

## Chemistry

Chemistry is the study of matter and its properties. Chemistry is a practical subject and therefore it is studied normally in a special place called the laboratory


### Aims of studying chemistry

- To acquire observational skills.
- To acquire the skills of accurately reporting what has been observed.
- To acquire the skills of handling instruments.(apparatus)
- To acquire the skills of predicting the outcome of an experiment.
- Chemistry also help us to make new products e.g. drugs, plastics etc

Chemistry is studied in a special place called the laboratory and when a student is in the laboratory, the following simple rules must be followed.

- Never enter the laboratory when a teacher or laboratory technician is absent.
- Never eat or taste any thing when in the laboratory.
- Never mix any chemical when you are in the laboratory unless when told by the teacher.
- Never clean or wash glass apparatus with water when they are still hot to avoid cracking.
- Incase of any accident, report immediately to the teacher or laboratory assistant.
- Never add water to acid but add acid to water. This is to avoid generating a lot of heat.
- Never heat chemical more than the time told by the teacher.

### Common laboratory apparatus

Apparatus	Use
<b>Test tube</b> 	<ul style="list-style-type: none"><li>• Used to heat or warm chemicals during qualitative analysis.</li></ul>

## Bunsen burner

A Bunsen burner is used in the laboratory to produce flame required for heating. A flame is only produced when a Bunsen burner is connected to a gas cylinder containing gas that can burn when lit. The Bunsen burner consists of the following parts:

- i) Chimney
- ii) Air hole
- iii) Collar/metal ring. This controls the size of the air hole.
- iv) The base. It holds the Bunsen burner firmly on the table
- v) Pin hole jet. It allows gas into the chimney.
- vi) Side tube. It connects the Bunsen burner through a gas tube to the tap which is on the table.

The following precautions must be followed when lighting a Bunsen burner:

- 1) Close the air hole.
- 2) Turn the gas on using the gas tap.
- 3) Light the gas immediately at the top of the chimney.
- 4) Adjust the size of the hole until the flame is about 7 cm high.
- 5) Open the air hole slowly until the flame is non-luminous but not noisy.

## Burning back

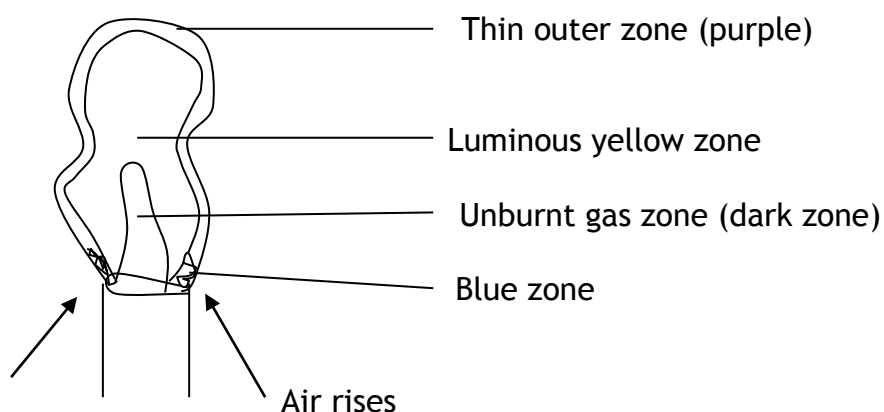
This normally occurs when the air holes of a Bunsen burner are open and the gas is turned down such that the rate at which the gas is burnt exceeds the rate at which it is supplied from the jet. During this process, the flame moves down the chimney and is seen to come from the jet at the base of the Bunsen burner.

When a flame burns back, turn off the gas, close the air hole and light it again using the normal procedure.

## Flame

A flame is burning gas that gives out heat and light. There are two types of flames formed depending on whether the air holes are opened or closed. Namely

- a) **Luminous flame.** It is formed when the air holes of Bunsen burner are closed such that no air enters the tube. This flame is large and bright. It has four zones as illustrated below.

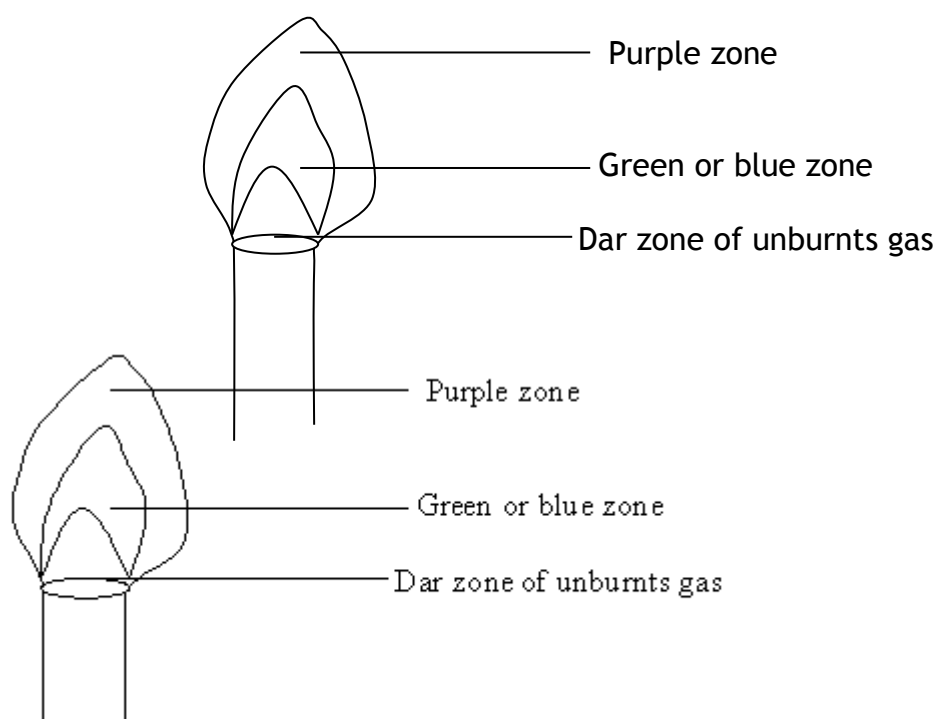


- The luminous yellow zone. The gas in this zone does not burn completely because there is not enough air. It forms about 70% of the flame.
- The thin outer zone. In this zone, the gas burns completely as there is enough air.
- The unburnt gas zone. This zone is dark and cool.
- The blue zone. This is found at the bottom of the flame and it receives plenty of air. Therefore, the gas in this zone burns completely making it very hot.

### Characteristics of a luminous flame

- It is not very hot
- It is yellow and unsteady.
- It is sooty or smoky.
- It has four zones.
- It does not burn back.
- The flame is easily visible.
- The flame is big in size.

b) **Non-luminous flame.** It is formed when the air hole of a Bunsen burner is opened such that air enters the tube and mixes with the gas that burns quickly and completely to produce a lot of heat. The flame has three zones as illustrated in figure below



- The dark zone of the unburnt gas. This consists of a mixture of gas and air that comes out of the chimney.
- The green or blue zone. In this zone there is a partial or incomplete combustion because of limited supply air.

- The pale blue or purple zone. In this zone, there is complete combustion and a lot of heat is produced.

### Characteristics of non-luminous flame

- i) It is very hot
- ii) It is blue and steady
- iii) It is non-sooty
- iv) It is noisy
- v) It has three zones
- vi) It usually burns back
- vii) It is small in size
- viii) It is not easily visible.

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### Differences between luminous flame and non-luminous flame

Luminous flame	Non-luminous flame
<ul style="list-style-type: none"> <li>• It is not very hot</li> <li>• It is yellow and unsteady</li> <li>• It is sooty or smoky</li> <li>• It has four zones</li> <li>• It does not burn back</li> <li>• Flame is easily seen</li> <li>• Flame is big in size</li> <li>• It burns silently</li> </ul>	<ul style="list-style-type: none"> <li>• Very hot</li> <li>• Blue and steady</li> <li>• Non-sooty</li> <li>• It has three zones</li> <li>• It burns back</li> <li>• Flame is not easily visible</li> <li>• It is small in size</li> <li>• It is noisy</li> </ul>

## MATTER

Matter is any thing that occupies space and has weight. There are three states of matter. Namely solid, liquid and gas.

### Kinetic theory of matter

This theory explains the three states of matter. It states that all matter is made up of particles and these particles possess kinetic energy which cause them to be in a state of constant motion.

### Properties of solids

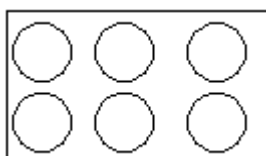
- i) The particles in solids are closely packed and orderly arranged.



- ii) The particles in the solids are held together by a strong force of attraction which makes it difficult for the particles to move freely. For this reason, solids are hard and strong.
- iii) The particles in solids can only vibrate within a fixed position making solids to have definite shape and volume.
- iv) Solids have high melting and boiling points.

### Properties of liquids

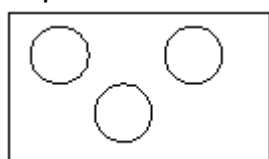
- i) The particles in liquids are still closely packed but less closely packed in solids.



- ii) The particles have less force of attraction than in solids. For this reason, they can move from one point to another freely.
- iii) They do not have definite shape and volume but they can take the shape of the container in which they are put.

### Properties of gases

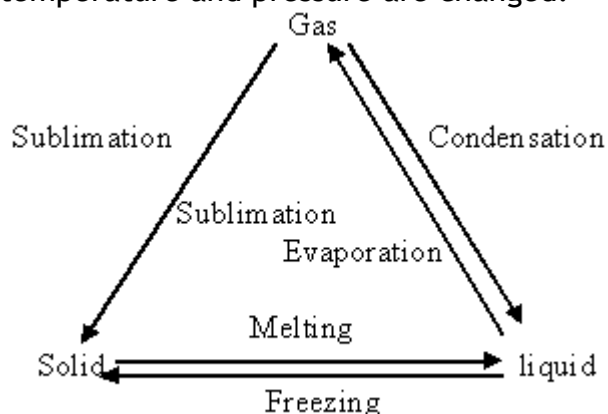
- i) The gas particles are widely separated with almost no force of attraction between the particles.



- ii) The particles are in state of constant motion (random movement in all directions)
- iii) They have definite volume and shape but fill up the volume of the container in which they are put.

### Interconversion of matter

The three states of matter can be converted into the other when the conditions of temperature and pressure are changed.



In summary;

- Change of state from solid to gas or gas to solid is called sublimation.
- Change of state from solid to liquid is called melting.
- Change of state from liquid to solid is called freezing.
- Change of state from liquid to gas is called evaporation.
- Change of state from gas to liquid is called condensation.

## Testing for purity of a substance

A pure substance does not contain impurities. The following criteria are used to test for purity of a given substance in the laboratory.

- Boiling point
- Melting point
- Density

## Boiling point

Boiling point of a liquid is the constant temperature at which a liquid changes completely into vapour. Boiling takes place when the vapour pressure exerted by the liquid is equal to the external atmospheric pressure.

A pure liquid will always boil at a constant temperature for example pure water boils at  $100^{\circ}\text{C}$  at 760 mmHg and ethanol boils at  $78^{\circ}\text{C}$  at 760 mmHg.

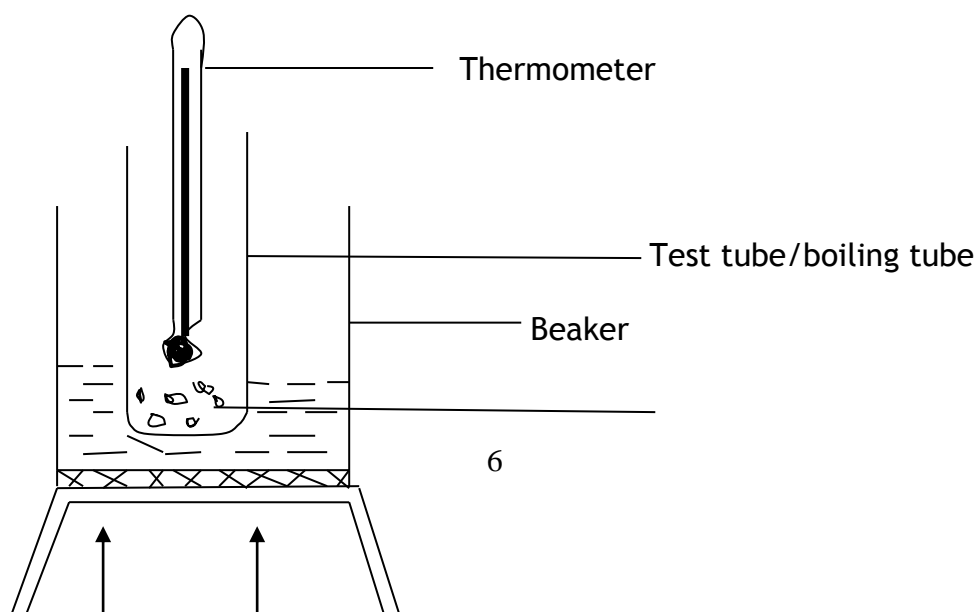
The boiling point of a pure substance is higher than expected in the presence of impurities.

## Experiment to determine the boiling of a liquid

### Requirements

- Thermometer
- Heat source
- Beaker
- Water
- Test tube/boiling tube
- Sample to be investigated

### Experimental set up



be determined

Liquid whose boiling point is to



## Procedure

- i) The liquid sample is put in test tube/boiling tube with a thermometer held slightly above the liquid.
- ii) The test tube/boiling tube together with its contents is transferred to boiling water in a beaker.
- iii) The temperature of the vapour when it becomes constant is noted. This is the boiling point of the liquid.

**NB//.** The thermometer is held above the liquid to avoid super heating i.e. higher boiling point than expected can be achieved if the thermometer is dipped into the liquid sample.

## Melting point

Melting point is the constant temperature at which a pure substance turns from solid state to liquid state at standard temperature and pressure.

The melting point of a pure substance is normally sharp while that of impure substance is not sharp because an impure substance first softens then later turns into liquid but a pure substance will just melt.

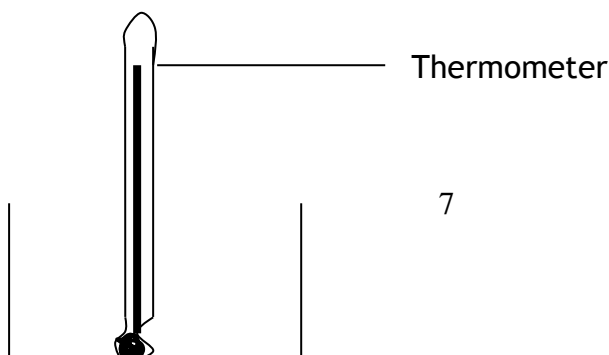
The presence of impurities tends to lower the melting point of a substance for example the melting point of ice water is  $0^{\circ}\text{C}$  but in presence of sodium chloride as impurity, it melts at  $-27^{\circ}\text{C}$ .

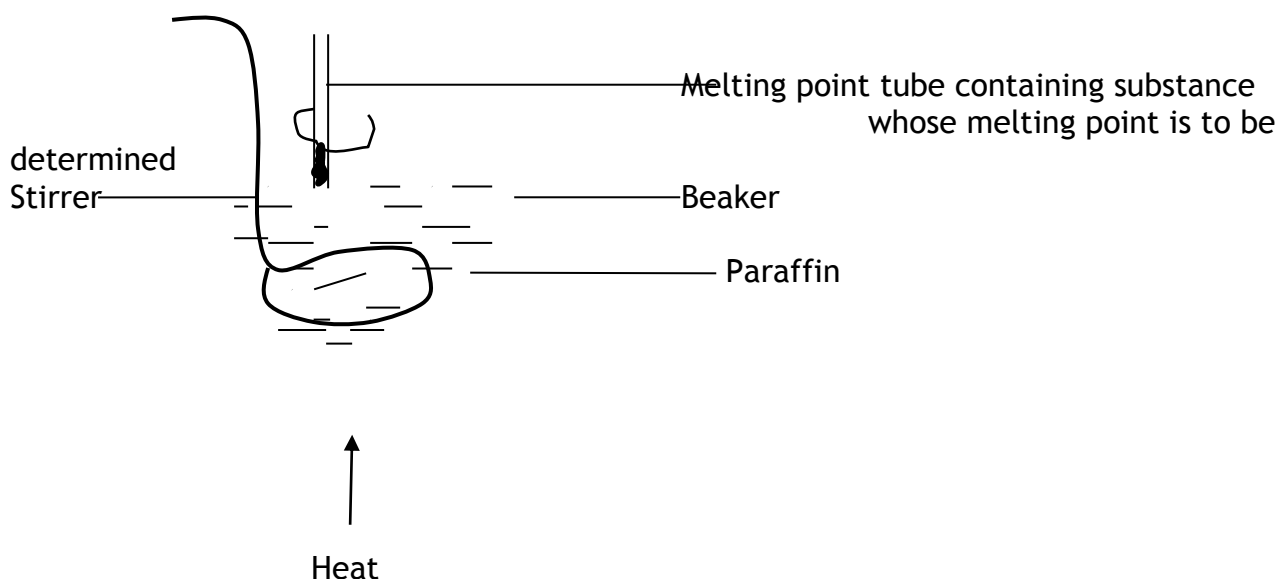
## Experiment to determine melting point of a substance

### Requirements

- Thermometer
- Paraffin
- Heat source
- Beaker
- Melting point tube/capillary tube
- Solid sample to be investigated

### Experimental set up





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## Procedure

- i) The solid sample is put in thin walled capillary tube sealed at one end.
- ii) The capillary tube with its content tight to a thermometer and put in beaker containing paraffin.
- iii) The paraffin is heated with constant stirring until the solid just melts. This temperature is noted and it is the melting point of the solid sample.

NB//. Stirring is done to ensure uniform distribution of heat so that melting can take place at once.

## Physical and chemical change

Matter undergoes several changes classified as physical or chemical changes.

