeat the experiment to complete table 1.

Table I . **Sample results**

Experiment	I	II
Temperature of acid T_1 (°C)	22.5	22.5
Temperature of base T_2 (°C)	22.0	23.0
Final temperature of solution T_4(°C)	35.5	36.0
Initial temperature of solution T_3(°C)	22.25	22.75
Temperature change(T_5)	13.25	13.75

(a)Calculate T_6 the average temperature change

$T_6 = \frac{13.25 + 13.75}{2} = 13.5\,^oC$

(b)Why should the apparatus be very clean?

Impurities present in the apparatus reacts with acid /base lowering the overall temperature change and hence ΔH^{θ}_n.

(c)Calculate the:

(i)number of moles of the acid used

number of moles = molarity x volume $\Rightarrow \frac{2 \times 50}{1000} = \frac{}{1000}$ **0.1moles**

(ii)enthalpy change ΔH of neutralization.

ΔH = (**m**)mass of solution(acid+base) x (**c**)specific heat capacity of solution x Δ**T**(T_6) => (50 +50) x 4.2 x 13.5 = **5670J**oules = **5.67kJ**

(iii) the molar heat of neutralization the acid.

ΔH_n = Enthalpy change ΔH => 5.67kJ = **56.7kJ mole^{-1}**
Number of moles 0.1moles

(c)Write the ionic equation for the reaction that takes place

$OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$

(d)The theoretical enthalpy change is 57.4kJ. Explain the difference with the results above.

The theoretical value is higher

Heat/energy loss to the surrounding/environment lowers Δ**T**/T_6 and thus ΔH_n

Heat/energy is absorbed by the reaction vessel/calorimeter/plastic cup lowers ΔT and hence ΔH_n

(e)Compare the ΔH_n of the experiment above with similar experiment repeated with neutralization of a solution of:

(i) potassium hydroxide with nitric(V) acid

The results would be the same/similar.

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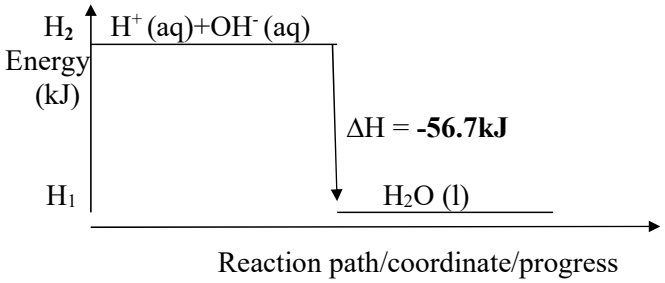
Both are neutralization reactions of strong acids and bases/alkalis that are fully /wholly dissociated into many free H^+ / H_3O^+ and OH^- ions.

(ii) ammonia with ethanoic acid

The results would be lower/ΔH_n would be less.

Both are neutralization reactions of weak acids and bases/alkalis that are partially /partly dissociated into few free H^+ / H_3O^+ and OH^- ions. Some energy is used to ionize the molecule.

(f)Draw an energy level diagram to illustrate the energy changes



Theoretical examples

1.The molar enthalpy of neutralization was experimentary shown to be 51.5kJ per mole of 0.5M hydrochloric acid and 0.5M sodium hydroxide. If the volume of sodium hydroxide was 20cm3, what was the volume of hydrochloric acid used if the reaction produced a 5.0°C rise in temperature?

Working:

Moles of sodium hydroxide = molarity x volume => $\frac{0.5 \text{ M} \times 20\text{cm}^3}{1000}$ = $\frac{}{1000}$ = **0.01 moles**

Enthalpy change ΔH = $\frac{\Delta H_n}{\text{Moles sodium hydroxide}}$ => $\frac{51.5}{0.01 \text{ moles}}$ = **0.515kJ**

Mass of base + acid = $\frac{\text{Enthalpy change } \Delta H \text{ in Joules}}{\text{Specific heat capacity} \times \Delta T}$

=> $\frac{0.515\text{kJ} \times 1000}{4.2 \times 5}$ = **24.5238**g

Mass/volume of HCl = Total volume – volume of NaOH

=>24.5238 - 20.0 = **4.5238 cm3**

3. **ΔH_n of potassium hydroxide was practically determined to be 56.7kJmole^{-1}.Calculate the molarity of 50.0 cm3 potassium hydroxide used to neutralize 25.0cm3 of dilute sulphuric(VI) acid raising the temperature of the solution from 10.0°C to 16.5°C.**

ΔH = (**m**)mass of solution(acid+base) x (**c**)specific heat capacity of solution x Δ**T**

=> (50 +25) x 4.2 x 6.5 = **2047.5J**oules

Moles potassium hydroxide =$\frac{\text{Enthalpy change } \Delta H}{\Delta H_n}$

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$$\frac{2047.5\text{Joules}}{56700\text{Joules}} = \mathbf{0.0361\ moles}$$

Molarity of KOH = $\frac{\text{moles} \times 1000}{\text{Volume used}}$ => $\frac{0.0361\ \text{moles} \times 1000}{50\text{cm}^3} = \mathbf{0.722M}$

3.Determine the specific heat capacity of a solution of a solution mixture of 50.0cm3 of 2M potassium hydroxide neutralizing 50.0cm3 of 2M nitric(V) acid if a 13.25°C rise in temperature is recorded.(1mole of potassium hydroxide produce 55.4kJ of energy)

Moles of potassium hydroxide = molarity KOH x volume $\frac{}{1000}$

=> $\frac{2\ \text{M} \times 50\text{cm}^3}{1000} = \mathbf{0.1\ moles}$

Enthalpy change ΔH = ΔH_n x Moles potassium hydroxide

=> 55.4kJ x 0.1 moles = 5.54kJ x 1000=**5540J**oules

Specific heat capacity = $\frac{\text{Enthalpy change } \Delta H \text{ in Joules}}{\text{Mass of base + acid} \times \Delta T}$

=> $\frac{5540}{(50+50) \times 13.25} = \mathbf{4.1811J^{-1}g^{-1}K^{-1}}$

Graphically ΔH_n can be determined as in the example below:

Procedure

Place 8 test tubes in a test tube rack .Put 5cm3 of 2M sodium hydroxide solution into each test tube.

Measure 25cm3 of 1M hydrochloric acid into 100cm3 plastic beaker.

Record its initial temperature at volume of base =0. Put one portion of the base into the beaker containing the acid.

Stir carefully with the thermometer and record the highest temperature change to the nearest 0.5°C.

Repeat the procedure above with other portions of the base to complete table 1 below

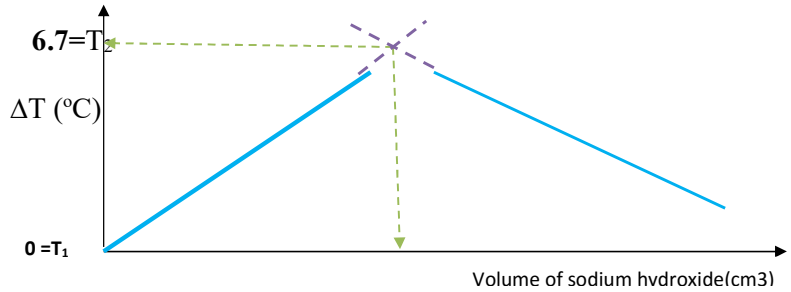
Table 1:Sample results.

olume of acid(cm3)	25.0	25.0	25.0	25.0	25.0	25.0	25.0	25.0	25.0
Volume of alkali(cm3)	0	5.0	10.0	15.0	20.0	25.0	30.0	35.0	40.0
Final temperature(°C)	**22.0**	**24.0**	**26.0**	**28.0**	**28.0**	**27.0**	**26.0**	**25.0**	**24.0**
Initial temperature(°C)	**22.0**	**22.0**	**22.0**	**22.0**	**22.0**	**22.0**	**22.0**	**22.0**	**22.0**
Change in temperature	**0.0**	**2.0**	**4.0**	**6.0**	**6.0**	**5.0**	**4.0**	**3.0**	**2.0**

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(a)Complete the table to determine the change in temperature.
(b)Plot a graph of volume of sodium hydroxide against temperature change.



From the graph show and determine :

(i)the highest temperature change ΔT

$\Delta T = T_2 - T_1$ => highest temperature-T_2 (from extrapolating a correctly plotted graph) less lowest temperature at volume of base=0 :T_1

=>$\Delta T = 6.7 - 0.0 =$ **6.7^0C**

(ii)the volume of sodium hydroxide used for complete neutralization

From a correctly plotted graph – **16.75cm3**

(c)Calculate the number of moles of the alkali used

Moles NaOH = molarity x volume => 2M x 16.75cm3 = **0.0335 moles**
1000 1000

(d)Calculate ΔH for the reaction

ΔH = mass of solution(acid+base) x c x ΔT

=>(25.0 + 16.75) x 4.2 x 6.7 = 1174.845 J = **1.174845kJ**
1000

(e)Calculate the molar enthalpy of neutralization of the alkali.

$\Delta H_n = \frac{\Delta H_n}{\text{Number of moles}} = \frac{1.174845\text{kJ}}{0.0335} =$ **35.0701kJ**

(d)Standard enthalpy/heat of solution ΔH^{θ}_s

The standard enthalpy of solution $\Delta H^{\theta}_s$is defined as the energy change when one mole of a substance is dissolve in excess distilled water to form an infinite dilute solution. An infinite dilute solution is one which is **too** dilute to be diluted further. Dissolving a solid involves two processes:

(i) **breaking** the **crystal** of the solid into **free ions**(cations and anion).This process is the **opposite** of the **formation** of the crystal itself. The energy required to form one mole of a crystal structure from its **gaseous ions** is called **Lattice energy**/heat/enthalpy of lattice ($\Delta \mathbf{H_l}$). Lattice energy /heat/enthalpy of lattice (ΔH_l) is an endothermic process ($+\Delta \mathbf{Hl}$).

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The table below shows some ΔH_l in kJ for the process MX(s) -> M^+ (g) + X^- (g)

	Li	Na	K	Ca	Mg
F	+1022	+900	+800	+760	+631
Cl	+846	+771	+690	+2237	+2493
Br	+800	+733	+670	+2173	+2226

(ii)**surrounding** the free ions by polar **water** molecules. This process is called **hydration.** The energy produced when one mole of ions are completely hydrated is called **hydration energy**/ heat/enthalpy of hydration(ΔH_h).Hydration energy /enthalpy of hydration(ΔH_h) is an **exothermic** process(ΔH_h).

The table below shows some ΔH_h in kJ for some ions;

ion	Li^+	Na^+	K^+	Mg^{2+}	Ca^{2+}	F^-	Cl^-	Br^-
ΔH_h	-1091	-406	-322	-1920	-1650	-506	-364	-335

The **sum** of the **lattice** energy $+\Delta H_l$ (endothermic) and **hydration** energy $-\Delta H_h$ (exothermic) gives the heat of solution$-\Delta H_s$

$\Delta H_s = \Delta H_l + \Delta H_h$

Note

Since ΔH_l is an endothermic process and ΔH_h is an exothermic process then ΔH_s is**:**

(i)**exothermic** if ΔH_l is **less** than ΔH_h and hence a solid **dissolve** easily in water.

(ii)**endothermic** if ΔH_l is **more** than ΔH_h and hence a solid does not **dissolve** easily in water.

(a)Dissolving sodium chloride crystal/s:

(i) **NaCl** ----breaking the crystal into free ions---> **Na^+(g)+ Cl^-(g)** ΔH_l **=+771 kJ**

(ii) Hydrating the ions;

$Na^+(g)$ + **aq** -> Na(aq) **ΔHh = - 406 kJ**

$Cl^-(g)$ + aq -> $Cl^-(aq)$ **ΔHh = - 364 kJ**

$\Delta H_s = \Delta H_h + \Delta H_s$ -> (- 406 kJ + - 364 kJ) + +771 kJ = **+ 1.0 kJmole^{-1}**

NaCl does not **dissolve** easily in water because overall ΔH_s is **endothermic**

Solubility of NaCl therefore increases with increase in temperature.