- 1) **Physical or non-permanent changes.** These occur when no new chemical substances are formed. Examples include
  - (a) Burning of candle wax
  - (b) Melting of ice
  - (c) Sublimation iodine crystals. When iodine crystals are heated, they change into vapour which condenses to form iodine crystals
  - (d) Sublimation of ammonium chloride
- 2) **Chemical or permanent changes.** These occur when new chemical substances are formed. Examples include
  - (a) Burning of wood, papers or grass
  - (b) Heating of copper (II) sulphate crystals
  - (c) Heating of potassium permanganate crystals

NB//. Chemical changes are accompanied by change in colour.

## Differences between chemical change and physical change

Chemical change	Physical change
<ul style="list-style-type: none"> <li>▪ A new chemical substance is formed</li> <li>▪ The change is irreversible (permanent)</li> </ul>	<ul style="list-style-type: none"> <li>▪ No new chemical substance is formed</li> <li>▪ The change is reversible (non-</li> </ul>



<ul style="list-style-type: none"> <li>▪ The mass of the original substance changes</li> <li>▪ Energy is either absorbed or given out</li> </ul>	permanent) <ul style="list-style-type: none"> <li>▪ The mass of the original substance does not change</li> <li>▪ Energy is neither absorbed nor given out</li> </ul>
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## Elements, mixtures and compounds

### Elements

These are substances that cannot be split into simpler substances by chemical means. There are about 104 different types of elements. These elements are further classified as

- Metals e.g. sodium, calcium, magnesium, aluminium etc
- Non-metals e.g. sulphur, carbon, nitrogen etc

### Characteristics of metals

- Metals are generally hard compared to non-metals
- They have high melting and boiling points
- Metals conduct electricity and heat e.g. iron, copper, aluminum etc
- Some metals are magnetic e.g. iron
- Metals are malleable and ductile i.e. they can be bent without necessarily breaking.
- Metals exist as solids except mercury which exists as a liquid

### Characteristics of non-metals

- They have low melting and boiling points
- They are brittle
- They exist as soft solid, liquids and gases at room temperature
- They are insulators i.e. they do not conduct electricity and heat
- They are non-magnetic

An element is normally represented in chemistry by a symbol.

A symbol of an element is one or two alphabetical letters that represents one atom of that element.

### Compound

A compound is a substance which consists of two or more elements chemically combined together e.g.

Compound	Elements
<ul style="list-style-type: none"> <li>• Water</li> <li>• Carbon dioxide</li> <li>• Alcohol</li> </ul>	<ul style="list-style-type: none"> <li>• Hydrogen and oxygen</li> <li>• Carbon and oxygen</li> <li>• Carbon, hydrogen and oxygen</li> </ul>

### Mixture

This is a substance which consists of two or more elements or compounds not chemically combined together e.g.

Mixture	Elements or compounds
<ul style="list-style-type: none"> <li>• Air</li> <li>• Steel</li> <li>• Sea water</li> </ul>	<ul style="list-style-type: none"> <li>• Nitrogen, carbondioxide and rare gases</li> <li>• Iron and carbon</li> <li>• Salt, water and soil</li> </ul>

## Atom

Atom is the smallest indivisible particle of an element that can take part in a chemical change.

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## Molecule

A molecule is the smallest particle of an element or compound which exists freely. Generally atoms combine form elements and elements combine to form molecules or compounds.

## Differences between compounds and mixtures

Compounds	Mixtures
<ul style="list-style-type: none"> <li>• The elements in a compound cannot be separated by physical means</li> <li>• Its properties are different from elements in it</li> <li>• Energy is usually given out or absorbed</li> <li>• Composition is not variable</li> </ul>	<ul style="list-style-type: none"> <li>• The substances in a mixture can be separated by physical means</li> <li>• Properties of a mixture are quite average of the substances in it</li> <li>• Energy is not usually given out or absorbed</li> <li>• Composition of a mixture is variable</li> </ul>

## Separation of mixtures

Mixtures are separated into their pure components by physical processes.

Separation of mixtures is important because;

- It enables purification of a substance
- It enables isolation/extraction of a substance
- It enables identification of a substance

There are five physical processes involved in the separation of mixtures

- Distillation
- Decantation
- Magnetic separation
- Filtration
- Chromatography
- Sublimation
- Crystallization

## Types of mixtures

There are four types of mixtures

- 1) Liquid mixture
  - (a) Miscible liquid mixtures e.g. ethanol and water
  - (b) Immiscible liquid mixtures e.g. paraffin and water
- 2) Gas mixtures e.g. air is made up of oxygen, carbon dioxide and nitrogen
- 3) Solid mixtures e.g. iron and sulphur
- 4) Solid-liquid mixtures e.g. water and sand

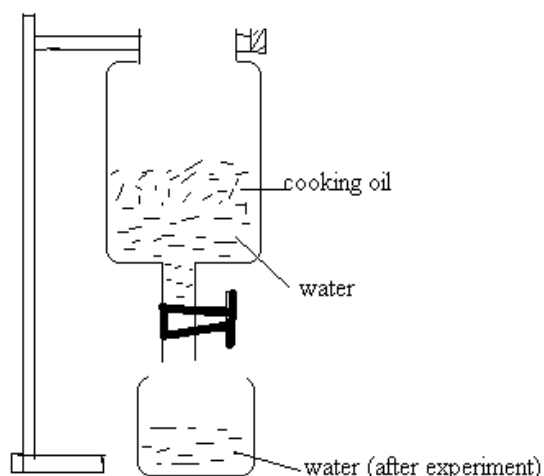
## Separation of liquid mixtures

### (a) Immiscible liquids

These are two or more liquids which do not form a uniform solution when mixed together instead they form separate layers. Examples include

- Paraffin and water
- Cooking oil and water
- Petrol and water etc

These liquids can be separated by using a separating funnel as shown below



### Procedure

- A mixture of cooking oil and water is poured into a separating funnel and allowed to settle.
- Cooking oil will float on top of water since it is less dense than water.
- Immediately the tap of the separating funnel is opened. The water runs down where it is collected in a beaker.
- The tap is closed when the two liquids have fully been separated.

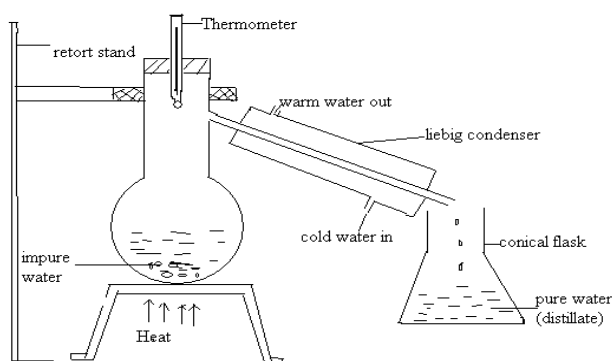
### (b) Miscible liquids

These are two or more liquids that form a uniform solution when put together for e.g. water and ethanol, petrol and diesel, petrol and paraffin, diesel and cooking oil, impure water. Etc

They can be separated using a process called distillation. Distillation is classified into two types;

### i) Simple distillation

This is the vapourisation followed by condensation of only one liquid. It is normally used during the purification of impure water. Below is an experimental set up for the process of simple distillation



### Procedure

- The impure water containing the sand, soil and salt is put in a distilling flask as shown in the set up above.
- The water is boiled using a Bunsen burner to allow the water vapour to rise.
- The vapour is cooled and condensed by cold water entering the Liebig condenser.
- The condensed vapour forms a liquid called distillate which is collected in a conical flask.
- The material which remains in the distilling flask is the residue.

**NB//.** Broken porcelain is put at the bottom of the distilling flask to allow even boiling of the impure water.

### ii) Fractional distillation

This is the vaporisation followed by condensation of more than liquid. These liquids have close boiling points examples of mixtures that can be separated using this method include;

### Separation of solid-liquid mixtures

The methods used to separate solid liquid mixtures include;

- Crystallization
- Filtration
- Decantation

Example's of mixtures which can be separated by the above methods

- ✓ Sand mixed with salt solution
- ✓ Sand and water

✓ Chalk dust mixed with water

- a) **Crystallization.** This is a process of recovering a salt from a solution for example a mixture sand and common salt dissolved in water.

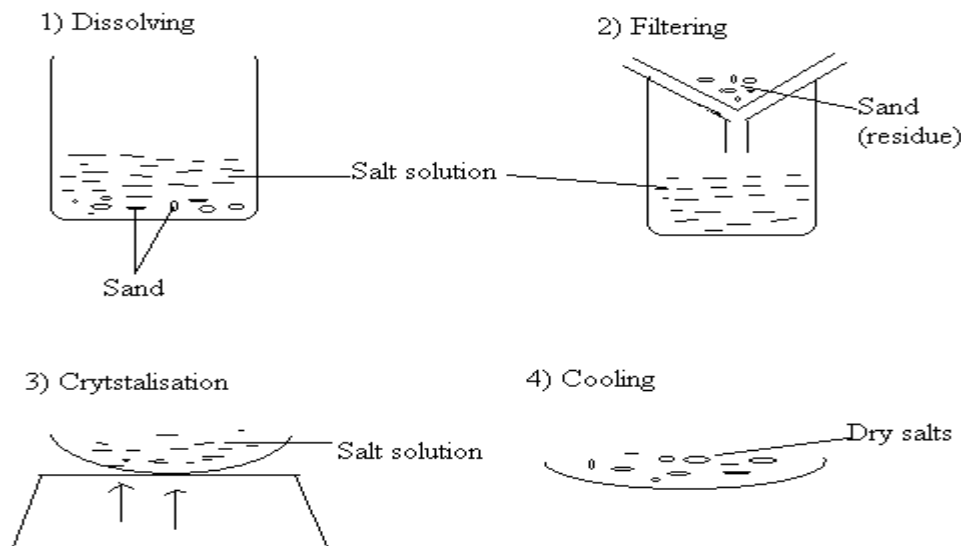
## Separation of a mixture of sand and salt

### Procedure

- Water is added to a mixture of sand and salt in a beaker and the mixture shaken thoroughly.
- The salt dissolves leaving behind the sand.
- Pour the resultant mixture into a filter funnel containing a filter paper. The salt solution will pass through the filter into a beaker as the sand particles are retained on the filter paper as residue.
- The salt solution is then transferred into an evaporating dish/basin and heated to evaporate water from the solution until the salt crystals begin to form.
- The hot solution with the few crystals is removed from the heat source and allowed to cool so that complete crystallization can take place.
- The crystals formed are filtered from the mother solution and dried between filter papers.

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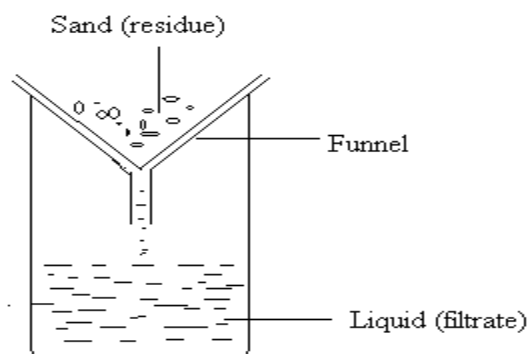
### Illustrations



**NB//.** Direct heating of the evaporating dish is not a good way of removing excess water since it can result in the salt crystals jumping out of the dish.

- b) **Filtration.** This is a process of separating a solid from liquid or solution by use of a filter that holds back the solid but allows the liquid to pass through or trickle through. The liquid that passes through a filter during filtration is called **filtrate** while the solid that remains in the filter is the **residue**.

### Illustration



- c) **Decantation.** This is the separation of solid liquid mixture by allowing the solid to settle at the bottom of the container followed by careful pouring out/off the liquid to another container.

### Illustration

### Separation of solid mixtures

The components present in the solid mixtures determine the method one should use. The commonly employed methods are as follow;

- Sublimation.
- Use of magnet

- a) **Sublimation.** This is the direct change of solid to gas when heated or gas to solid when cooled without passing through the liquid state. For this method to work, one of the components must be able to sublime. Examples of substances that under undergo sublimation include;

- ✓ Ammonium chloride
- ✓ Iodine crystals
- ✓ Iron (III chloride
- ✓ Aluminium chloride

Therefore, any solid mixture with one of the above can be separated by sublimation for example;

- ✓ Iodine crystals and sand
- ✓ Iron (III chloride and sodium chloride
- ✓ Zinc oxide and ammonium chloride

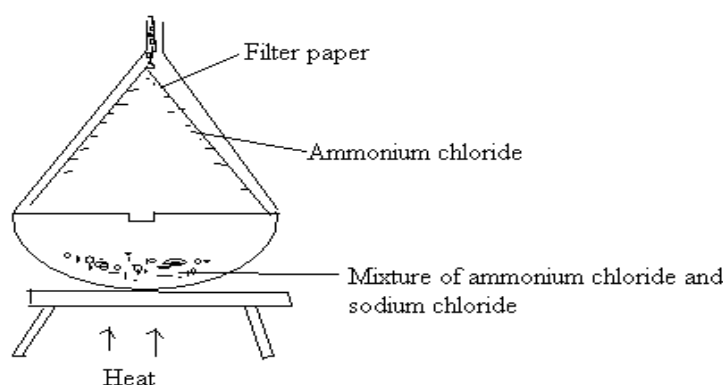
### Separation of a mixture of ammonium chloride and sodium chloride

## Requirements

- Evaporating dish
- Heat source
- Funnel and filter paper
- Mixture of ammonium chloride and sodium chloride

## Experimental set up

### Separation of ammonium chloride from sodium chloride



## Procedure

- A mixture of ammonium chloride and sodium chloride is poured into an evaporating dish and covered with a filter paper having holes.
  - An inverted funnel is placed on the filter paper.
  - The mixture is heated, first gently and then strongly to drive off the ammonium chloride.
  - The ammonium chloride vapour passes through the holes in the filter paper and condenses on cool parts of the funnel.
- b) **Use of magnet.** The use of a magnet is possible if one of the components in the mixture is diamagnetic i.e. has the ability to be attracted by a magnet on close contact. Below are some examples of diamagnetic substances;
- ✓ Iron fillings
  - ✓ Cobalt
  - ✓ Copper etc

The following are examples of solid mixtures that can be separated by use of magnet

- ✓ Sulphur and iron fillings
- ✓ Copper and zinc oxide
- ✓ Cobalt and copper (II) sulphate crystals
- ✓ Sugar and iron fillings

## Separation of iron-sulphur mixture by use of a bar magnet

- A bar magnet is dipped into the iron-sulphur mixture in a container and rubbed thoroughly throughout the mixture.
- The iron fillings are attracted to the magnet and pulled out leaving behind the sulphur.

## Chromatography

This is the process of separation of miscible liquids by using their different rates of movement over a porous medium. This method is heavily employed in the separation of;

- ✓ Ink components
- ✓ Plant pigments
- ✓ Dyes
- ✓ Paints
- ✓ Amino acids etc

There three common types of chromatography used in the separation of liquid mixtures are;

- Paper chromatography
- Thin layer chromatography
- Column chromatography

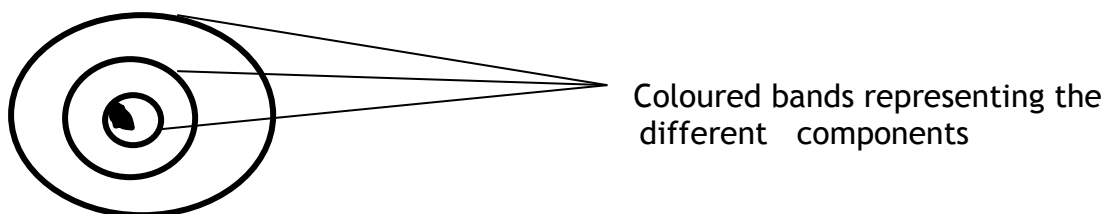
### Paper chromatography

In this type of chromatography, a filter is used as the porous medium over which the different components move.

This chromatography is mainly used to separate the different components in ink.

### Procedure

- A small drop of ink is placed at the center of a filter paper and allowed to dry.
- Several drops of ink are then added using a pipette to the dry ink spot and allowed to spread throughout.
- A number of different substances separate out into a series of bands. Each band represents a component separated from the rest.

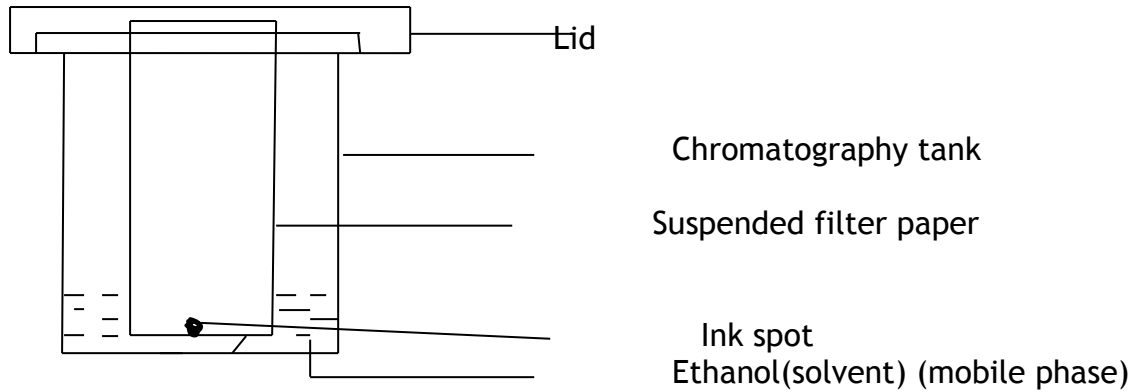


In chromatography, there are two phases;

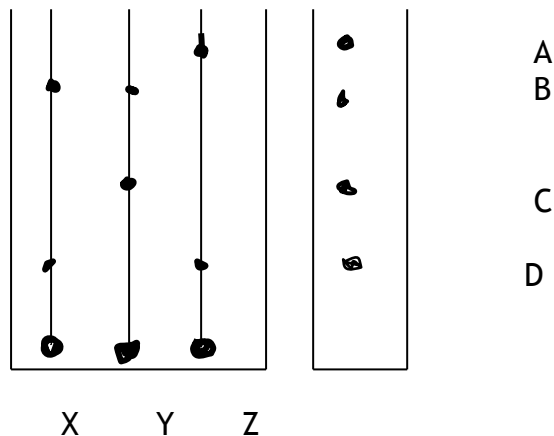
- Stationary phase. This is the filter filter
- Mobile phase. This is the solvent in which the different components can dissolve.