Increase in temperature increases the energy to break the crystal lattice of NaCl to free Na^+(g)+ Cl^-(g)

(b)Dissolving magnesium chloride crystal/s// $MgCl_2$ (s) ->$MgCl_2$ (aq)

(i) $MgCl_2$ --breaking the crystal into free ions-->Mg^{2+}(g)+ $2Cl^-$(g) ΔH_l =+2493 kJ

(ii) Hydrating the ions;

$Mg^{2+}(g)$ + aq -> Mg^{2+}(g) (aq) ΔHh = - 1920 kJ

$2Cl^-(g)$ + aq -> $2Cl^-(aq)$ ΔHh = (- 364 x 2) kJ

$\Delta H_s = \Delta H_h + \Delta H_s$ -> (- 1920 kJ + (- 364 x 2 kJ)) + +2493 kJ = **-155.0 kJmole^{-1}**

$MgCl_2$ (s) **dissolve** easily in water because overall ΔH_s is **exothermic** .

Solubility of $MgCl_2$ (s) therefore decreases with increase in temperature.

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(c)Dissolving Calcium floride crystal/s// CaF_2 (s) -> CaF_2 (aq)

(i) **CaF_2 -->**$Ca^{2+}(g)+ 2F^-(g)$ ΔH_l =+760 kJ

(ii) Hydrating the ions;

$Ca^{2+}(g)$ + aq -> $Ca^{2+}(g)$ (aq) ΔHh = - 1650 kJ

$2F^-(g)$ + aq -> $2F^-(aq)$ ΔHh = (- 506 x 2) kJ

$\Delta H_s = \Delta H_h + \Delta H_s$ -> (- 1650 kJ + (- 506 x 2 kJ)) + +760 kJ = **-1902.0 kJmole^{-1}**

CaF_2 (s) **dissolve** easily in water because overall ΔH_s is **exothermic** .

Solubility of CaF_2 (s) therefore decreases with increase in temperature.

(d)Dissolving magnesium bromide crystal/s// $MgBr_2$ (s) ->$MgBr_2$ (aq)

(i) $MgCl_2$ --breaking the crystal into free ions-->$Mg^{2+}(g)+ 2Br^-(g)$ ΔH_l =+2226 kJ

(ii) Hydrating the ions;

$Mg^{2+}(g)$ + aq -> $Mg^{2+}(g)$ (aq) ΔHh = - 1920 kJ

$2Br^-(g)$ + aq -> $2Br^-(aq)$ ΔHh = (- 335x 2) kJ

$\Delta H_s = \Delta H_h + \Delta H_s$ -> (- 1920 kJ + (- 335 x 2 kJ)) + +2226 kJ = **-364.0 kJmole^{-1}**

$MgBr_2$ (s) **dissolve** easily in water because overall ΔH_s is **exothermic** .

Solubility of $MgBr_2$(s) therefore decreases with increase in temperature.

Practically the heat of solution can be determined from dissolving known amount /mass/volume of solute in known mass /volume of water/solvent.

From the temperature of solvent **before** and **after** dissolving the change in temperature(ΔT) during dissolution is determined.

To determine the ΔH_s ammonium nitrate

Place 100cm3 of distilled water into a plastic beaker/calorimeter. Determine its temperature and record it at time =0 in table I below.

Put all the 5.0g of ammonium nitrate (potassium nitrate/ammonium chloride can also be used)provided into the plastic beaker/calorimeter, stir using a thermometer and record the highest temperature change to the nearest 0.5^{o}Cafter every ½ minute to complete table I.

Continue stirring the mixture throughout the experiment.

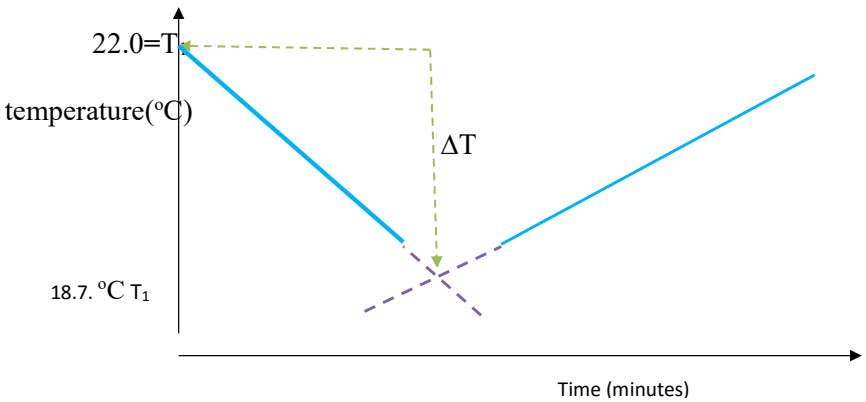
Sample results: Table I

Time (minutes)	0.0	½	1	1 ½	2	2 ½	3	3 ½
Temperature()^{o}C	**22.0**	**21.0**	**20.0**	**19.0**	**19.0**	**19.5**	**20.0**	**20.5**

(a)Plot a graph of temperature against time(x-axis)

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(b)From the graph show and determine the highest temperature change ΔT

$\Delta T = T_2 - T_1$ => lowest temperature-T_2 (from extrapolating a correctly plotted graph) less highest temperature at volume of base=0 :T_1

=>$\Delta T = 18.7 - 22.0 =$ **3.3^0C**

(c)Calculate the number of moles of ammonium nitrate(V) used

Moles NH_4NO_3 = mass used => 5.0 = **0.0625 moles**
Molar mass 80

(d)Calculate ΔH for the reaction

ΔH = mass of water x c x ΔT

->100 x 4.2 x 3.3 = +1386 J = **+1.386kJ**
1000

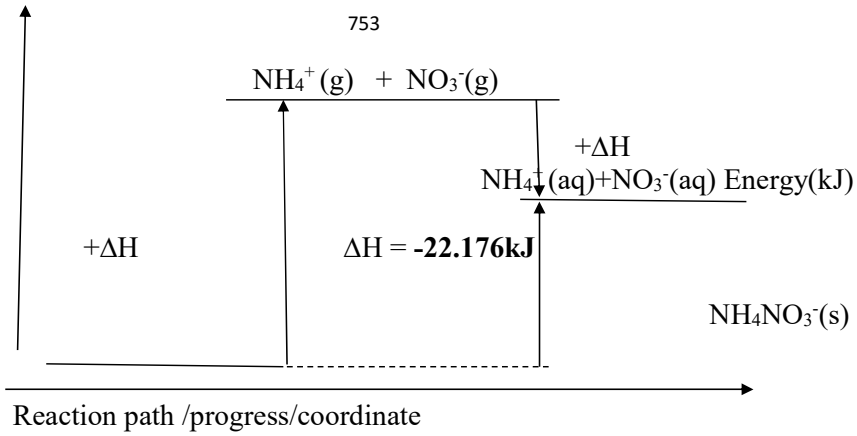
(e)Calculate the molar enthalpy of dissolution of ammonium nitrate(V).

ΔH_s = ΔH = +1.386kJ = + 22.176kJ $mole^{-1}$
Number of moles 0.0625 moles

(f)What would happen if the distilled water was heated before the experiment was performed.

The ammonium nitrate(V)would take less time to dissolves. Increase in temperature reduces lattice energy causing endothermic dissolution to be faster

(g)Illustrate the process above in an energy level diagram



(h) 100cm3 of distilled water at 25°C was added carefully 3cm3 concentrated sulphuric(VI)acid of density 1.84gcm^{-3}.The temperature of the mixture rose from 25^{0}C to 38°C.Calculate the molar heat of solution of sulphuric(VI)acid (S=32.0,H=1.0,0=16.0)

Working

Molar mass of H_2SO_4 = 98g

Mass of H_2SO_4= Density x volume => 1.84gcm^{-3} x 3cm3 = **5.52 g**

Mass of H_2O = Density x volume => 1.00gcm^{-3} x 100cm3 = **100 g**

Moles of H_2SO_4= $\frac{\text{mass}}{\text{Molar mass of } H_2SO_4}$ => $\frac{5.52\text{ g}}{98\text{g}}$ = **0.0563 moles**

Enthalpy change ΔH= (mass of acid + water) x specific heat capacity of water x ΔT

=> (100 +5.52 g) x 4.2 x 13°C = $\frac{5761.392\text{ J}}{1000}$ = **5.761392 kJ**

ΔH_s of H_2SO_4= $\frac{\Delta H}{\text{Moles of } H_2SO_4}$ => $\frac{5.761392\text{ kJ}}{0.0563\text{ moles}}$ = **-102.33378kJmoles^{-1}**

(e)Standard enthalpy/heat of formation ΔH^{θ}_f

The molar enthalpy of formation ΔH^{θ}_f is defined as the energy change when one mole of a compound is formed from its elements at 298K(25°C) and 101325Pa(one atmosphere)pressure. ΔH^{θ}_f is practically difficult to determine in a school laboratory.

It is determined normally determined by applying Hess' law of constant heat summation.

Hess' law of constant heat summation states that **"the total enthalpy/heat/energy change of a reaction is the same regardless of the route taken from reactants to products at the same temperature and pressure".**

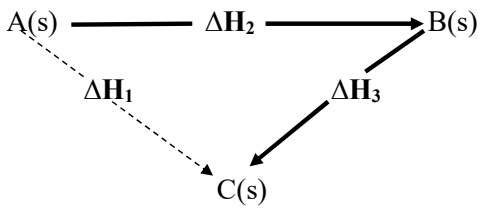
Hess' law of constant heat summation is as a result of a series of experiments done by the German Scientist Henri Hess(1802-1850).

He found that the total energy change from the reactants to products was the same irrespective of the intermediate products between. i.e.

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$A(s) \xrightarrow{-\Delta H_1} C(s) = A(s) \xrightarrow{-\Delta H_2} B(s) \xrightarrow{-\Delta H_3} C(s)$

Applying Hess' law of constant heat summation then:



The above is called an **energy cycle diagram**. It can be used to calculate any of the missing energy changes since:

(i) $\Delta H_1 = \Delta H_2 + \Delta H_3$

(ii) $\Delta H_2 = \Delta H_1 + -\Delta H_3$

(iii) $\Delta H_3 = -\Delta H_1 + \Delta H_2$

Examples of applying Hess' law of constant heat summation

1.Calculate the molar enthalpy of formation of methane (CH_4) given that ΔH^{θ}_c of carbon-graphite is -393.5kJmole^{-1},Hydrogen is -285.7 kJmole^{-1} and that of methane is 890 kJmole^{-1}

Working

Carbon-graphite ,hydrogen and oxygen can react to first form methane.

Methane will then burn in the oxygen present to form carbon(IV)oxide and water.

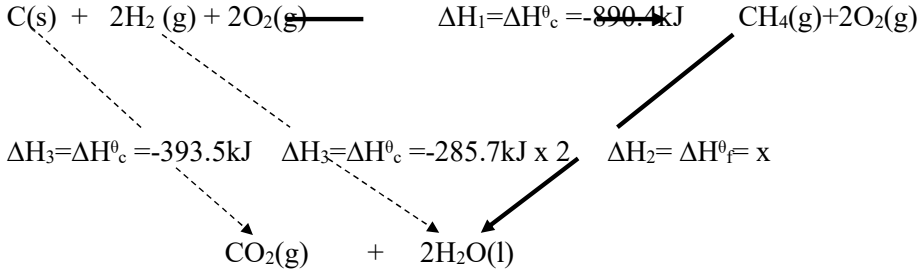
Carbon-graphite can burn in the oxygen to form carbon(IV)oxide.

Hydrogen can burn in the oxygen to form water.