Paper chromatography can also be demonstrated a filter paper placed vertically in a chromatography tank.





### Specimen result



According to the diagram above, there are four different components which make up mixtures X, Y and Z.

- Mixture X is made up of components B and D
- Mixture Y is made up of components B and C
- Mixture Z is made up of components D and A

### Importance of chromatography

- It is used in the purification of drugs
- It is used in the purification of food
- It used to separate pigments from plants

### Crystal

A crystal is a solid which consists of particles arranged in an orderly and repetitive manner. Crystals have the following characteristics;

- They have definite shape
- They have flat sides
- They have sharp edges

## Formation of crystals from a solution

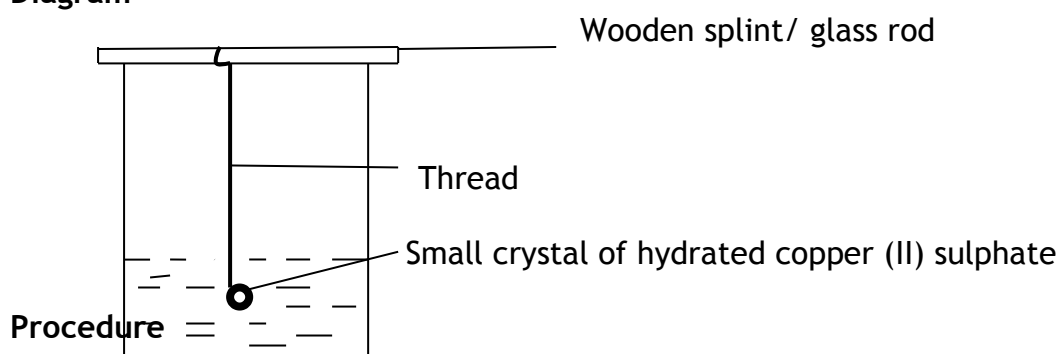
Crystals are formed from both cold and hot solutions. When a hot solution cools, crystals are normally formed. The liquid left when crystals are formed from a hot solution is called mother liquor.

**Experiment to how a large crystal of hydrated copper (II) sulphate is grown.**

### Requirements

- ✓ Beaker
- ✓ Glass rod
- ✓ Thread
- ✓ Hydrated copper (II) sulphate crystals

### Diagram



- Dissolve some crystals of copper (II) sulphate in some water in a beaker to form a saturated solution.
- Choose one good crystal of hydrated copper (II) sulphate and tie a thread around it.
- Hang the crystal in the saturated solution of copper (II) sulphate.
- Leave the crystal in the solution for several weeks in the solution at constant temperature.
- The crystal will continue to grow as the water evaporates from the beaker.

**NB//.** The experiment can also be done with chromium potassium sulphate 12-water (chrome alum)

### Water of crystallization

This is the definite amount of water with which some substances chemically combine when they form crystals from their solutions. A crystal cannot be formed without water. The presence of water in the crystal gives the crystal its **colour** and **shape**.

When the water is removed from the crystal by simply heating, the crystal will lose its colour and shape.

Crystals or substances which contain water of crystallization are hydrates and are said to be **hydrated**.

Crystals without water of crystallization are **anhydrous**.

Examples of hydrated substances which are hydrated and anhydrous

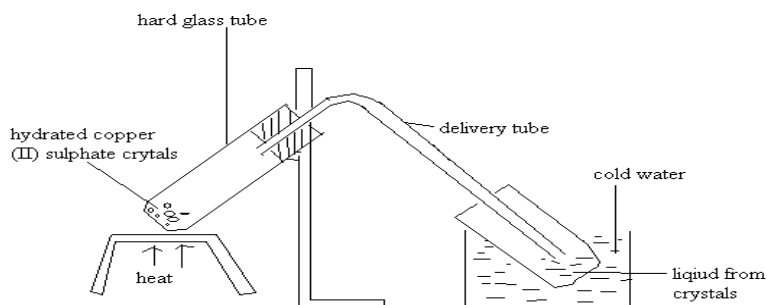
Hydrated substances	Anhydrous substances
<ul style="list-style-type: none"> <li>✓ Sodium carbonate 10-water</li> <li>✓ Copper (II) sulphate -water</li> <li>✓ Sodium sulphate 10-water</li> <li>✓ Iron (II) sulphate 7-water</li> <li>✓ Magnesium sulphate 7-water</li> <li>✓ Calcium chloride 6-water</li> </ul>	<ul style="list-style-type: none"> <li>✓ Sodium chloride</li> <li>✓ Potassium chloride</li> <li>✓ Ammonium carbonate</li> <li>✓ Ammonium chloride</li> </ul>

**Experiment to show that hydrated copper (II) sulphate contains water of crystallization**

### Requirements

- ✓ Boiling tubes
- ✓ Heat source
- ✓ Hydrated copper (II) sulphate crystals
- ✓ Delivery tube
- ✓ Beaker having cold water

### Experimental set up



### Procedure

- Put some crystals of hydrated copper (II) sulphate in a boiling tube and slant it as shown above.
- Connect the boiling tube to a delivery tube placed in another boiling tube dipped in a beaker of cold water.
- Heated the crystals in the boiling tube gently.
- The blue colour of the crystals begins to fade and eventually turns white. This is because the water evaporates from the crystals.
- The water vapour is condensed in the test tube as shown above to form a liquid (water).

In the laboratory, the presence of water can be confirmed using the following chemicals;

- ✓ **Anhydrous Cobalt (II) chloride paper.** Water turns this paper from blue to pink.
  - ✓ **Anhydrous copper (II) sulphate.** Water turns this chemical from white to blue.
- NB//.** Water is neutral to litmus paper i.e. it has no effect on it.

## Symbols, valency, chemical formula and equations

### Symbol of elements

The symbol of an element is one or two letters which mean one atom of the element. It is usually the first letter of the English or Latin name of the elements. However, some elements have two letters. Writing the symbol of an element which has two letters, the first letter is written as capital letter and second one as a small letter. The Latin names of some elements are;

Element	Latin name
Potassium	Kalium
Sodium	Natrium
Iron	Ferrum
Lead	Plumbum
Copper	Cuprum
Mercury	Hydrargyrum
Silver	Argentum

### Valency

The valency of an element or radical is the number of hydrogen atoms which combine with or displace one atom of that element or one group of the radical. The valency tells us what number of atoms of one element will combine with a fixed number of atoms of another element to form a molecule of a compound. Elements or radicals combine with others to form compounds.

Valency is defined in terms of hydrogen because hydrogen atom is taken as a standard and has a valency of one.

Element	Symbol	Valency
Hydrogen	H	1
Helium	He	0
Lithium	Li	1
Beryllium	Be	2
Carbon	C	4
Nitrogen	N	3
Oxygen	F	2
Fluorine	Ne	1
Neon	Na	0
Sodium	Mg	1
Magnesium	Al	2
Aluminium	P	3
Phosphorous	S	3, 5
Sulphur	S	2, 6
chlorine	Cl	1
Potassium	K	1

Calcium	Ca	2
Manganese	Mn	2, 3, 4, 5, 6, 7
Iron	Fe	2, 3, 6
Copper	Cu	1, 2
Zinc	Zn	2
Mercury	Hg	1, 2
Gold	Au	3
Barium	Ba	2
Iodine	I	1
Bromine	Br	1

## Radicals

A radical is an atom or group of atoms that exist in several compounds but do not exist on their own. Radicals behave as if though they were atoms of elements by virtue of having valencies.

Radical	Symbol	Valency
Sulphate	$SO_4$	2
Carbonate	$CO_3$	2
Nitrate	$NO_3$	1
Nitrite	$NO_2$	1
Sulphite	$SO_3$	2
Hydrogen sulphite	$HSO_3$	1
Hydrogen sulphate	$HSO_4$	1
Hydrogen carbonate	$HCO_3$	1
Hydroxide	$OH$	1
Ammonium	$NH_4$	1
Phosphate	$PO_4$	3
Chloride	$Cl$	1
Oxide	$O$	2

## Chemical formula

This consists of symbols and whole numbers together denoting one molecule of the element or compound.eg

Chemical formula	One molecule of ;
$O_2$	Oxygen consists of two atoms of oxygen (2O)
$H_2$	Hydrogen consists of two atoms of hydrogen (2H)
$CO_2$	Carbondioxide consists of one atom of carbon and two atoms of oxygen
$NH_3$	Ammonia consists of one atom of nitrogen and three atoms of hydrogen
$H_2O$	Water consists of two atoms of hydrogen and one atom of oxygen

**NB//.** The atoms of gases like oxygen, hydrogen, chlorine, fluorine, bromine, iodine exist in pairs (diatomic)

The formula 2O means two atoms of oxygen not chemically combined.

## Working out chemical formulae

Symbols, radicals and valencies are used when writing the chemical formulae of compounds. The following rules should be followed when chemical formulae of compounds.

- Write the symbols for the elements and radicals.
- Write the valency above and to the right of the symbols and radicals
- Reverse the valencies and write the numbers (other than 1) below and to the right.

### Examples

Write the chemical formulae of the following compounds.

- (a) Sodium chloride
- (b) Sodium sulphate
- (c) Copper (II) chloride
- (d) Carbondioxide
- (e) Water
- (f) Copper (II) carbonate
- (g) Ammonium sulphate
- (h) Copper (II) nitrate
- (i) Iron (II) sulphate
- (j) Iron (III) sulphate
- (k) Potassium sulphate
- (l) Potassium phoaphate
- (m) Sulphuric acid
- (n) Nitric acid
- (o) Hydrogen chloride
- (p) Lead (II) chloride
- (q) Magnesium hydroxide
- (r) Sodium hydroxide
- (s) Calcium hydroxide
- (t) Zinc carbonate
- (u) Ammonia

### Exercise

Write the chemical formulae for the following compounds.

- (a) Zinc oxide
- (b) Sodium hydrogen sulphate
- (c) Magnesium hydrogen carbonate
- (d) Sodium sulphate
- (e) Iron (II) nitrate
- (f) Barium sulphate
- (g) Potassium chloride
- (h) Lead (II) nitrate

(i) Calcium hydrogen sulphate

If an element has variable valencies, the valency of the element in a particular compound is indicated by a roman numeral e.g. copper (II) oxide, copper (I) oxide, iron (II) hydroxide, iron (III) hydroxide.

## Chemical equations

A chemical equation represents a chemical change by means of symbols and formulae i.e. a statement that represents a chemical reaction. Included in a chemical equation are;

- Reactants
- Products
- Proportion of reactants and products
- Direction of reaction
- State symbols. They indicate the states of matter in which the reactants and products are present i.e.

Solid (s)

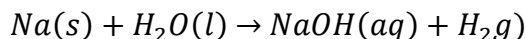
Liquid (l)

Gas (g)

Aqueous (aq) which means the solution contains water.

The state symbols of reactants and products are written to the right of the formulae. The amount of energy evolved or absorbed during a chemical reaction is indicated in most cases. Consider the reaction between sodium and water to form sodium hydroxide and hydrogen gas.

*Sodium + Water → sodium hydroxide + hydrogen gas*

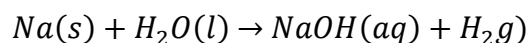


Reactants are always on the left hand side of the equation and products on the right hand side. + means “reacts with”. In the above example, it means that sodium reacts with water to form sodium hydroxide and hydrogen gas.

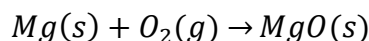
## Word equations

In this case, names of products and reactants are used instead of symbols and formulae.eg

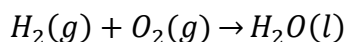
(i) *Sodium + Water → sodium hydroxide + hydrogen gas*



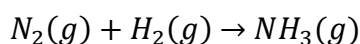
(ii) *Magnesium + oxygen → magnesium oxide*



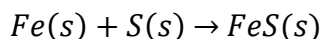
(iii) *Hydrogen + oxygen → water*



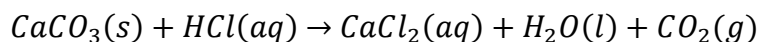
(iv) *Nitrogen + hydrogen → ammonia*



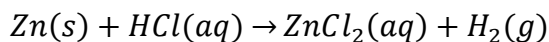
(v) *Iron + sulphur → iron sulphide*



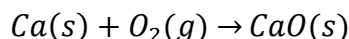
(vi) *Calcium carbonate + hydrochloric acid → calcium chloride + water + carbondioxide*



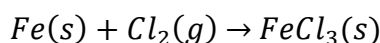
(vii) *Zinc + hydrochloric acid → zinc chloride + hydrogen gas*



(viii) *Calcium + oxygen → calcium oxide*

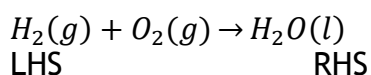


(ix) *iron + chlorine → iron (III)chloride*



### Balancing chemical equations

This is the making of the total number of each kind of atom on the left hand side of the equation equal to the number on the right hand side.eg



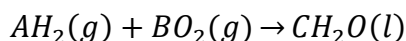
$$H = 2H = 2$$

$$O = 2O = 1$$

Therefore the equation is not balanced because the number of oxygen atoms on the left hand side is not equal to those on the right hand side.

### Rules for balancing molecular chemical equations

- Represent the number of molecules of all the compounds by different letters e.g. A, B, C etc.
- Equate all the number of atoms of the same kind from either sides of the equation.
- Solve the equations for the different letters.



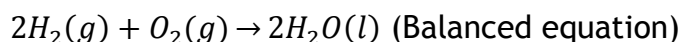
$$H, 2A = 2C$$

$$O, 2B = C$$

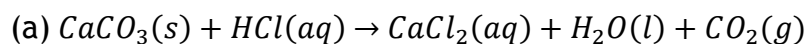
$$\text{Let } A = 1, \text{ then } C = 1 \text{ and } B = \frac{1}{2}$$

Equation becomes  $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$

Multiply the equation through out by 2



### Balance the following equations





- (b)  $\text{Zn}(s) + \text{HCl}(aq) \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g)$
- (c)  $\text{CuSO}_4(aq) + \text{NaOH}(aq) \rightarrow \text{Cu}(\text{OH})_2(s) + \text{Na}_2\text{SO}_4(aq)$
- (d)  $\text{Na}_2\text{O}(s) + \text{H}_2\text{O}(l) \rightarrow \text{NaOH}(aq)$
- (e)  $\text{H}_2\text{O}_2(aq) \rightarrow \text{H}_2\text{O}(l) + \text{O}_2(g)$
- (f)  $\text{MgO}(s) + \text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + \text{H}_2\text{O}(l)$

## Exercise

Balance the following equations

- (a)  $\text{Pb}(\text{NO}_3)_2(s) \rightarrow \text{PbO}(s) + \text{NO}_2(g) + \text{O}_2(g)$
- (b)  $\text{KNO}_3(s) \rightarrow \text{KNO}_2(s) + \text{O}_2(g)$
- (c)  $\text{Fe}(s) + \text{H}_2\text{O}(g) \rightarrow \text{Fe}_3\text{O}_4(s) + \text{H}_2(g)$
- (d)  $\text{Al}(\text{OH})_3(s) \rightarrow \text{Al}_2\text{O}_3(s) + \text{H}_2\text{O}(l)$
- (e)  $\text{Mg}(s) + \text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + \text{H}_2(g)$
- (f)  $\text{H}_2\text{S}(g) + \text{O}_2(g) \rightarrow \text{S}(s) + \text{H}_2\text{O}(l)$

**Note//.** (i) The number below and to the right of the atoms in the bracket only multiplies those atoms

- Aluminium sulphate,  $\text{Al}_2(\text{SO}_4)_3$   
*Number of Al atoms* =  $1 \times 2 = 2$   
*Number of sulphur atoms* =  $1 \times 3 = 3$   
*Number of oxygen atoms* =  $4 \times 3 = 12$
- Magnesium nitrate,  $\text{Mg}(\text{NO}_3)_2$   
*Number of Mg atoms* = 1  
*Number of N atoms* =  $1 \times 2 = 2$   
*Number of O atoms* =  $3 \times 2 = 6$
- Ammonium carbonate,  $(\text{NH}_4)_2\text{CO}_3$   
*Number of N atoms* =  $1 \times 2 = 2$   
*Number of H atoms* =  $4 \times 2 = 8$   
*Number of C atoms* = 1  
*Number of O atoms* =  $1 \times 3 = 3$

(ii) The number in front of the formula multiplies all the atoms in the formula.