$C(s) + 2H_2(g) + 2O_2(g) \xrightarrow{-\Delta H_1} CH_4(g) + 2O_2(g) \xrightarrow{-\Delta H_2} \mathbf{CO_2(g) + 2H_2O(l)}$

$C(s) + 2H_2(g) + 2O_2(g) \xrightarrow{-\Delta H_3} \mathbf{CO_2(g) + 2H_2O(l)}$

Energy cycle diagram



Substituting:

$\Delta H_3 = \Delta H_1 + \Delta H_2$

$-393.5 + (-285.7 \times 2) = -890.4\text{kJ} + x$

$x = -74.5$ kJ

Heat of formation ΔH^{θ}_f $CH_4 = -74.5$ kJmole^{-1}

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2. Calculate the molar enthalpy of formation of ethyne (C_2H_2) given : ΔH^{θ}_c of carbon-graphite = -394kJmole^{-1},Hydrogen = -286 kJmole^{-1} , (C_2H_2) = -1300 kJmole^{-1}

Working

Carbon-graphite ,hydrogen and oxygen can react to first form ethyne.

Ethyne will then burn in the oxygen present to form carbon(IV)oxide and water.

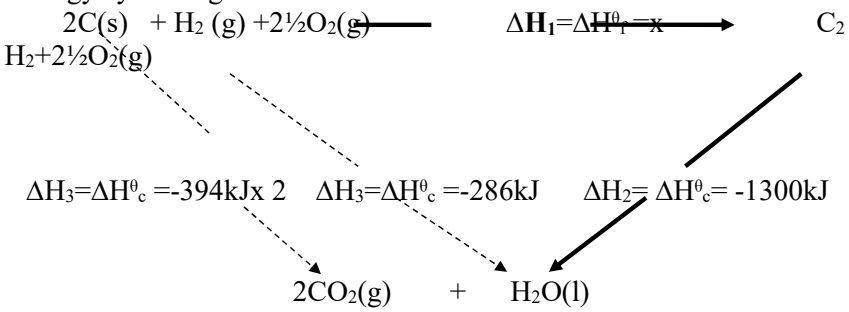
Carbon-graphite can burn in the oxygen to form carbon(IV)oxide.

Hydrogen can burn in the oxygen to form water.

$2C(s)+ H_2(g)+2\ ½\ O_2(g)$ --ΔH_1--> $C_2H_2(g) +2\ ½\ O_2(g)$ --ΔH_2--> **$CO_2(g)+H_2O(l)$**

$2C(s)+ H_2(g)+ 2\ ½\ O_2(g)$ --ΔH_3--> **$2CO_2(g)+H_2O(l)$**

Energy cycle diagram



Substituting:

$\Delta H_3 = \Delta H_1 + \Delta H_2$

$(-394 \times 2) + -286 = -1300kJ + x$

$x = +244\ kJ$

Heat of formation $\Delta H^{\theta}_f\ CH_4$ = **+244 kJmole^{-1}**

3. Calculate the molar enthalpy of formation of carbon(II)oxide (CO) given : ΔH^{θ}_c of carbon-graphite = -393.5kJmole^{-1}, ΔH^{θ}_c of carbon(II)oxide (CO)= -283 kJmole^{-1}

Working

Carbon-graphite reacts with oxygen first to form carbon (II)oxide (CO).

Carbon(II)oxide (CO) then burn in the excess oxygen to form carbon(IV)oxide.

Carbon-graphite can burn in excess oxygen to form carbon (IV) oxide.

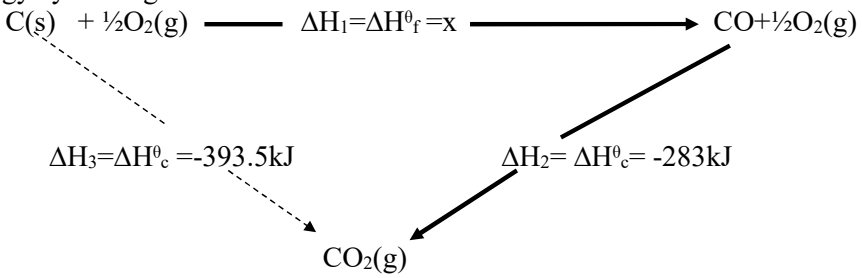
$C(s)+\ ½O_2(g)$ --ΔH_1--> $CO(g) + ½\ O_2(g)$ --ΔH_2--> **$CO_2(g)$**

$C(s)+\ O_2(g)$ --ΔH_3--> **$CO_2(g)$**

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Energy cycle diagram



Substituting:

$\Delta H_3 = \Delta H_1 + \Delta H_2$

$-393.5kJ = -283kJ + x$

$x = -110\ kJ$

Heat of formation ΔH^{θ}_f CO = -110 kJmole^{-1}

4.Study the information below:

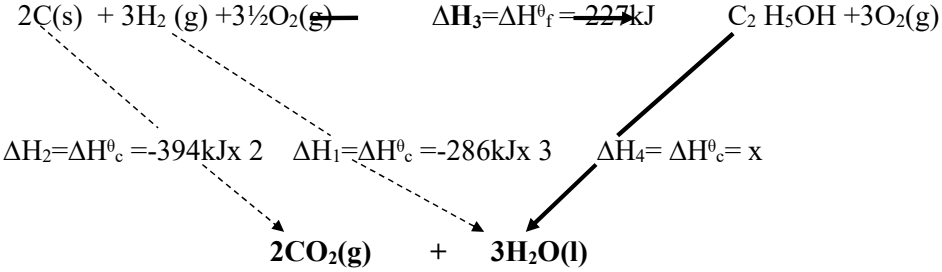
$H_2(g) + ½\ O_2(g)$ -> $H_2O(l)$ ΔH_1=-286 kJmole^{-1}

$C(s) + O_2(g)$ -> $CO_2(g)$ ΔH_2=-393 kJmole^{-1}

$2C(s) + H_2(g) + ½\ O_2(g)$ ->$C_2H_5OH(l)$ ΔH_3=-277 kJmole^{-1}

Use the information to calculate the molar enthalpy of combustion ΔH_4 of ethanol

Energy cycle diagram



Substituting:

$\Delta H_1 + \Delta H_2 = \Delta H_3 + \Delta H_4$

$(-394 \times 2) + -286 \times 3 = -277 + x$

$\Delta H_4 = -1369\ kJ$

Heat of combustion ΔH^{θ}_c C_2H_5OH = **-1369 kJmole^{-1}**

5.Given the following information below:

$CuSO_4(s) + (aq)$ -> $CuSO_4(aq)$ ΔH=-66.1 kJmole^{-1}

$CuSO_4(s) + (aq) + 5H_2O(l)$-> $CuSO_4.5H_2O$ (aq) ΔH=-77.4 kJmole^{-1}

Calculate ΔH for the reaction;

$CuSO_4(aq) + 5H_2O$-> $CuSO_4.5H_2O$ (aq) ΔH=-77.4 kJmole^{-1}

Working

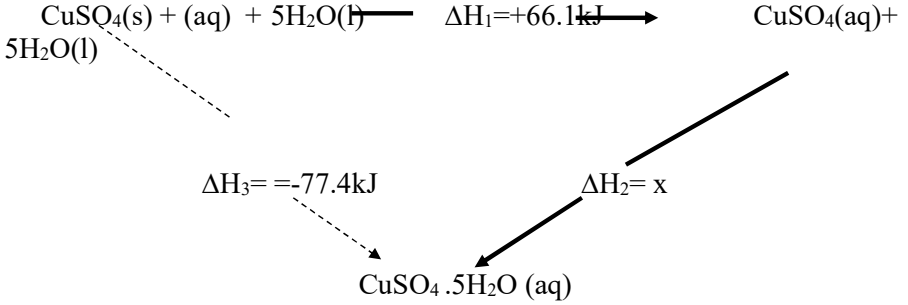
$CuSO_4(s) + (aq) + 5H_2O(l)$-> $CuSO_4(aq)$+ $5H_2O(l)$-> $CuSO_4.5H_2O$ (aq)

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$CuSO_4(s) + (aq) + 5H_2O(l) \rightarrow CuSO_4.5H_2O\ (aq)$

Energy cycle diagram



Substituting:

$\Delta H_3 = \Delta H_2 + \Delta H_1$

$(-77.4kJ = x + +66.1kJ$

$\Delta H_4 =$ **-10.9 kJ**

Heat of dissolution of $CuSO_4$ = **$-10.9kJmole^{-1}$**

Practically, Hess' law can be applied practically as in the following examples

a)Practical example 1

Determination of the enthalpy of formation of $CuSO_4.5H_2O$

Experiment I

Weigh accurately 12.5 g of copper(II)sulphate(VI)pentahydrate. Measure 100cm3 of distilled water into a beaker. Determine its temperature T_1 .Put all the crystals of the copper(II)sulphate(VI)pentahydrate carefully into the beaker. Stir using a thermometer and determine the highest temperature change T_2 Repeat the procedure again to complete table 1.

Table 1:Sample results

Experiment	I	II
Highest /lowest temperature T_2	27.0	29.0
Initial temperature T_1	24.0	25.0
Change in temperature ΔT	3.0	4.0

Experiment II

Weigh accurately 8.0g of anhydrous copper(II)sulphate(VI). Measure 100cm3 of distilled water into a beaker. Determine its temperature T_1 .Put all the crystals of the copper(II)sulphate(VI)pentahydrate carefully into the beaker. Stir using a thermometer and determine the highest temperature change T_2 Repeat the procedure again to complete table II.

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Table II :Sample results

Experiment	I	II
Highest /lowest temperature T_2	26.0	27.0
Initial temperature T_1	25.0	25.0
Change in temperature ΔT	1.0	2.0

Questions

(a)Calculate the average ΔT in

(i)Table I

$\Delta T = T_2 - T_1 \Rightarrow \frac{3.0 + 4.0}{2} =$ **3.5 °C**

(ii)Table II

$\Delta T = T_2 - T_1 \Rightarrow \frac{1.0 + 2.0}{2} =$ **1.5 °C**

(b)Calculate the number of moles of solid used in:

(i)Experiment I

Moles of $CuSO_4.5H_2O$ = $\frac{\text{Mass}}{\text{Molar mass}} \Rightarrow \frac{12.5}{250} =$ **0.05 moles**

(ii)Experiment II

Moles of $CuSO_4$ = $\frac{\text{Mass}}{\text{Molar mass}} \Rightarrow \frac{8.0}{160} =$ **0.05 moles**

(c)Calculate the enthalpy change for the reaction in:

(i)Experiment I

Enthalpy change of $CuSO_4.5H_2O$= mass of Water(**m**) x **c** x Δ**T**

$\Rightarrow \frac{100cm3}{1000} \times 4.2 \times 3.5\,^{o}C =$ **-1.47kJ**

(ii)Experiment II

Enthalpy change of $CuSO_4$ = mass of water(**m**) x **c** x Δ**T**

$\Rightarrow \frac{100cm3}{1000} \times 4.2 \times 1.5\,^{o}C =$ **-0.63kJ**

(c)Calculate the molar enthalpy of solution $CuSO_4.5H_2O$ (s) form the results in (i)experiment I.

$\Delta Hs = CuSO_4.5H_2O = \frac{\Delta H}{\text{Number of Moles}} \Rightarrow \frac{-1.47kJ}{0.05 \text{ moles}} =$ **29.4kJ**

(ii)experiment II.

$\Delta Hs = CuSO_4 = \frac{\Delta H}{\text{Number of Moles}} \Rightarrow \frac{-0.63kJ}{0.05 \text{ moles}} =$ **12.6kJ**

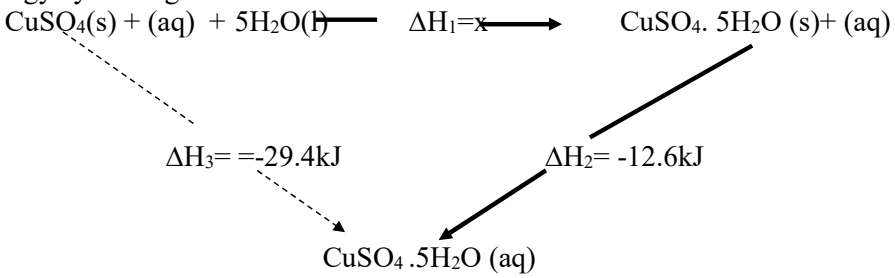
(d) Using an energy level diagram, calculate the molar enthalpy change for the reaction:

$CuSO_4.5H_2O$ (s) -> $CuSO_4$(s) + $5H_2O$(l)

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Energy cycle diagram



$\Delta H_3 = \Delta H_1 + \Delta H_2$

$\Rightarrow -29.4kJ = -12.6kJ + x$

$\Rightarrow -29.4kJ - (+12.6kJ) = x$

x = 16.8kJ

b)Practical example II

Determination of enthalpy of solution of ammonium chloride

Theoretical information.