- $3\text{Al}_2(\text{SO}_4)_3$  means two molecules of aluminium sulphate  
*Number of Al atoms* =  $1 \times 2 \times 3 = 6$   
*Number of sulphur atoms* =  $1 \times 3 \times 3 = 9$   
*Number of oxygen atoms* =  $4 \times 3 \times 3 = 36$
- $2\text{Cu}(\text{NO}_3)_2$  means two molecules of copper (II) nitrate  
*Number of Cu atoms* =  $1 \times 2 = 2$   
*Number of N atoms* =  $2 \times 1 \times 2 = 4$   
*Number of O atoms* =  $2 \times 3 \times 2 = 12$

## Methods of gas collection

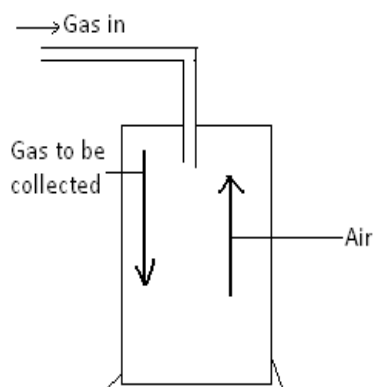
The following apparatus are used to collect gases;

- Gas jar
- Bee-hive stand
- Water trough

In the gas jar, there is air is displaced by the gas to be collected. For this reason, there are two methods of collecting gas depending on the densities of the gases. They include;

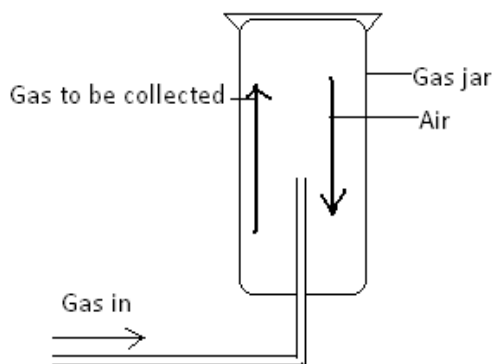
- 1) **Downward delivery (upward displacement of air).** This method is used to collect gases which are denser than air because they would displace air from the gas jar since they are denser than air. Examples of gases collected by this method are;

- Chlorine gas
- Carbon dioxide gas
- Nitrogen dioxide gas
- Sulphur trioxide gas
- Hydrogen chloride gas.
- Sulphur dioxide



- 2) **Upward delivery (downward displacement of air).** This method is used to collect gases which are less dense than air. Examples include;

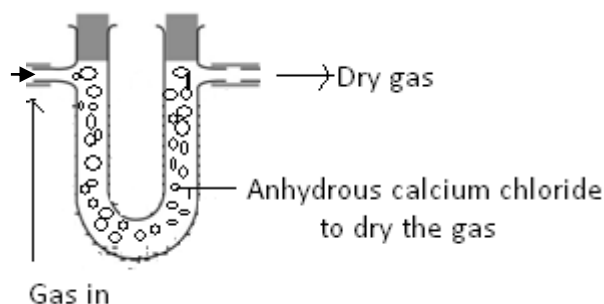
- Oxygen gas
- Hydrogen gas
- Nitrogen gas
- Carbon monoxide gas
- Ammonia gas



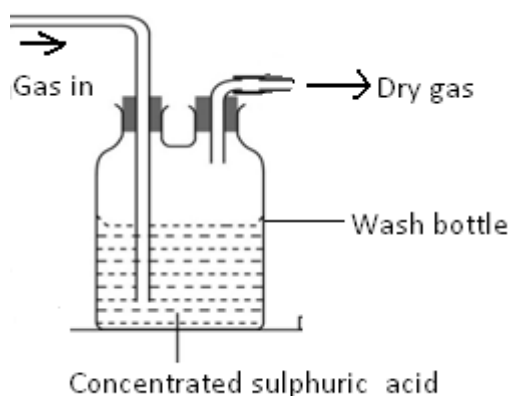
### How to dry a gas

This is the process of removing suspended water vapour molecules from the gas collected. The following chemical can be used to dry gases.

- 1) **Anhydrous calcium chloride.** It is used to dry all gases except ammonia gas because it reacts with it forming a complex. During the process, the anhydrous calcium chloride is put in U-tube and the wet gas is passed through it.



- 2) **Concentrated sulphuric acid.** Concentrated sulphuric acid has no water. This acid is used to dry all gases except ammonia because it reacts with it to form ammonium sulphate. During this process, the acid is put in a wash bottle and the gas is passed through the acid where it is dried.



- 3) **Calcium oxide (quick lime).** It is used to dry ammonia gas because ammonia reacts with other dry agents.

## WATER AND HYDROGEN

### Hydrogen gas

The major source of hydrogen in nature is water. Hydrogen is an element which exists as a gas at room temperature. The symbol is. However, in nature, hydrogen exists as diatomic molecule meaning one molecule contains two atoms.

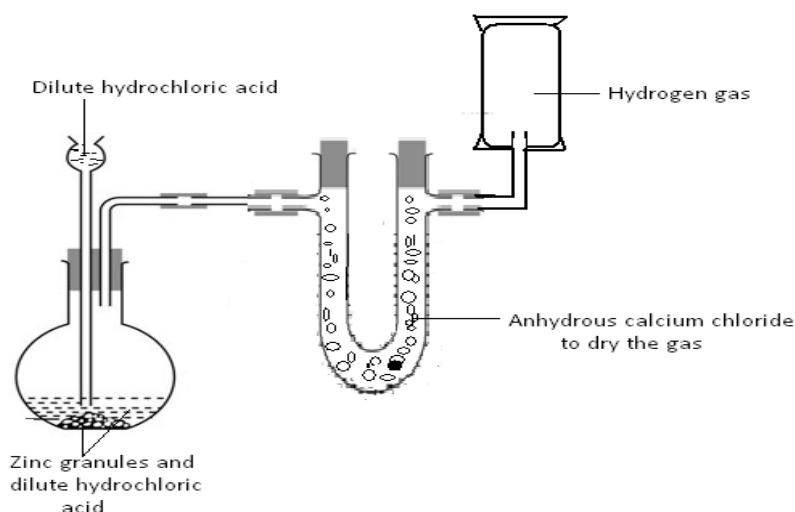
## Laboratory preparation of hydrogen gas

### Requirements

- Zinc granules.
- Dilute hydrochloric acid
- Copper two sulphate
- Flask
- Anhydrous calcium chloride gas jar
- U tube
- Delivery tube.

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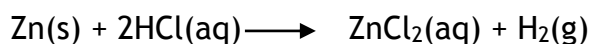
### Experimental set up



### Procedure

- Zinc granules are put in flask fitted with a tap funnel as shown above.
- The whole set up is connected with a u-tube and a gas jar.
- Dilute hydrochloric acid is slowly added to the zinc granules in the flask by opening the tap. immediately, hydrogen gas is produced. This gas is seen as colorless bubbles.
- The gas fills the flask and it's led out of the flask through a delivery tube to a U-tube containing anhydrous calcium chloride which dries the gas.
- This gas is collected by upwards delivery [downward displacement of air] since it's less dense than air.

### Equation



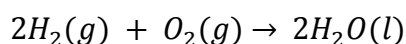
The production of hydrogen gas can be increased by;

- 1) Adding copper two sulphate. It acts as a catalyst increasing the rate of the reaction.  
Catalyst is a substance which alters the rate of a chemical reaction but it remains chemically unchanged at the end of the reaction.
- 2) Using zinc powder instead of zinc granules. The zinc powder has a greater surface area than the zinc granules hence more molecules would react to produce hydrogen gas.

### Testing for hydrogen gas

Hydrogen is tested by using a lighted splint. When a lighted splint is brought in close contact with hydrogen gas either in the test tube or in the gas jar, the hydrogen will explode with a pop sound.

### Equation



### Physical properties of hydrogen

- Hydrogen is a colourless gas.
- Hydrogen is the lightest gas known,
- It is insoluble in water.
- Hydrogen is odourless[has no smell]
- Hydrogen is a neutral gas ie it's neither acidic nor alkaline.

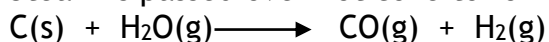
### Uses of hydrogen gas

- ❖ Hydrogen is used in filling air balloons because it's the lightest gas.
- ❖ Hydrogen is used in the hardening of oil, a process called hydrogenation. During this process, oil is converted to margarine by reacting it with hydrogen in the presence of a catalyst called nickel at 180 c and 2atms[pressure].the oil absorbs hydrogen to form fats.
- ❖ It's used in the manufacture of petrol from coal
- ❖ Hydrogen is used in the manufacture of a gas called ammonia in a process called Haber process. Ammonia is an important gas because it is used in the manufacture of fertilizers, nitric acid and explosives. During this process, hydrogen is reacted with nitrogen gas to give ammonia in the presence of iron as a catalyst.
- ❖ It's used in the manufacture of hydrochloric acid .in this process; hydrogen is reacted with chlorine to form hydrochloric acid.

## Industrial preparation of hydrogen gas

On a large scale, hydrogen gas is manufactured by Bosch process. The raw materials used include coke and steam.

Steam is passed over hot coke to form carbonmonoxide and hydrogen gas.

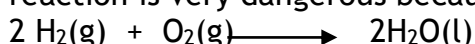


A combination of carbonmonoxide and hydrogen gas is water gas. The carbonmonoxide is absorbed in another chemical leaving hydrogen gas behind.

## Chemical properties of hydrogen gas

### a) Reaction of hydrogen gas with oxygen

Hydrogen burns with a blue flame in the presence of oxygen to form water. This reaction is very dangerous because it leads explosion of the test tube



The water is formed inform steam which later condenses on cooler parts of the test tube to form liquid water.

### b) Reaction of hydrogen gas with metal oxides.

An oxide is a compound formed between a metal or non metal with oxygen

Hydrogen removes oxygen atoms from oxides of metals in a process called reduction.

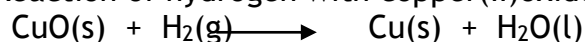
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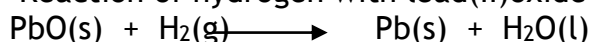
Reduction is the removal of oxygen from a metal oxide by hydrogen. The hydrogen in this case is referred to as the reducing agent

#### Examples

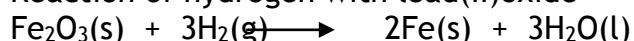
- ✓ Reaction of hydrogen with copper(II)oxide



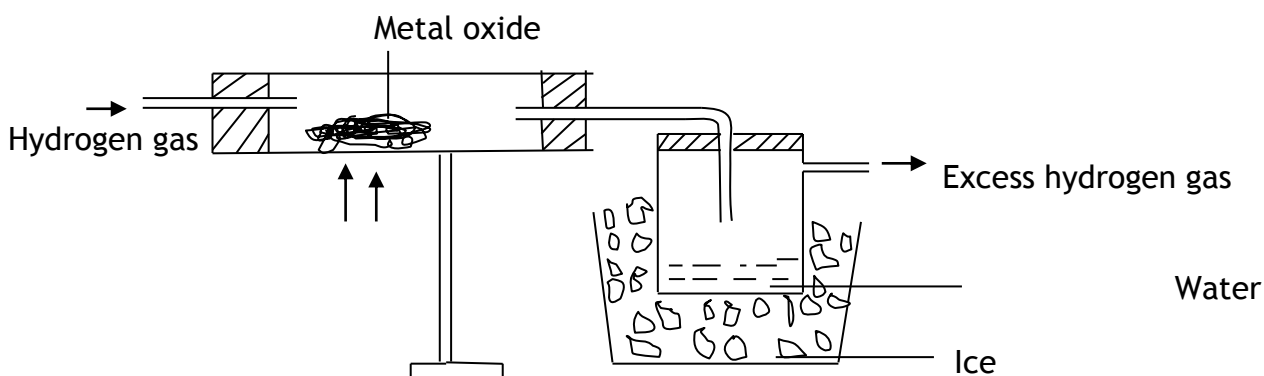
- ✓ Reaction of hydrogen with lead(II)oxide

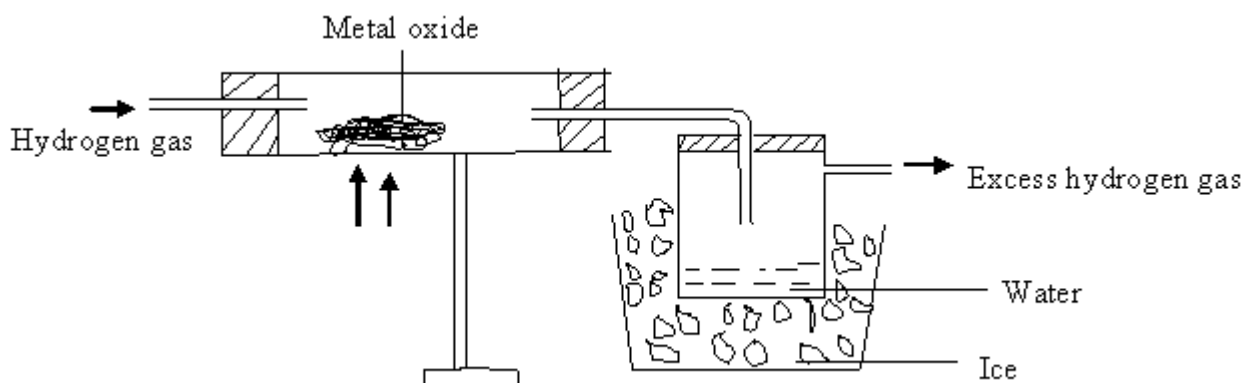


- ✓ Reaction of hydrogen with lead(II)oxide



The reaction of hydrogen with metal oxide can be demonstrated by the experiment below.





NB// the combustion should be slanted as shown above so that condensed water does not run back in to the combustion tube.

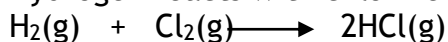
Ice is used to condense water.

Before the experiment the colour of the oxide was black and no water was formed in the test tube.

After the experiment, the colour of the oxide changed from black to brown due to the formation of copper and a liquid was formed in the test tube.

**c) Reaction of hydrogen with chlorine.**

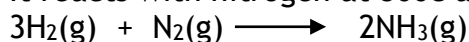
Hydrogen reacts with chlorine when heated to form hydrogen chloride gas.



When hydrogen chloride gas is dissolved in water it becomes hydrochloric acid.

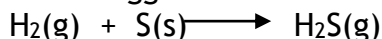
**d) Reaction of hydrogen with nitrogen**

It reacts with nitrogen at 500°C and 200 atm to form ammonia gas.



**e) Reaction of hydrogen with sulphur**

Hydrogen reacts with sulphur forming hydrogen sulphide gas which has a smell of rotten egg.



## WATER

Water is a chemical compound consisting of hydrogen in the ratio of 2: 1. It always contains dissolved substances and suspended solids. Therefore, water is never pure.

There are four natural sources of water.

- a) Rain water
- b) Spring water
- c) River water
- d) Sea water

**1) Rain water**

This water is fairly pure although it contains dissolved gases from the atmosphere e.g oxygen and carbon dioxide. It also contains dust particles.

**2) Spring/well water**

The water originates from the rocks underground. This water contains a number of dissolved minerals such as iron , calcium etc

**3) River water**

This water contains dissolved salts, suspended particles. It is mainly used for domestic purposes after treatment

#### 4) Sea water

This water contains more dissolved salts than any other sources of water. These salts include sodium chloride, magnesium chloride, calcium chloride, calcium sulphate etc the presence of these salts makes sea water very dense and salty

Sea water is major source of a number of elements and compounds e.g. bromine, potassium, sodium etc

#### Water cycle

The water cycle show how water circulates around the earth .the cycle begins when the sun evaporates water from seas, oceans, lakes and rivers. The water vapour rises to the atmosphere where it cools to form the rain clouds. The rain clouds further cool and at a certain temperature they turn into droplets which falls on the earth's surface inform of rainfall.

The rain water then runs into streams and rivers where it is taken to the lakes and seas .some of this rain water penetrates into underground where it forms underground water reservoir. The water contains a number of dissolved substances which are useful to living things.

#### Physical properties of water

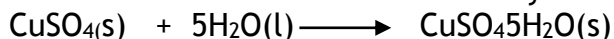
- Water is a liquid at room temperature
- Water is a colourless liquid
- Has no smell
- Water is tasteless substance i.e. it has flat taste.
- Pure water boils at 100c
- Water has a density of 1g/dm<sup>3</sup> at 4c
- Water is neutral compound i.e. it is neither acidic nor alkaline compound
- Water has a PH of 7

NB// it is pleasant to drink water which contains dissolved gases or solids because they give water its taste.

#### How to test for water in the laboratory

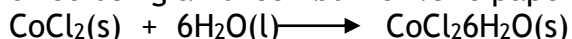
##### 1) Anhydrous copper(II)sulphate

Water turns the white colour of Anhydrous copper(II)sulphate to blue.



##### 2) Anhydrous Cobalt(II)chloride paper

This paper is made by soaking it in a solution of cobalt(II)chloride. The paper is then dried using a Bunsen burner .this paper turns from blue to pink when put in water.



#### Water in the atmosphere

- a) **Deliquescence:** this is the absorbing of water vapor from the atmosphere by a solid to form a solution. Substances which absorb water vapour from the atmosphere and change to a liquid or solution is called deliquescent.

Examples of deliquescent substances are.

-sodium hydroxide



- potassium hydroxide
- iron(II)chloride
- phosphorous pentaoxide

b) **Hygroscopy** .this is absorption of water vapour from the atmosphere by a substance without forming a solution or liquid. The substance absorbing water vapour simply becomes damp or wet.

**Hygroscopic substances.** These are substances which absorb water vapour from the atmosphere without forming a solution eg

- calcium oxide (quick lime)
- copper(II)oxide
- concentrated sulphuric acid
- anhydrous copper(II)sulphate

c) **Efflorescence.** These are substances which lose water of crystallization when exposed to the atmosphere e.g.

- hydrated sodium carbonate
- hydrated sodium sulphate

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### Action of water

The reactivity of metals with water is not the same. Other metals react very fast with water while others react very slowly with water.

Generally

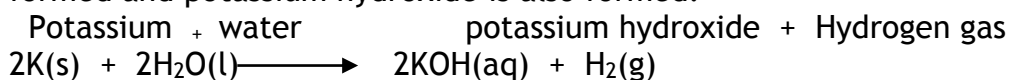
Metals + water  $\longrightarrow$  hydrogen gas + metal hydroxide/metal oxide

The arrangement of the reactivity of metals with water beginning from the most reactive to the least reactive is called the reactivity series or activity series. The reactivity series is given below

K	
Na	
Ca	
Mg	
Al	Reactivity decreases
Zn	
Fe	
Pb	
H	
Cu	
Hg	
Ag	
Au	

The reactivity series is the arrangement of metals according to their rate of reaction with water. It begins with the most reactive metal and ends with the least reactive metal.

a) Potassium. The reaction of potassium with water is very violent. Hydrogen gas is formed and potassium hydroxide is also formed.

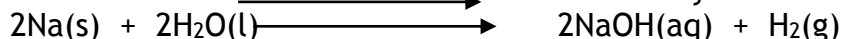


During this reaction, a lot of heat is produced or liberated which makes yhe hydrogen to start burning giving out light. The potassium dissolves in to silvery balls and eventually disappears. This reaction can be illustrated by the diagram below



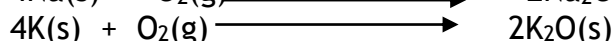
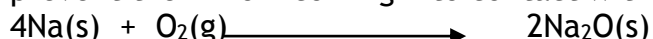
b) Sodium, sodium metal dissolves in water and this reaction is fast but not faster than that of potassium with water. Sodium hydroxide and hydrogen gas are formed

Sodium + water  $\longrightarrow$  Sodium hydroxide + hydrogen gas



During the reaction the sodium metal melts into silvery balls but the hydrogen liberated does not burn like the one of potassium therefore no heat and light is given out.

NB// Na and K are very reactive elements. For this reason they are stored under oil to prevent them from coming into contact with air.



c) Calcium. Calcium is a very reactive metal when put in water, it will sink in water giving out calcium hydroxide and hydrogen gas.

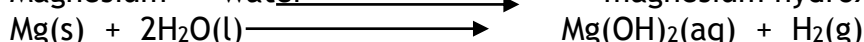
Calcium + water  $\longrightarrow$  calcium hydroxide + hydrogen gas



d) Magnesium. Magnesium reacts very slowly with cold water forming magnesium hydroxide and hydrogen gas but very fast with boiling water or steam forming white powder of magnesium oxide and a colourless gas that produces a pop sound on exposure to a burning splint.

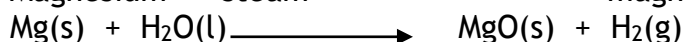
With cold water,

Magnesium + water  $\longrightarrow$  magnesium hydroxide + hydrogen

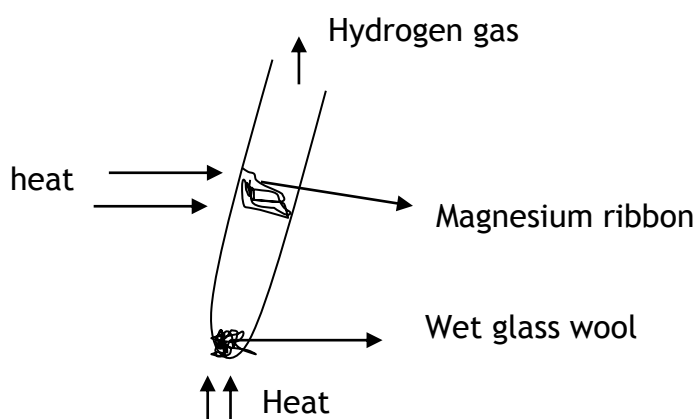


With steam or boiling water,

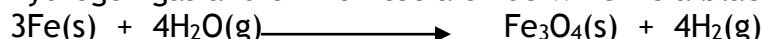
Magnesium + steam  $\longrightarrow$  magnesium oxide + hydrogen gas.



The reaction of magnesium with steam can be demonstrated using the set up below



e) Iron. Iron does not react with cold water but react with steam when heated to form hydrogen gas and tri-iron tetra oxide which is a black solid.



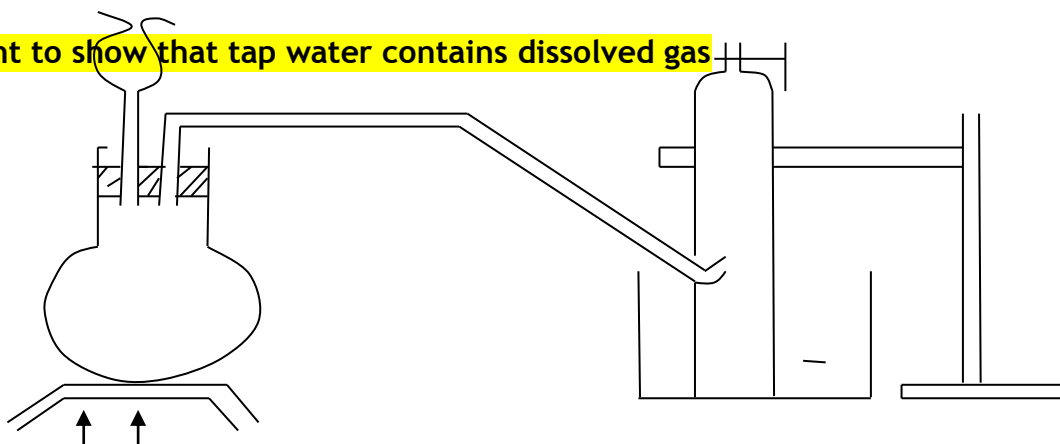
The reaction of iron with steam can be shown as below.

### Procedure

Iron is put in a combustion tube as shown above. And it is heated and then steam from boiling water is passed over the heated iron. Immediately a reaction takes place. The products formed eg hydrogen gas is collected in gas jar. The tri-iron tetra oxide remains in the combustion tube.

### Experiment to show that tap water contains dissolved gas

Diagram



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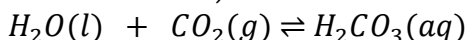
Water contains gases e.g. oxygen. When water is heated, the heat dries out the gases in the water. The oxygen rises and can be collected in the burette as shown above

### Hardness of water

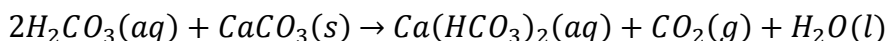
Hard water is one which does not readily or easily form lather with soap.

### How water becomes hard

When rain falls, it dissolves the carbon dioxide in the air to form carbonic acid.



When this water reaches the earth's surface, it passes through soil or rocks containing lime stone (calcium carbonate) or magnesium carbonate. The carbonic acid in the rain water reacts with the calcium carbonate or magnesium carbonate to form calcium hydrogen carbonate or magnesium hydrogen carbonate. The chemicals formed make water temporary hard.



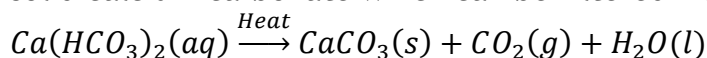
Another chemical which causes hardness in water is magnesium hydrogen carbonate.

**Soft water** is one which readily forms lather with soap e.g. rain water, distilled water, deionised water. There are two types of hardness of water.

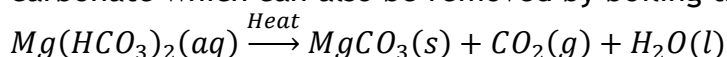
- (a) **Temporary hardness.** This is hardness due to the presence of calcium hydrogen carbonate in water. It is called temporary hardness because it can be removed by simply boiling the water.

## Methods of removing temporary hardness

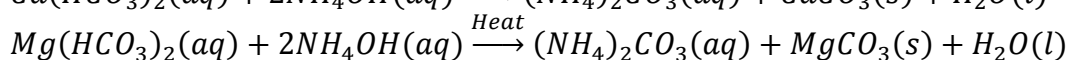
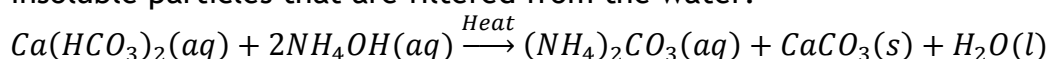
- **Boiling of water.** The heat decomposes calcium hydrogen carbonate into solid calcium carbonate which can be filtered from water.



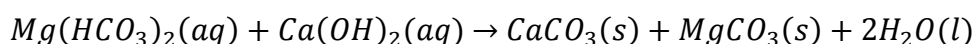
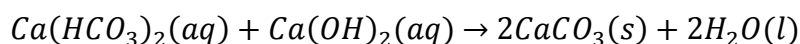
Another substance that causes temporary hardness is magnesium hydrogen carbonate which can also be removed by boiling the water.



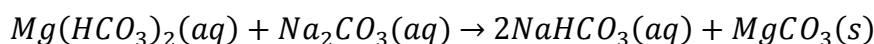
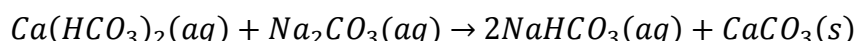
- **By addition of ammonia solution to the water.** Ammonia solution reacts with calcium hydrogen carbonate or magnesium hydrogen carbonate to form insoluble particles that are filtered from the water.



- **By adding calcium hydroxide (slaked lime) to the water.** Insoluble particles are formed and are filtered out of the water.



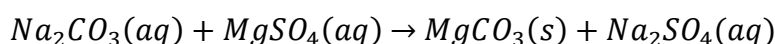
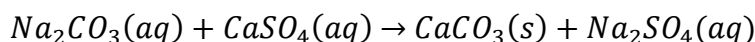
- **By adding sodium carbonate (washing soda).** Insoluble particles are formed and are filtered out of the water.



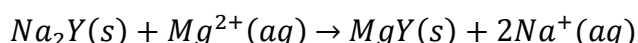
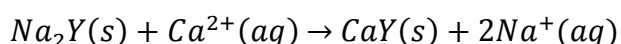
- (b) **Permanent hardness.** This is the hardness caused by the presence of dissolved calcium sulphate and magnesium sulphate in water. The hardness cannot be removed by boiling because the compounds in them do not decompose when heated hence the name permanent hardness.

## Methods of removal of permanent hardness

- By addition of sodium carbonate (washing soda) to the water. This removes both temporary and permanent hardness.



- **Permutit process.** This method is used to treat large supplies of water. Permutit is a complex substance with formula  $Na_2Y$ . It is normally packed in a container and hard water is allowed to pass through the permutit. During the process, ion exchange takes place i.e. the calcium and magnesium ions in hard water are exchanged with the sodium ions of the permutit as shown by the equations below.



This method is very expensive.

- **Distillation**

## Disadvantages of hard water

- The scum formed (calcium and magnesium stearate) formed stains clothes leaving them with changed colour. Scum is insoluble calcium stearate and magnesium stearate precipitated when soap reacts with water.
- It leads to the formation of furs in kettles and boilers. Some times it is referred to as the boilers scale. The presence of furs in the boiler makes the water to take longer time to boil hence leading to wastage of fuel.

## Advantages of hard water

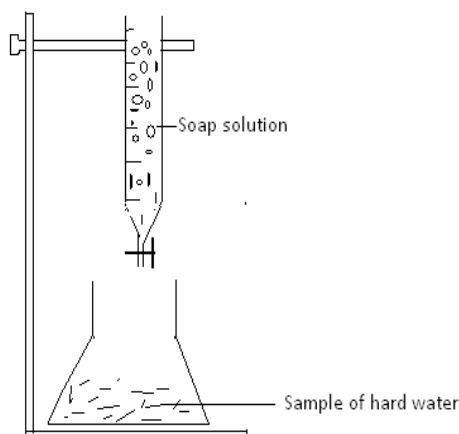
- It improves the taste of water, beer hence used in brewing industries.
- It leads to formation of hard tissues in animals because it contains calcium ions used for their formation e.g. bones, teeth, egg shells, scales.
- It prevents lead poisoning especially when lead pipes are used in transporting water.

## Experiment to determine hardness of water

### Requirements

- Burette
- Conical flasks
- Funnel
- Retort stand
- Samples of hard water
- Liquid soap

### Setup



### Procedure

- Soap solution is put in the burette as shown in the diagram.
- A known volume of water is also put in a beaker.
- The soap solution is then run into the different samples of water before boiling and after boiling the sample.
- The experiment is done until the soap forms stable bubbles of lather.
- The volume of soap used in each sample is recorded.

### Sample results

Assuming the 10 cm<sup>3</sup> of samples were used in all the cases.

Type of hardness	Volume of soap before boiling	Volume of soap after boiling
Soft water	2	2
Permanent hard water	12	12
Temporary hard water	10	3

### Explanation

- (a) Soft water. It requires small and constant amount of soap solution before and after boiling. This implies that it is not hard water.
- (b) Temporary hard water. It required bigger amount of soap before boiling than after boiling. This means that the hardness was removed by boiling.
- (c) Permanent hard water. It required large and constant amount of soap before and after boiling meaning that boiling did not remove the hardness.

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## ACIDS, BASES AND SALTS

### ACID

An acid is a substance that dissolves in water to produce hydrogen ions as the only positively charged ions. Examples of acid include;

- Hydrochloric acid
- Sulphuric acid
- Nitric acid
- Ethanoic acid
- Carbonic acid
- Phosphorous acid

### Characteristics of an acid

- Turns blue litmus paper red and has no effect on red litmus paper.
- It reacts with metals to liberate hydrogen gas and a salt is formed.
- It reacts with carbonates and hydrogen carbonates to liberate carbondioxide, salt and water.
- It reacts with a base to form salt and water only.

### Concentration of acids

An acid can either be dilute or concentrated.

- 1) **Dilute acid.** This is one which contains a large amount of water and very small quantity of the acid.
- 2) **Concentrated acid.** This contains a very large amount of acid dissolved in a little amount of water.

### Strength of acid

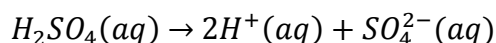
The strength of any acid is determined by the number or amount of hydrogen ions an acid can produce when dissolved in water i.e. some acids produce very many hydrogen ions while others produce very few hydrogen ions. On this basis, acids are of two types;

- 1) **Strong acids.** They are acids which when dissolved in water dissociates/ionizes completely to produce very many free hydrogen ions.

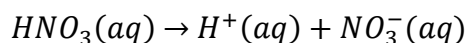
**NB//.** Dissociation/ionisation is the breakdown of a substance into particles which are capable of moving freely in solution.

**Examples of strong acids include;**

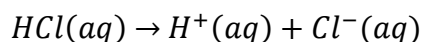
- ❖ Sulphuric acid



- ❖ Nitric acid



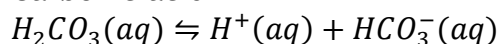
- ❖ Hydrochloric acid



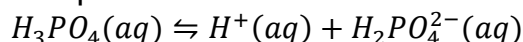
- 2) **Weak acids.** They partially dissociates/ionizes in water to form few hydrogen ions.

**Examples include;**

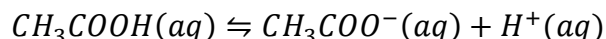
- ❖ Carbonic acid



- ❖ Phosphoric acid



- ❖ Ethanoic acid



**NB//.**  $\rightleftharpoons$  means partial/incomplete reaction and the reaction is reversible.

An acid can also be referred to as mineral or organic acid.

**Examples of mineral acids**

- Sulphuric acid
- Nitric acid
- Hydrochloric acid
- Carbonic acid
- Phosphoric acid
- Phosphorous acid
- Sulphurous acid

**Examples of organic acids**

- Citric acid (lemon juice)
- Lactic acid (milk)
- Oxalic acid (moss)
- Ethanoic acid (alcohols)
- Tartaric acid (baking acid)
- Methanoic acid (wasps, bees)

### Basicity of an acid

This is the number of hydrogen ions produced by one molecule of an acid when dissolved in water.

**NB//.** Basicity is not always equal to the number of hydrogen ions in one molecule of the acid e.g. the basicity of ethanoic acid is one (monobasic) yet it has four hydrogen ions three of which cannot be replaced.

Some acids produce 2, 3 or more hydrogen ions per molecule.

### Types of basicity

- (a) **Monobasic acid.** It is an acid which dissociates to produce only one hydrogen ion when dissolved in water. The acid is said to have a basicity of one e.g.
- Hydrochloric acid
  - Nitric acid
  - Ethanoic acid
- (b) **Dibasic acid.** It is an acid which ionizes to produce only two hydrogen ions from its molecule when dissolved in water. The basicity of this acid is two. E.g.
- Sulphuric acid
  - Carbonic acid
- (c) **Tribasic acid.** This is an acid which dissociates to produce only three hydrogen ions when dissolved in water. E.g.
- ❖ Phosphoric acid

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### Chemical properties of acids

- 1) **Reaction with metals.** All acids react with metals to liberate hydrogen gas and a salt is formed.

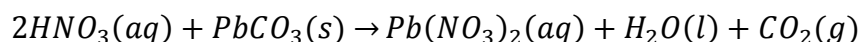
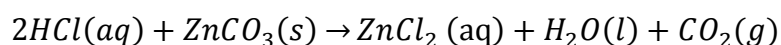
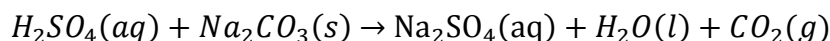
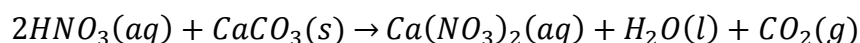
*Acid + metal → salt + hydrogen gas*

**NB//.** Metals below hydrogen in the activity series do not react with acids.

K	}	Do react with acids
Na		
Al		
Mg		
Zn		
Fe		
Pb		
H	}	Do not react with acids
Cu		
Hg		
Au	}	

- 2) **Reaction with carbonates.** All carbonates react with acids to form salt, water and carbon dioxide gas.