Ammonium chloride dissolves in water to form ammonium chloride solution. Aqueous ammonia can react with excess dilute hydrochloric acid to form ammonium chloride solution. The heat change taking place can be calculated from the heat of reactions:

(i) $NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(s)$

(ii) $NH_4Cl(s) + (aq) \rightarrow NH_4Cl(aq)$

(iii) $NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$

Experiment procedure I

Measure 50cm3 of water into a 100cm3 beaker. Record its temperature T_1 as initial temperature to the nearest 0.5°C in table I. Add exactly 5.0g of ammonium chloride crystals weighed carefully into the water. Stir and record the highest temperature change T_2 as the final temperature change. Repeat the above procedure to complete table I.

Sample results TableI

Experiment	I	II
final temperature(°C)	19.0	20.0
initial temperature(°C)	22.0	22.0
temperature change ΔT(°C)	3.0	2.0

Experiment procedure II

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Measure 25cm3 of 2M aqueous ammonia into a 100cm3 beaker. Record its temperature T_1 as initial temperature to the nearest 0.5°C in table II. Measure 25cm3 of 2M hydrochloric acid solution. Add the acid into the beaker containing aqueous ammonia. Stir and record the highest temperature change T_2 as the final temperature change. Repeat the above procedure to complete table II.

Sample results:Table II

Experiment	I	II
final temperature(°C)	29.0	29.0
initial temperature(°C)	22.0	22.0
temperature change ΔT(°C)	7.0	7.0

Sample Calculations:

(a)Calculate the average ΔT in

(i)Table I

$\Delta T = T_2 - T_1 \Rightarrow \frac{-3.0 + -2.0}{2} =$ **2.5 °C**

(ii)Table II

$\Delta T = T_2 - T_1 \Rightarrow \frac{7.0 + 7.0}{2} =$ **7.0 °C**

(b)Calculate the enthalpy change for the reaction in:

(i)Experiment I

Enthalpy change ΔH = mass of Water(**m**) x **c** x **ΔT**

$\Rightarrow \frac{50cm3 \times 4.2 \times 2.5\,°C}{1000} =$ **+0.525kJ**

(ii)Experiment II

Enthalpy change of $CuSO_4$ = mass of water(**m**) x **c** x **ΔT**

$\Rightarrow \frac{25+25cm3 \times 4.2 \times 7\,°C}{1000} =$ **+1.47kJ**

(c)Write the equation for the reaction taking place in:

(i)Experiment I

$NH_4Cl(s) + (aq) \rightarrow NH_4Cl(aq)$

(ii)Experiment I

$NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$

(d)Calculate the enthalpy change ΔH for the reaction:

$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$ given that:

(i) $NH_3(g) + (aq) \rightarrow NH_3(aq)$ ΔH= -40.3kJ

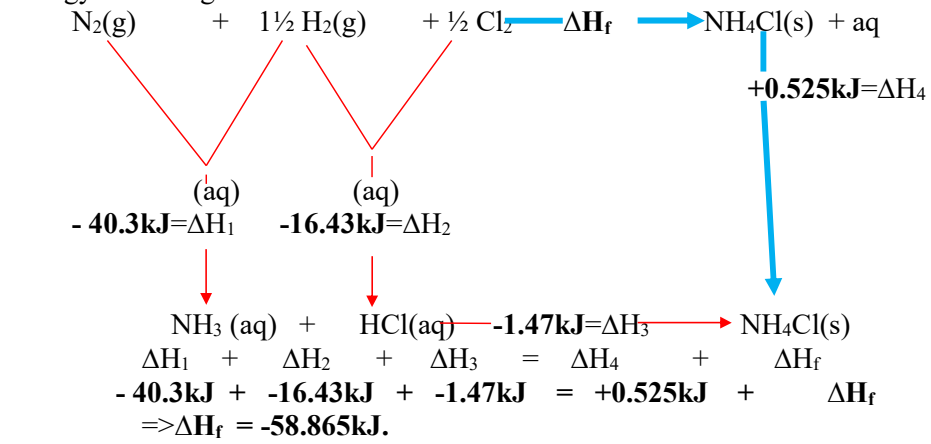
(ii) $(aq) + HCl(g) \rightarrow HCl(aq)$ ΔH= -16.45kJ

(e)Applying Hess' Law of constant heat summation:

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Energy level diagram



$\Delta H_1 + \Delta H_2 + \Delta H_3 = \Delta H_4 + \Delta H_f$

- 40.3kJ + -16.43kJ + -1.47kJ = +0.525kJ + ΔH_f

$\Rightarrow \Delta H_f$ = -58.865kJ.

Practice theoretical examples:

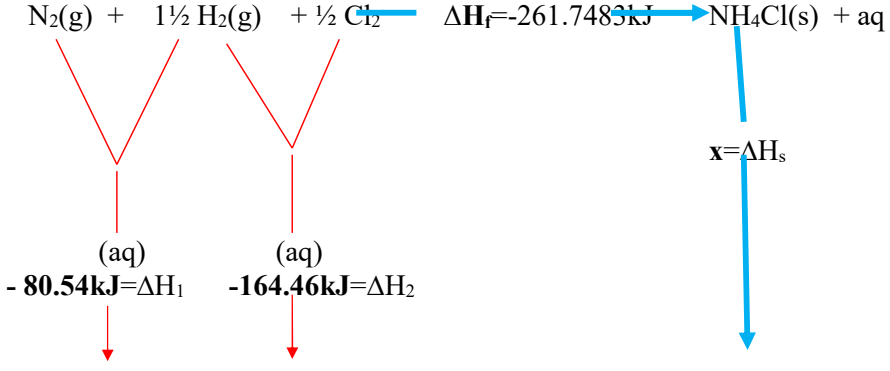
1. Using an energy level diagram calculate the ΔH_s of ammonium chloride crystals given that.