*Acid + carbonate → salt + water + carbon dioxide*



**NB//.** Sulphuric acid does not react with calcium carbonate and barium carbonate.

- 3) **Reaction with hydrogen carbonates.**



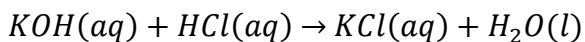
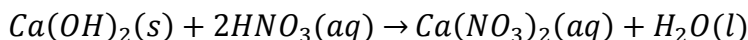
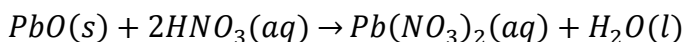
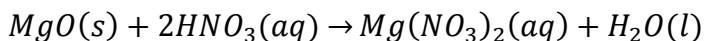
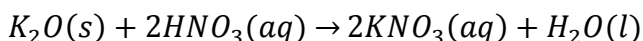
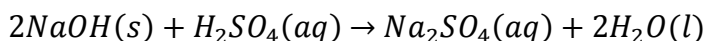
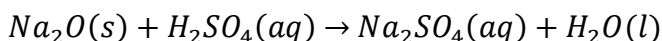
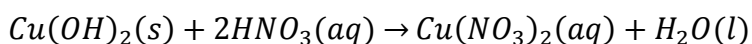
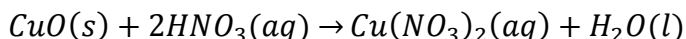
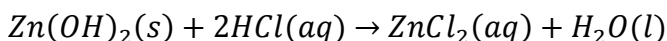
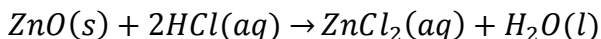
#### 4) Reaction with bases.

##### Bases

A base is metal oxide or hydroxide which reacts with an acid to form salt and water only. The reaction is called **neutralization** reaction.

*Base + acid → salt + water*

Eg



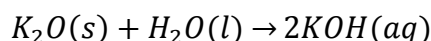
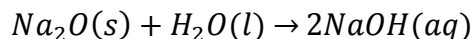
There are two types of bases;

- Metal oxides
- Metal hydroxides

##### Metal oxides

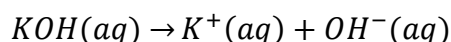
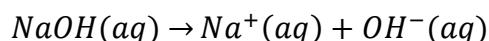
- Solubility of metal oxides in water. All metal oxides are insoluble in water except sodium oxide and potassium oxide which are soluble.
- They dissolve in water to give metal hydroxides.

*Soluble basic oxide + water → metal hydroxide*



##### Metal hydroxides

- All metal hydroxides are insoluble in water except sodium hydroxide and potassium hydroxide.



**NB//.** If a solid compound dissolves in water, it implies that the water has caused the solid to split into ions.

## Alkalis

An alkali is a metal hydroxide which dissolves in water to give hydroxide ions as the only negatively charged ions. There are only two alkalis present i.e. sodium hydroxide and potassium hydroxide. The rest of the hydroxides are not alkalis e.g. calcium hydroxide, magnesium hydroxide, copper (II) hydroxide.

### Properties of bases (alkalis)

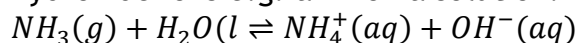
- Their solutions feel soapy/greasy.
- Their solutions are bitter.
- Their solutions turn red litmus paper to blue.
- They react with acids to form salt and water and this reaction called neutralization reaction.

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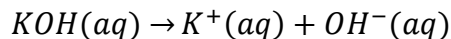
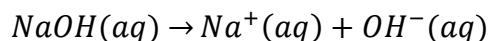
### Strength of a base

Bases are said to be weak or strong.

- 1) **Weak base.** This is base which partially dissociates or ionizes to produce a few hydroxide ions e.g. ammonia solution.



- 2) **Strong base.** A strong base is one which completely dissociates or ionizes in solution to produce many hydroxide ions e.g. sodium hydroxide and potassium hydroxide.



**NB//.** Hydroxide ions are responsible for the change of colour of litmus paper from red to blue while hydrogen ions are responsible for the change of blue litmus paper to red.

### Indicators

An indicator is a substance which shows whether a solution is acidic, alkaline or neutral. Common examples of indicators include;

Indicators	Colour in acid	Colour in base
Methyl orange	Red	Yellow
Phenolphthalein	Colourless	Pink
Methyl red	Red	Yellow
Litmus paper	Red	Blue
Bromothymol red		
Litmus solution	Red	Blue
Universal indicator	red	blue

### Preparation of indicators in the laboratory from flowers of red cabbage

In the laboratory, indicators can be obtained from plant extracts especially brightly coloured flowers. Eg hibiscus. The procedure is as follows;

- Collect some brightly coloured flowers e.g. hibiscus or red cabbage and using a mortar crash the flowers thoroughly into a paste and add ethanol to the paste.

- Stir the solution well then filter the solution into a boiling tube. The filtrate obtained is the indicator and then discard the residue.
- Record the colour of the filtrate. (which type of indicator is this?)

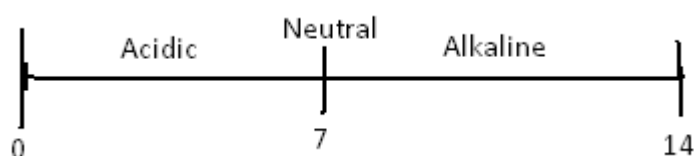
**NB//.** Ethanol is used as a solvent because indicators are insoluble in water but soluble in ethanol which is a non polar solvent. Therefore the indicator obtained can be used to test for the acidity and alkalinity of different solutions.

## Universal indicator

This is mixture of indicators in solution. It enables us to classify solutions into neutral, alkaline and acidic.

## pH Scale

pH refers to the degree of alkalinity or acidity of a solution. This is shown by a pH meter. A pH meter classifies solutions as alkaline, acidic or neutral. In pH meter, the scale ranges from 0-14 as shown below.



- Acidic solutions have pH of less than 7.
  - Alkaline solutions have high pH i.e. greater than 7.
  - Neutral solutions have pH of 7.
- Strongly acidic. This is a solution with the lowest pH value ie 0-3.
  - Weakly acidic. A solution with slightly higher pH value ie 4-6.9
  - Strongly alkaline (basic). This is a solution with the highest pH value ie 10 - 14.
  - Weakly alkaline this is a solution with pH value just above 7 i.e. 7.1 to 9.9

## Salts

A salt is a compound formed when all the replaceable or part of the replaceable hydrogen ions of an acid has been replaced by a metal or ammonium ion. The name of the salt is derived from the acid from which it is formed.

- All salts formed from sulphuric acid are called sulphates eg  
Copper (II) sulphate  
Zinc sulphate  
Sodium sulphate  
Barium sulphate  
Iron (II) sulphate  
Iron (III) sulphate  
Ammonium sulphate  
Aluminium sulphate
- All salts derived from hydrochloric acids are called chlorides.eg  
Sodium chloride  
Magnesium chloride  
Zinc chloride  
Iron (III) chloride

- All salts derived from nitric acids are called nitrates.  
Aluminium nitrate  
Copper (II) nitrate  
Sodium nitrate  
Ammonium nitrate  
Calcium nitrate

## Types of salts

There are two types of salts;

- Normal salt
- Acid salt

### Normal salt

This is formed when all the replaceable hydrogen ions of an acid have been replaced by a metal or ammonium ion i.e. these salts do not contain hydrogen ions from the acid. eg sodium sulphate, sodium chloride, copper (II) sulphate.

### Acid salt

This is formed when only part of the replaceable hydrogen ions has been replaced by a metal or ammonium ion. Therefore an acid salt is only formed by sulphuric acid because it has more than one replaceable hydrogen ions. Nitric acid and hydrochloric acid cannot form an acid salt because they contain only one replaceable hydrogen ion. E.g. of acid salts include; sodium hydrogen sulphate, ammonium hydrogen sulphate, sodium hydrogen carbonate.

## Methods of preparing salts in the laboratory

The method used depends on the solubility of the salt in water. Generally;

- All sulphates are soluble in cold water except barium sulphate, lead (II) sulphate and calcium sulphate.
- All chlorides are soluble in water except lead (II) chloride, silver chloride.
- All nitrates are soluble in water.
- All carbonates are insoluble in water except sodium carbonate, potassium carbonate and potassium carbonate.

Based on the solubility of salts there are two methods of preparing salts;

- (i) Crystallization
- (ii) Precipitation (double decomposition)

### Crystallization

This method is used to prepare soluble salts. During this method, salt crystals are formed from a salt solution. The following procedures are used during the process of crystallization.

- A salt solution is made by dissolving a metal oxide or a metal in an acid until no more dissolves. The undissolved solids are then filtered out to obtain a filtrate which contains the salt solution.

- The filtrate is then put in an evaporating dish and heated to drive off the excess water from the salt solution.
- The hot solution is removed from the source of heat and allowed to cool in order for the crystals to form. The liquid that remains on top of the crystals after cooling is called mother liquor.
- The crystals formed are then filtered out from the mother liquor, dried on filter paper to obtain the salt.

#### Diagrams

**NB//.** To ensure that all the acid put in the beaker is reacted completely, the metal oxide or metal must be added until no more dissolves. This implies that all the acid put in the beaker is reacted.

#### Double decomposition (precipitation)

This method is used to prepare insoluble salts i.e. salts which do not dissolve in cold water. During this method, two solutions are mixed and a solid called a precipitate is formed.

Precipitation is a process of obtaining a solid by mixing two different solutions. The solid formed is called a precipitate. The precipitate is then filtered to obtain the salt.

#### Illustrations

#### Preparation of soluble salts

*metal + acid → salt + hydrogen gas*

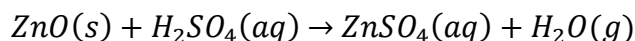
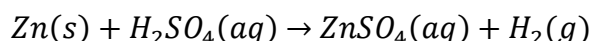
*acid + metal oxide → salt + water*

*acid + metal carbonate → salt + water + carbon dioxide*

*acid + metal hydroxide → salt + water*

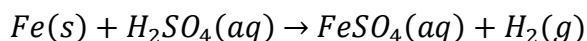
#### Preparation of zinc (II) sulphate crystals

- Dilute sulphuric acid is put in a beaker and warmed.
- Zinc granules or zinc oxide is added to the warm acid little at a time with constant stirring until no more can dissolve. This means that all the acid has reacted.
- The undissolved zinc or zinc oxide is filtered out and a colourless filtrate is obtained.
- The filtrate is then transferred into an evaporating dish and heated to evaporate excess water from the solution.
- The hot solution is removed from the heat source and allowed to cool for the crystals of zinc sulphate to be formed.
- The crystals are then filtered out and dried between the filter papers.



**NB//.** The acid is warmed to increase the rate of reaction.

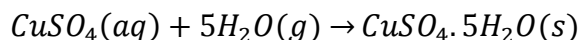
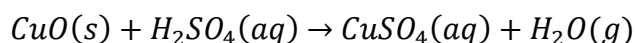
### Preparation of iron (II) sulphate from iron and dilute sulphuric acid



- Dilute sulphuric acid is put in a beaker and warmed.
- Iron fillings are added to the warm acid little at a time with constant stirring until no more can dissolve. This means that all the acid has reacted.
- The undissolved iron fillings are filtered out and the green filtrate obtained has iron (II) sulphate.
- The filtrate is then transferred into an evaporating dish and heated to evaporate excess water from the solution.
- The hot solution is removed from the heat source and allowed to cool for the crystals of iron sulphate to be formed.
- The crystals are then filtered out and dried between the filter papers.

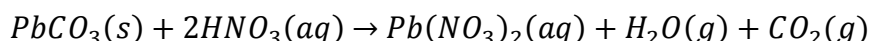
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### Preparation of copper (II) sulphate crystals from copper (II) sulphate and dilute sulphuric acid



- Dilute sulphuric acid is put in a beaker and warmed.
- Copper (II) oxide crystals are added to the warm acid little at a time with constant stirring using a glass rod until no more can dissolve. This means that all the acid has reacted.
- The undissolved copper (II) oxide is filtered out and a blue filtrate of copper (II) sulphate is obtained.
- The filtrate is then transferred into an evaporating dish and heated to evaporate excess water from the solution.
- The hot solution is removed from the heat source and allowed to cool for the crystals of copper sulphate to be formed.
- The crystals are then filtered out and dried between the filter papers.

### Preparation of lead (II) nitrate crystals from lead (II) carbonate and dilute nitric acid



- Dilute nitric acid is put in a beaker and warmed.
- Lead (II) carbonate crystals are added to the warm acid little at a time with constant stirring using a glass rod until no more can dissolve. During this reaction carbondioxide gas is evolved. This means that all the acid has reacted.

- The undissolved copper (II) oxide is filtered out to obtain a colourless filtrate which contains lead (II) nitrate.
- The filtrate is then transferred into an evaporating dish and heated to evaporate excess water from the solution. This is done until the crystals begin to form.
- The hot solution is removed from the heat source and allowed to cool for more crystals of lead (II) nitrate to be formed.
- The crystals are then filtered out and dried between the filter papers.

**NB//.** Lead (II) nitrate can also be prepared from lead (II) oxide dilute nitric acid.

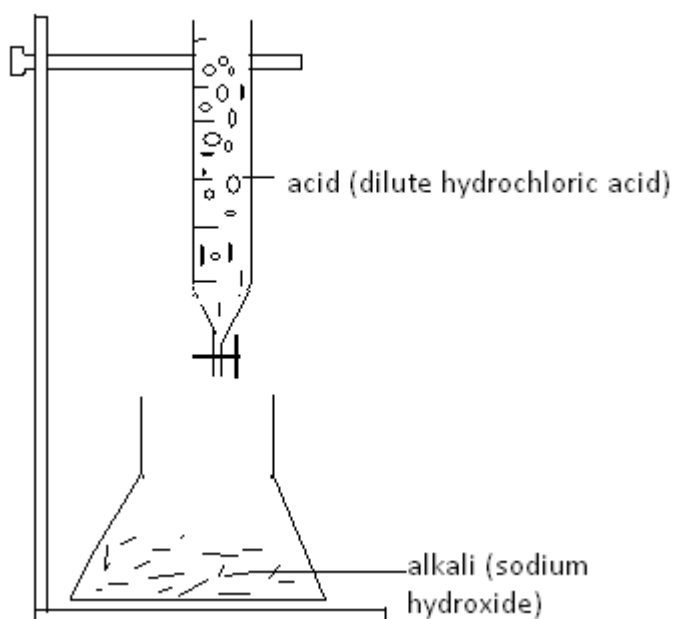
The same procedures outlined above are used for preparation of other salts involving the action of acid of carbonates e.g. copper (II) nitrate, iron (II) nitrate.

### Action of dilute acid on alkalis (metal hydroxide)

*Acid + alkali → salt + water*

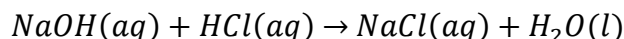
This reaction is called neutralization and it is used to prepare common salt (sodium chloride) by a technique called titration where an acid is put in a burette and the alkali is put in a conical flask. The acid is then run to the alkalis in the conical flask by opening the tap of the burette.

### Titration process



### Preparation of sodium chloride from dilute hydrochloric acid and sodium hydroxide

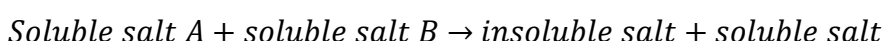
- Fill the burette with dilute hydrochloric acid and note the level of the acid.
- Pipette about 25 cm<sup>3</sup> of sodium hydroxide and transfer it into a clean conical flask and add 2-3 drops of phenolphthalein indicator to this solution.
- Then run the acid from the burette into the sodium hydroxide solution until the colour of the solution turns from pink to colourless. The colourless solution is sodium chloride.



- Transfer the colourless solution into an evaporating dish and heat to evaporate the excess water until the crystals begin to form.
- Remove the hot solution from the heat source and allow it to cool for the crystals to form completely.
- Filter out the crystals and dry them between filter papers. Repeat the process of crystallization to purify the crystals.

### Preparation of insoluble salts by double decomposition (precipitation)

Insoluble salts are prepared by a method called double decomposition/precipitation method. During the process, two soluble salts are mixed and an insoluble salt is formed or precipitated.



### ATOMIC STRUCTURE

Atomic structure deals with the study of the composition of an atom. An atom is the smallest indivisible particle of an element that can take part in a chemical change. Studies have revealed that an atom consists of three fundamental particles. They include;

- Protons
- Neutrons
- Electrons

These particles differ from each other in terms of mass and charges which are shown below;

Particle	Mass (atomic mass unit)	Charge
Proton	1	+1
Neutron	1	0
Electron	$\frac{1}{1840}$	-1

$\frac{1}{1840}$  means that the electron is  $\frac{1}{1840}$  times lighter than protons and neutrons. An atom is said to consist of two regions i.e. nucleus and orbit/shell. This can be diagrammatically represented as below.

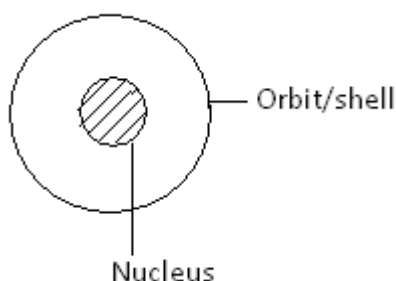


Figure 1 structure of atom

### Characteristics of the nucleus

- It contains the protons and neutrons which are collectively called nucleons.



- It occupies the central part of the atom. This part is smaller compared to the orbit.
- The nucleus is extremely dense (heavy) because of the presence of protons and neutrons.
- The mass of the atom (atomic mass/mass number) is the sum of the number of protons and neutrons in the nucleus of the atom.
- The number of protons is not always equal to the number of neutrons in some atoms but the number of protons is always equal to the number of electrons in a neutral atom.eg

In sodium,  $p = 11, n = 12$  and  $e = 11$

**NB//.** Hydrogen is the only atom which does not have a neutron in its nucleus. For this reason it is considered the lightest atom.  $p = 1, n = 0$  and  $e = 1$

### Orbit/shell

This is the space outside the nucleus of an atom electrons revolve or are found. The greater the number of electrons in an atom the greater the number of orbits. Studies have shown that electrons rotate round the orbits i.e. they are not stationary within the orbit.