ΔH_f of NH_3 (aq) = -80.54kJ mole^{-1}

ΔH_f of HCl (aq) = -164.46kJ mole^{-1}

ΔH_f of NH_4Cl (aq) = -261.7483kJ mole^{-1}

ΔH_s of NH_4Cl (aq) = -16.8517kJ mole^{-1}



$\Delta H_1 + \Delta H_2 + \Delta H_3 = \Delta H_4 + \Delta H_f$

- 80.54kJ + -164.46kJ + -16.8517kJ = -261.7483kJ + ΔH_f

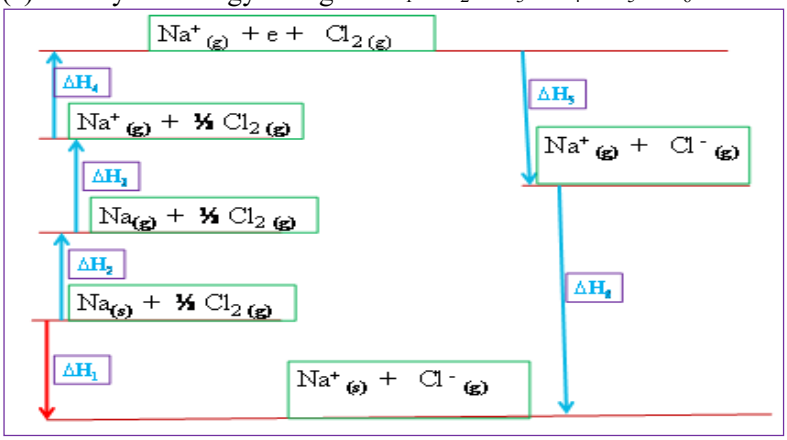
$\Rightarrow \Delta H_f$ = **-33.6kJmole^{-1}**.

Study the energy cycle diagram below and use it to:

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(a)Identify the energy changes ΔH_1 ΔH_2 ΔH_3 ΔH_4 ΔH_5 ΔH_6



ΔH_1 - enthalpy/heat of formation of sodium chloride (ΔH_f)
ΔH_2 -enthalpy/heat of atomization of sodium (ΔH_{at})
ΔH_3 -enthalpy/heat of ionization/ionization energy of sodium (ΔH_i)
ΔH_4 -enthalpy/heat of atomization of chlorine (ΔH_{at})
ΔH_5 -enthalpy/heat of electron affinity of chlorine (ΔH_e)
ΔH_6 enthalpy/heat of lattice/Lattice energy of sodium chloride(ΔH_l)

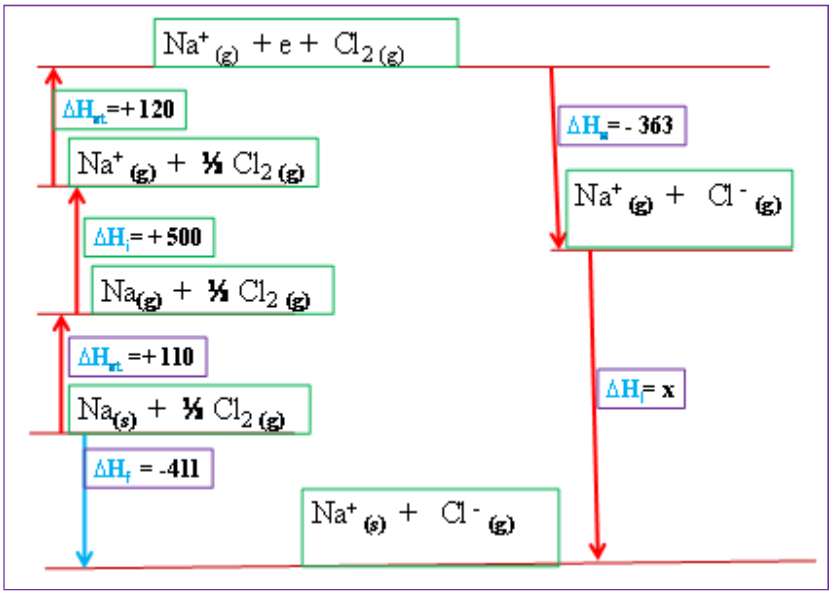
(b) Calculate ΔH_1 given that ΔH_2 =+108kJ , ΔH_3=+500kJ, ΔH_4 =+121kJ ,ΔH_5 =-364kJ and ΔH_6 =-766kJ

Working:
$\Delta H_1 = \Delta H_2 + \Delta H_3 + \Delta H_4 + \Delta H_5 + \Delta H_6$
Substituting:
ΔH_1= +108kJ + +500kJ + +121kJ +-364kJ + -766kJ
ΔH_1= **-401kJmole^{-1}**

(c) Given the that:
(i) Ionization energy of sodium = + 500kJmole^{-1}
(ii)ΔH_{at} of sodium = + 110kJmole^{-1}
(iii) Electron affinity of chlorine = - 363kJmole^{-1}
(iv)ΔH_{at} of chlorine = + 120kJmole^{-1}
(v) ΔH_f of sodium chloride = -411kJ , calculate the lattice energy of sodium chloride using an energy cycle diagram.

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Working:
Applying Hess law then:
$\Delta H_f = \Delta H_a + \Delta H_i + \Delta H_a + \Delta H_e + \Delta H_l$
Substituting:
-411= +108kJ + +500kJ + +121kJ +-364kJ + x
-411 + -108kJ + -500kJ + -121kJ + +364kJ = x
x= **-776kJmole^{-1}**

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