- Electrons occupy the orbits according to their energy levels. The orbits nearer to the nucleus have the lowest energy while those orbits far from the nucleus have the highest energy.
- An electron can only remain rotating within the same orbit when the centripetal force and centrifugal forces are balanced.
- The chemical property of an element (atom) is determined only by the electrons and not the protons and neutrons.

### Atomic number (z) and atomic mass (A)

An atom of any element differs from those of the other elements due to the difference brought about by atomic number and atomic mass of different elements.

#### Atomic number

This is defined as the number of protons in the nucleus of an atom and is indicated by letter z.

#### Atomic mass (mass number)

It refers to the total number of protons and neutrons in the nucleus in the nucleus of an atom. It is denoted by letter A. therefore an atom of an element can be represented in terms of its atomic number and mass number.eg

${}^A_ZY$  Where  $A = \text{mass number}$

$z = \text{atomic number}$

#### Summary

*Mass number = number of protons + Number of neutrons*

*Atomic number = Number of protons*

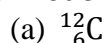
Therefore number of neutrons = Mass number – atomic number

Elements	Mass number	Atomic number	p	n	e
Hydrogen	1	1	1	0	1
Helium	4	2	2	2	2
Lithium	6	3	3	3	3
Beryllium	9	4	4	5	4
Boron	10	5	5	5	5
Carbon	12	6	6	6	6
Nitrogen	14	7	7	7	7
Oxygen	16	8	8	8	8
Fluorine	18	9	9	9	9
Neon	20	10	10	10	10
Sodium	23	11	11	12	11
Magnesium	24	12	12	12	12
Aluminium	27	13	13	14	13
Silicon	28	14	14	14	14
Phosphorous	31	15	15	16	15
Sulphur	32	16	16	16	16
Chlorine	35	17	17	18	17
Argon	36	18	18	18	18
Potassium	39	19	19	20	19
Calcium	40	20	20	20	20

Figure 2 showing mass number, atomic number, and number of protons, neutrons and electrons for the first 20 elements

### Worked examples

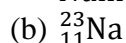
Work out the number of protons, electrons and neutrons for the following elements;



Number of protons = 6

Number of electron = 6

Number of neutrons =  $12 - 6 = 6$



Number of protons = 11

Number of electron = 11

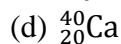
Number of neutrons = 12



Number of protons = 1

Number of electron =

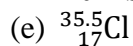
Number of neutrons = 0



Number of protons = 20

Number of electron = 20

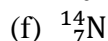
Number of neutrons = 20



Number of protons = 17

Number of electron = 17

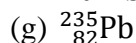
Number of neutrons = 18.5



Number of protons = 7

Number of electron = 7

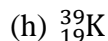
Number of neutrons = 7



Number of protons = 82

Number of electron = 82

Number of neutrons = 153



Number of protons = 19

Number of electron = 19

Number of neutrons = 20

## Isotopes and isotopy

An isotope refers to atoms of an element which have the same atomic number but different mass number. Isotopy refers to the existence of different atoms of an element with the same atomic number but different mass numbers.

### Characteristics of isotopes

- They have the same number of protons.
- They have different number of neutrons.
- They have same chemical properties.
- They have the same number of electrons.

Examples of elements which show isotopy include;

- Chlorine  $\begin{matrix} {}^{35}_{17}\text{Cl} & {}^{36}_{17}\text{Cl} & {}^{37}_{17}\text{Cl} \end{matrix}$
- Hydrogen  $\begin{matrix} {}^1_1\text{H} & {}^2_1\text{H} & {}^3_1\text{H} \end{matrix}$
- Oxygen  $\begin{matrix} {}^{16}_8\text{O} & {}^{18}_8\text{O} \end{matrix}$
- Lead  $\begin{matrix} {}^{206}_{82}\text{Pb} & {}^{208}_{82}\text{Pb} \end{matrix}$

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## Electronic configuration and structure

Electronic configuration is the arrangement of electrons within the orbit of an atom. The number of electrons in each orbit is given by the formula  $2n^2$  where  $n$  is the number of orbits.  $n = 1, 2, 3 \dots$

If  $n = 1$ , number of electrons = 2

If  $n = 2$ , number of electrons = 8

If  $n = 3$ , number of electrons = 18

If  $n = 4$ , number of electrons = 32

### Writing electronic configuration

$\begin{matrix} {}^1_1\text{H} & 1 \end{matrix}$	$\begin{matrix} {}^{39}_{19}\text{K} & 2: 8: 8: 1 \end{matrix}$	$\begin{matrix} {}^{32}_{16}\text{S} & 2: 8: 6 \end{matrix}$
$\begin{matrix} {}^{27}_{13}\text{Al} & 2: 8: 7 \end{matrix}$	$\begin{matrix} {}^{40}_{20}\text{Ca} & 2: 8: 8: 2 \end{matrix}$	$\begin{matrix} {}^{16}_8\text{O} & 2: 6 \end{matrix}$
$\begin{matrix} {}^{23}_{11}\text{Na} & 2: 8: 1 \end{matrix}$	$\begin{matrix} {}^{35}_{17}\text{Cl} & 2: 8: 7 \end{matrix}$	$\begin{matrix} {}^{14}_7\text{N} & 2: 5 \end{matrix}$
$\begin{matrix} {}^{24}_{12}\text{Mg} & 2: 8: 2 \end{matrix}$	$\begin{matrix} {}^{12}_6\text{C} & 2: 4 \end{matrix}$	
$\begin{matrix} {}^{31}_{15}\text{P} & 2: 8: 5 \end{matrix}$		

**NB//.** The number of electrons in the outer most shell is equal to the valency for elements from group 1 to group 4.

For elements whose outer orbits have more than four electrons subtract those from eight to get the valency.

### Electronic configuration of noble gases

Noble gases are said to be unreactive elements because they have a stable electronic configuration in which their outermost orbit has eight electrons giving them valency of zero. These noble gases are;

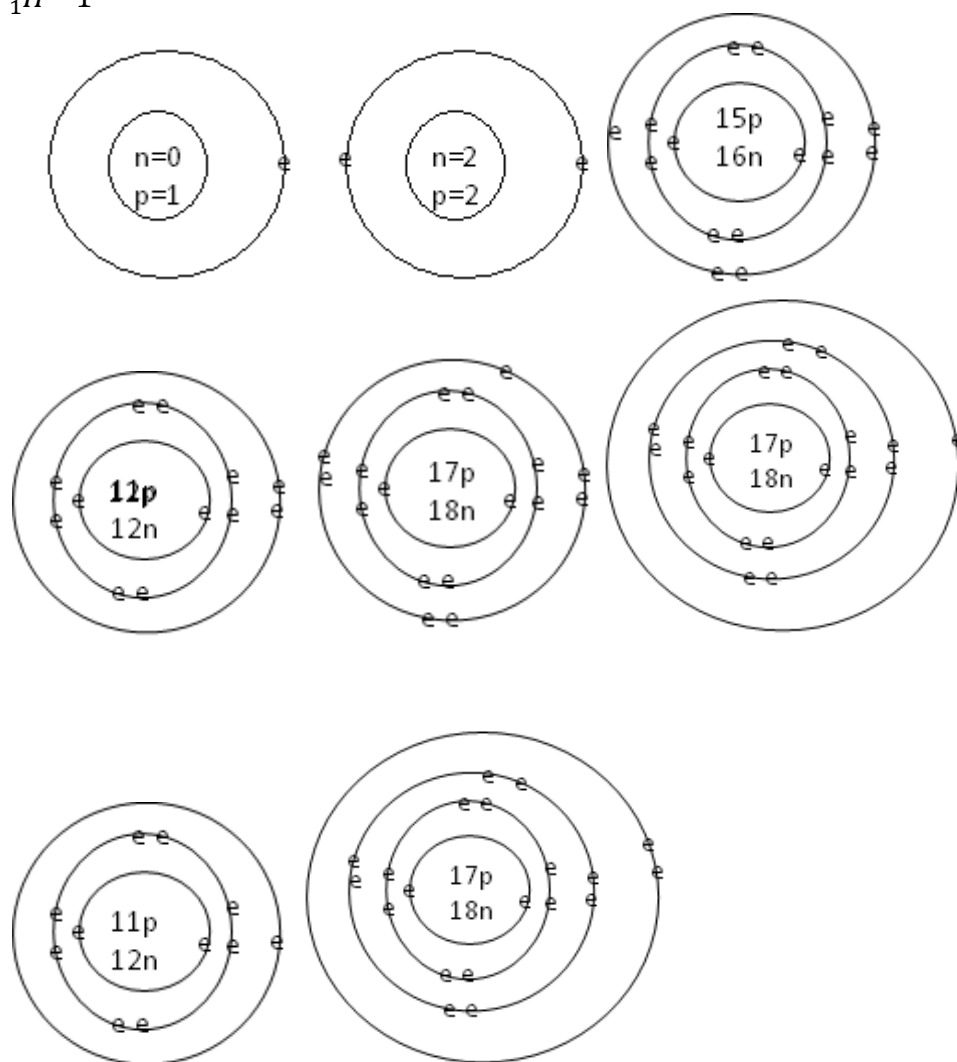
$\begin{matrix} {}^4_2\text{He} & 2 \end{matrix}$
$\begin{matrix} {}^{20}_{10}\text{Ne} & 2: 8 \end{matrix}$
$\begin{matrix} {}^{36}_{18}\text{Ar} & 2: 8: 8 \end{matrix}$
$\begin{matrix} {}^{52}_{26}\text{Kr} & 2: 8: 8: 8 \end{matrix}$

Elements do react in order to gain stable electronic configuration like the ones of noble gases. If an element has less than four electrons in its outermost orbit it will react by losing their outermost electrons so as to gain a stable electronic configuration of the noble gases. E.g.

Element	Before reaction	After reaction
Sodium	2: 8: 1	2: 8
Potassium	2: 8: 8: 1	2: 8: 8
Magnesium	2: 8: 2	2: 8
Aluminium	2: 8: 3	2: 8

### Electronic structure of an atom

This is a diagrammatic representation of an atom showing the number of protons and neutrons in the nucleus and also the number of electrons in the different orbits. Therefore in order to draw an electronic structure you must first write the electronic configuration of the atom. eg



The information obtained from the electronic configuration of an atom can be used in the following ways;

- To tell whether the element is a metal or non metal.
- For metals their outermost shell has less than four electrons.  
 $\left. \begin{array}{l} X, 2: 8: 2 \\ Y, 2: 8: 1 \\ Z, 2: 8: 3 \end{array} \right\} \text{Metals}$
- For non-metals their outermost shell contains more than four electrons.  
 $\left. \begin{array}{l} A, 2: 6 \\ B, 2: 5 \\ C, 2: 8: 7 \end{array} \right\} \text{Non-metals}$
- Metalloids are elements which behave neither like a non-metal nor a metal. They normally have four electrons in their outermost shell.  
 $\left. \begin{array}{l} C, 2: 4 \\ Si, 2: 8: 4 \end{array} \right\} \text{Metalloids}$
- For the outermost shell of an element is called its valency electrons because it determines the valency of the element. If an element has less than four electrons, then those become the valency of that element. eg  
 $X, 2: 8: 2$  valency is 2  
 $Y, 2: 8: 1$  valency is 1  
 $Z, 2: 8: 3$  valency is 3
- If an element has more than four electrons in its valency shell, then the valency of that element will be eight minus the number of electrons in the outermost shell.  
 $A, 2: 6$  valency is 2  
 $B, 2: 5$  valency is 3  
 $C, 2: 8: 7$  valency is 1

## Bonding and structure

Bonding is the combination of two or more atoms during a chemical reaction to form a stable compound. During the reaction, a chemical bond is formed between the atoms. There are three types of chemical bond that can be formed between two or more atoms.

- Electrovalent/ionic bond.
- Covalent bond.
- Metallic bond.
- Dative bond.

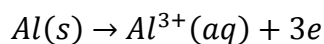
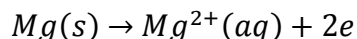
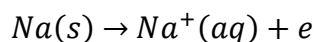
### Electrovalent/Ionic bonding

This is a type of bonding that occurs as a result of complete transfer of electrons from one atom to another. It normally takes place when a metal atom transfers electrons to a non-metal atom. I.e.

$A + B \rightarrow A^+ - B^-$  This represents transfer of electrons from A to B.

The type of bond formed as a result of transfer electrons is called electrovalent/ionic bond. Electrovalent bonding leads to formation of particles called ions. An ion is a charged particle formed by either loss or gain of electrons by an atom during a reaction. There are two types of ions;

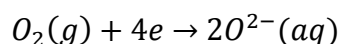
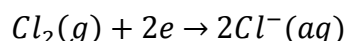
(a) **Positive ions (cations).** This is formed when a metal atom loses electrons because metals have high tendency to lose electrons.eg



The ions formed bear positive charge because they contain more protons than electrons.

In a neutral atom the number of electrons is always equal to the number of protons.

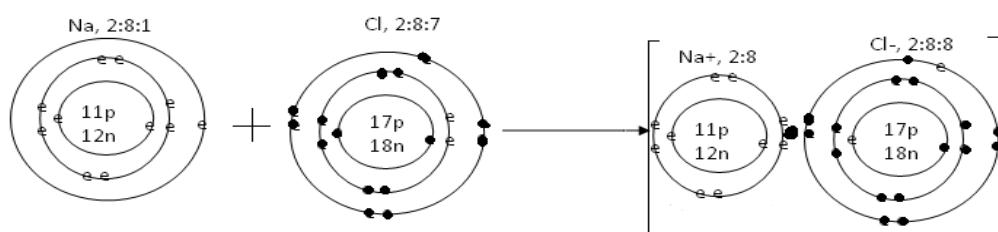
(b) **Negative ion (anions).** This ion is formed as a result of gain electrons by a non-metal atom. This is because non-metals have high tendency to accept electrons.



### Formation of ionic/electrovalent compounds

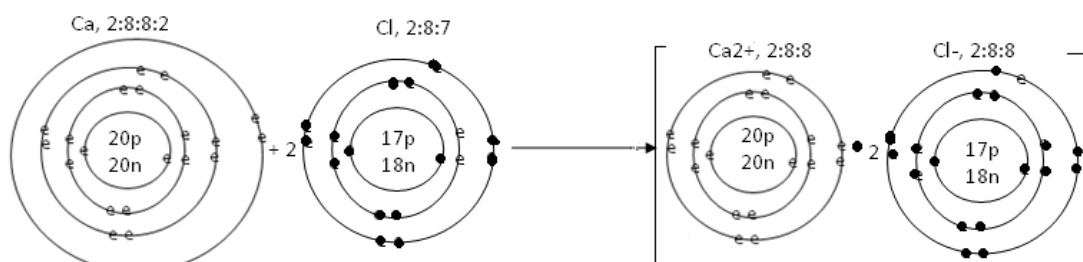
In the formation of ionic compounds, the number of electrons transferred by the metal is equal to its valency while the number of electrons received by the non-metal is also equal to the valency. The number of electrons which a metal atom can transfer to non-metal atom during the formation of an ionic compound is called **electrovalency**.eg the electrovalency of sodium is one, calcium is 2 and aluminium is three.

(a) **Formation of sodium chloride**

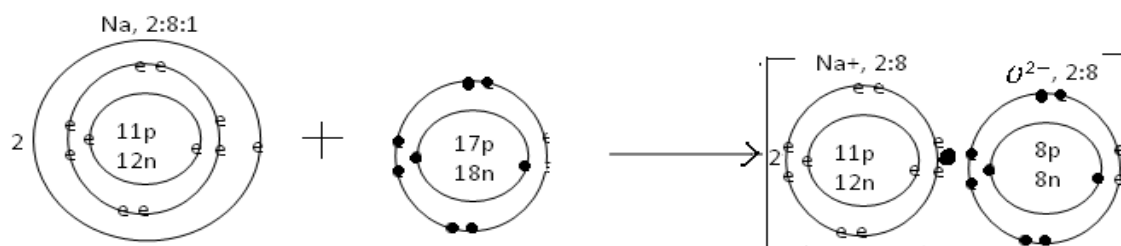


(b) **Formation of calcium oxide**

(c) **Formation of calcium chloride**



(d) **Formation of sodium oxide**



## Properties of ionic compounds

- They are very hard solids e.g. sodium chloride.
- They have high melting and boiling points because the ions in the compounds are held by very strong electrostatic force.
- They are very soluble in water but insoluble in organic solvents such as benzene, carbontetrachloride. The solubility in water is because water is polar. A polar substance is one which has charges.
- They conduct electricity when in molten or aqueous state.
- They are non-volatile i.e. they do not evaporate.

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## Covalent bonding

This is a type of bonding formed by sharing of electrons between two or more atoms. Each atom donates an electron or electrons towards the bond formation and the bond formed is called covalent bond. The atoms involved can be of the same type or different types. Compounds formed as a result of this bonding are referred to as covalent compounds.

## Formation covalent compounds

- Formation of chlorine
- Formation of hydrogen
- Formation of oxygen
- Formation of water
- Formation of hydrogen chloride
- Formation of ammonia
- Formation of hydrogen sulphide
- Formation of methane

## Properties of covalent compounds

- They exist mainly as gases, liquids and sometimes as soft solids e.g. candle wax, petroleum jelly, gases like carbon dioxide, oxygen and liquids like petrol, water, diesel.
- They have low melting and boiling points. This is because covalent bonds are weak and require very little heat to break.

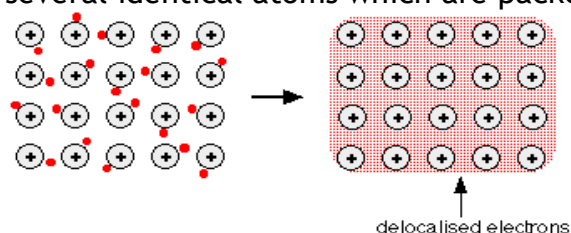
- They consist of molecules not ions and these molecules are held together by forces called van der Waals forces of attraction.
- They do not conduct electricity because they are made up of molecules not ions. In a solution, electric current is conducted by ion. Therefore ionic compounds conduct electricity because they contain free ions when in aqueous or molten state.

**NB//.** In metals electric current is conducted by electrons.

- They are insoluble in water but soluble in organic solvents (non-polar solvents) e.g. benzene, petrol, chloroform.

## Metallic bonding

This bond exists between metal atoms of the same kind. A piece of metal is made up of several identical atoms which are packed together in an orderly way.



Each atom constantly loses electrons equal to its valency and these electrons are free to move in any direction. No metal atom controls the movement of electrons therefore these electrons are referred to as delocalized electrons. Therefore metallic bonding is an attraction between the positively charged nucleus of the metal atom and the delocalized electrons.

The more the number of delocalized electrons released, the stronger the metallic bond and the stronger the metal e.g. aluminium is stronger than sodium because each aluminium atom releases three electrons while each sodium atom releases only one electron. This is why sodium metal is very soft and be cut using a knife.

The mobile electrons are responsible for the conductivity of electricity and heat by metals.

The strong metallic bond makes the metals to have very high melting and boiling points. The metal atoms can slide over each other when compressed. For this reason, they are said to be ductile and malleable i.e. they can be forged into various shapes without breaking.

## Structure of solids

Solids have three types of structure;

- (a) Giant ionic structure.
- (b) Giant atomic structure.
- (c) Simple molecular structure.

## Giant ionic structure

This structure is mainly for ionic compounds e.g. sodium chloride, magnesium chloride, copper (II) sulphate and potassium nitrate.



## Structure of sodium chloride

In this structure, each sodium ion is surrounded by six equidistant chloride ions likewise each chloride ion is surrounded by six equidistant sodium ions.

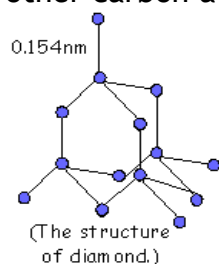
### Illustration

The above illustration is called the unit cell of sodium chloride structure.

## Giant atomic structure

This structure is mainly formed by diamond, graphite, silicondioxide, silicon. Consider diamond;

Diamond is a form of carbon. In the structure each carbon atom is surrounded by four other carbon atoms and this forms a tetrahedral structure as shown below;



In diamond, each carbon atom is bonded to one another by a strong covalent bond making diamond the hardest known substance. For this reason it can be used for drilling, cutting glasses etc

## Simple molecular structure

This structure is shown by covalent compounds only e.g. sulphur, carbondioxide, water etc. in this structure, the unit cell is the molecule. The molecules are joined together by weak van der waals forces. These forces are weak and can easily be broken. For this reason compounds with this structure have very low melting and boiling points.

## Periodic table

It was developed in order to put together elements with similar chemical properties. The first periodic table was developed by a Russian scientist called Mendeleev in the early 19<sup>th</sup> century. According to Mendeleev, he classified all the elements according to their mass numbers or atomic masses but this criterion failed to categorize elements with similar chemical properties together because mass number does not determine the chemical properties of an element.

However, in 1913, an English scientist called Mosley attempted to modify the periodic table of Mendeleev and he used atomic number to classify all the elements and all the elements with similar chemical properties found themselves together, and this periodic table was referred to as the modern periodic table.

The modern periodic table has a law called the periodic table law. It states that the chemical property of an element is function of its atomic number. This means that the chemical property of an element is determined by the atomic number which is equal to the number of electrons that element contains.

## The modern periodic table

The periodic table is divided into groups and periods of the elements.

- (a) **Groups.** These are vertical arrangement of elements into columns. These groups are indicated by roman numerals and the periodic table has 8 groups.
- (b) **Periods.** These are horizontal rows of elements in the periodic table. They are indicated by figures 1, 2 ,3 etc.

### Characteristics of elements in the same group

- Elements in the same group have the similar chemical and physical properties because they have the same valency.eg consider group 1 and group 2 elements.

#### Group 1 elements

Li, 2:1

Na, 2:8:1

K, 2:8:8:1

Rb, 2:8:8:8:1

Fr, 2:8:8:8:8:1

#### Group 2 elements

Be, 2:2

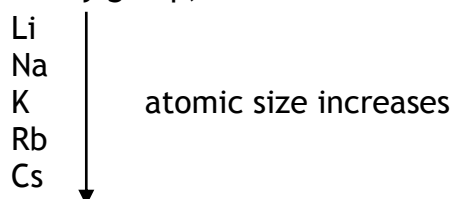
Mg, 2:8:2

Ca, 2:8:8:2

Sr, 2:8:8:8:2

Ba, 2:8:8:8:8:2

- Elements in the same group have the same number electrons in their outermost shell (valency).
- In any group, atomic size of the elements increases down the group.



In the above case, lithium is the smallest element in group 1 and caesium is the biggest.

- The reactivity of the elements in a group increases downwards.  
From the above, caesium is the most reactive and lithium is the least reactive.
- Sometimes elements in the different groups are known by different names eg
 

Group 1	alkali metals
Group 2	alkaline earth metals
Group 7	halogens
Group 8	noble/rare gases/inert gases
- The group number of an element corresponds to the number of electrons in its outermost shell. Eg

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### Example

- (a) An atom of element of X is represented by  ${}_{20}^{40}\text{X}$ .
- Write the electronic configuration of the element.  
2:8:8:2
  - State the group and period of the element.  
Group 2 because it has two electrons in the outermost shell.  
Period four because it has four shells/orbits occupied by electrons.
  - Write the formula of the chloride, oxide and nitrate of X.  
Chloride,  $\text{XCl}_2$   
Nitrate,  $\text{X}(\text{NO}_3)_2$   
Oxide,  $\text{XO}$
- (b) Element Y is represented as  ${}_{11}^{23}\text{Y}$
- Write the electronic configuration of Y.  
2:8:1
  - State the period and group of Y. Give reasons for your answers.  
Period three because it has three shells having electrons.  
Group 1 because it has two electrons in the last orbit/shell.
  - Write the formula of sulphate of Y.  
 $\text{Y}_2\text{SO}_4$

### Characteristics of elements in the same period

- Elements within the same period have the same number of shells occupied by electrons. eg all elements in period three have three shells occupied by electrons.

- The period number of an element corresponds to the number of shells occupied by electrons. eg oxygen has two shells occupied by electrons therefore it belongs to period 2 of the periodic table.
- The atomic size decreases from left to right along the period e.g. sodium is the biggest atom from period three.
- The reactivity of the elements decreases across the period as you move from left to right.
- Each period ends with a noble gas.

### Position of hydrogen

Hydrogen is placed in both group 1 and 7 of the periodic table. This shows a unique property of hydrogen.

- It is considered a group 1 because it has a valency of 1 and can form an ion with positive charge of 1.
- It is placed in group 7 because it reacts like group 7 elements e.g. sodium hydride is an ionic compound.

### Alkali metals (group 1 elements)

They have the following characteristics;

- They are the most reactive metals.
- They are called alkali metals because they produce hydroxides which are soluble in water e.g. sodium hydroxide, potassium hydroxide.
- They have a fixed valency of 1.
- All their compounds are white in colour and soluble in water.
- They are the most electropositive elements. Electropositivity refers to the ability of an element to lose electrons. However, electropositivity normally increases as you move down the group.

### Alkaline earth metals

They are called alkaline earth metals because they are the only metals found beneath the earth which are alkaline. Other metals found beneath the earth include gold, copper etc.

#### Characteristics

- They have a fixed valency of 2.
- They are reactive elements but less reactive than group 1 elements.

### Halogens (group 7 elements)

#### Characteristics

- They exist as diatomic molecules.
- They are non-metals and exist as gases at room temperature except iodine which exists as solid.
- They have a fixed valency of 1.
- They are the most reactive non-metals and very poisonous.

- These elements are said to be electronegative. This means the halogen atom has the ability to attract bonding electrons in a molecule towards itself. E.g. in hydrogen chloride, the chlorine atom will acquire negative charge and hydrogen positive charge.

## Noble gases

### Characteristics

- They have a valency of zero.
- They are monatomic gases (i.e. each molecule has one atom).
- They do not react because they have a valency of zero.

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## Transition elements

These are metals with special characteristics. Examples of transition elements include; copper, iron, nickel, manganese etc.

### Characteristic

- They have more than one valency while other elements have only one valency e.g. copper has 1 and 2, iron has 2 and 3.  
Copper (I) chloride, copper (II) chloride  
Iron (II) chloride, iron (III) chloride
- They form coloured compounds while other elements form compounds that are white eg  
Copper (II) sulphate -5-water is blue  
Iron (III) chloride is brown  
Copper (II) oxide is black  
Copper (II) nitrate is green
- They have catalytic ability i.e. most catalysts are transition metals/compounds.  
Manganese (IV) oxide for preparation of oxygen  
Iron used in Haber process to prepare ammonia  
Vanadium (V) oxide to prepare sulphuric acid by contact process

## Examples

Use part of the periodic table below to answer the questions that follow.

(I)

(VIII)

H	(II) (VII)	(III)	(IV)	(V)	(VI)	
A	Z				D	C
B						T
						Y

(a) (i) List elements which are non-metals.

A, B and Z

(ii) Give the most reactive metal.

B

(iii) State the least reactive element.

Y

(b) Give the formula of the compound formed between;

(i) A and D

A<sub>2</sub>D

(ii) B and C

BC

(iii) Z and D

ZD

(c) State the type of bond in the compounds in (b) above.

A<sub>2</sub>D electrovalent bond

BC electrovalent bond

ZD electrovalent bond

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### Determination of formula of compounds using empirical and molecular formula.

(a) **Empirical formula.** This is the simplest formula which expresses the ratio of each kind of atom present in a molecule. It does not give the actual number of each kind of atom present in that compound e.g. for glucose its empirical formula is CH<sub>2</sub>O.

(b) **Molecular formula.** Molecular formula of a compound is one which expresses the actual number of each kind of atom present in the compound. Sometimes certain compounds have the same empirical and molecular formula.

In general, empirical formula multiplied by whole number n gives the molar mass of the compound.

Compound	Empirical formula	Molecular formula
Water	H <sub>2</sub> O	H <sub>2</sub> O
Glucose	CH <sub>2</sub> O	C <sub>6</sub> H <sub>12</sub> O <sub>6</sub>
Ethanoic acid	CH <sub>2</sub> O	C <sub>2</sub> H <sub>4</sub> O <sub>2</sub>
Ethyne	CH	HC≡CH
Ethene	CH <sub>2</sub>	CH <sub>2</sub> =CH <sub>2</sub>

### Molar mass of a compound

This is the mass obtained by adding the mass of all the atoms present in the compound. The molecular formula is normally used to determine mass I the empirical formula. In order to determine the molar mass of a compound, the relative atomic mass of each of the atom must be given.

**Relative atomic mass** of an element is the mass of one atom of the element divided by  $\frac{1}{12}$ th the mass of one atom of carbon-12.

$$\text{Relative atomic mass} = \frac{\text{Mass of one atom of element}}{\left(\frac{1}{12}\right)^{\text{th}} \text{ mass of one atom of carbon} - 12}$$

Using the formula above, hydrogen was given a mass of  $1.0008 \cong 1$

Oxygen 16

Nitrogen 14

Relative atomic mass is a ratio therefore it does not have units.

### Determining molar mass given relative atomic masses

Find the molar masses of the following compounds.

- (a) Lead (II) nitrate. ( $Pb = 207, N = 14, O = 16$ )

$$Pb(NO_3)_2 = (207 \times 1) + (14 \times 2) + (63 \times 3 \times 2) = 331g/mol$$

- (b) Sodium carbonate. ( $Na = 23, C = 12, O = 16$ )

$$Na_2CO_3 = (23 \times 2) + (12 \times 1) + (16 \times 3) = 106g/mol$$

- (c) Sulphuric acid. ( $H = 1, S = 32, O = 16$ )

$$H_2SO_4 = (1 \times 2) + (32 \times 1) + (16 \times 4) = 98g/mol$$

- (d) Ammonium carbonate. ( $N = 14, H = 1, C = 12, O = 16$ )

$$(NH_4)_2CO_3 = (14 \times 2) + (1 \times 4 \times 2) + (12 \times 1) + (16 \times 3) = 96g/mol$$

## Exercise

Calculate the relative formula mass of the following compounds.

- Copper (II) sulphate pentahydrate
- Sodium hydrogen carbonate
- Ammonium chloride
- Calcium carbonate
- Iron (II) sulphate
- Methane

## Determining empirical and molecular formula of a compound

In order to determine the empirical formula of a compound, the percentage composition of the elements in the compound must be given and also the relative atomic masses of the elements should be given.

## Examples

- 1) A compound contains 20% by mass of hydrogen and 80% by mass of carbon.

Determine the empirical formula of the compound. ( $C = 12, H = 1$ )

Elements	C	H
Percentage composition	80	20
Number of moles	$\frac{80}{12}$	$\frac{20}{1}$
Simplest ratio	$\frac{6.7}{6.7}$	$\frac{20}{6.7}$
	1	3

Empirical formula is  $CH_3$

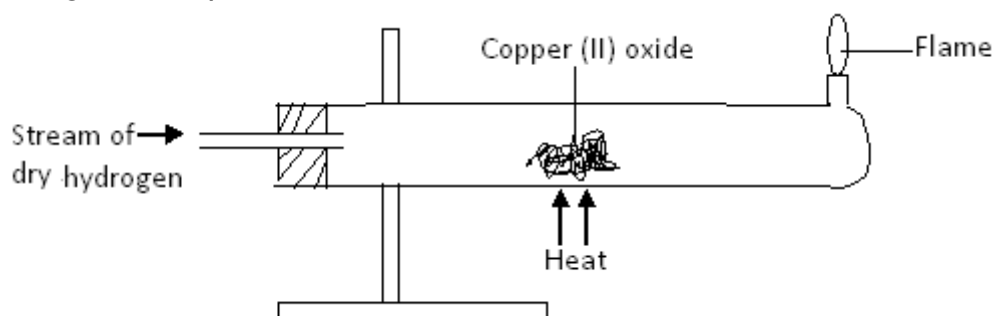
- 2) A compound contains 21.1% sodium, 40.5% sulphur and 30.4% oxygen. Find its empirical formula.

Elements	Na	S	O
Percentage composition	21.1	40.5	30.4
Number of moles	$\frac{21.1}{23}$	$\frac{40.5}{32}$	$\frac{30.4}{16}$
Simplest ratio	$\frac{0.97}{0.97}$	$\frac{1.3}{0.97}$	$\frac{1.9}{0.97}$
	1	1	2

Empirical formula is  $NaSO_2$

## Experiment to determine the formula of copper (II) oxide

Copper (II) oxide contains copper and oxygen chemically combined together. By removal of oxygen using a stream of hydrogen gas the masses of copper and oxygen can be determined hence the empirical formula is determined. The experiment is carried out using the setup below;



### Procedure

- Blow a hole in a test tube and weigh the test tube. Record its mass,  $m_1$  g.
- Put some copper (II) oxide in the test tube and weigh it again and record its mass,  $m_2$ g.
- Pass hydrogen gas through the test tube until the black copper (II) oxides turns brown. Immediately weigh the test tube together with the content and record its mass,  $m_3$ g. Excess hydrogen gas burns at the end of the tube as shown above.

### Calculations

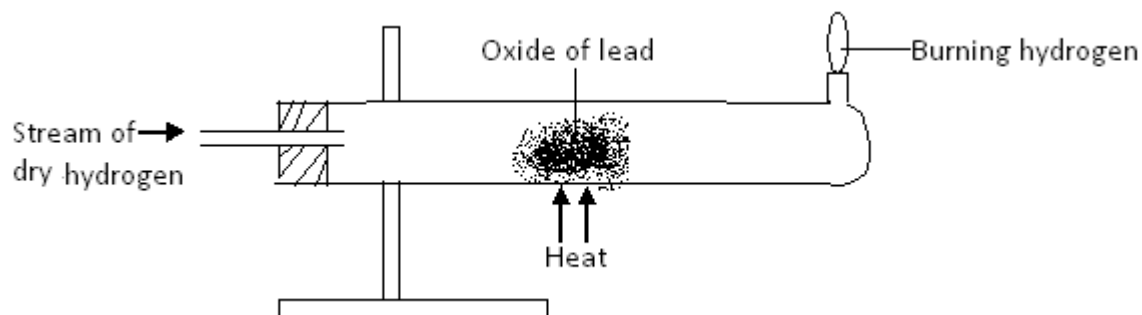
$$\text{Mass of copper} = (m_3 - m_1)g$$

$$\text{Mass of oxygen} = (m_2 - (m_3 - m_1))g$$

The empirical formula can be calculated using the normal approach.

### Examples

- 1) 1.60 g of the oxide of a metal M gave 1.2 g of the metal when reduced with hydrogen gas. Determine the empirical formula of the oxide. ( $M = 64, O = 16$ )
- 2) 15 g of an oxide of lead is strongly heated in a stream of hydrogen gas leaving 13 g of lead. Calculate the empirical formula of the oxide. ( $Pb = 207, O = 16$ )
- 3) 6.5 g of metal M combines with excess oxygen gas to give 8.1 g of the oxide. Calculate the simplest formula of the oxide. ( $M = 65, O = 16$ )
- 4) The figure below shows an apparatus used to determine the formula of oxide of lead.





Mass of tube = 15.647 g

Mass of tube + mass of oxide of lead = 15.886 g

Mass of tube + mass of lead = 15.85 g

(a) State what was observed in the tube as heating was continued.

**Dark brown solid turns to grey solid and colourless liquid that turns anhydrous copper (II) sulphate from white to blue condenses at the end.**

(b) Determine the empirical formula of the oxide of lead.

Mass of lead oxide = 15.886 - 15.647 = 0.239 g

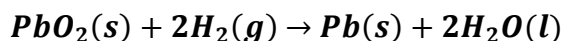
Mass of lead = 15.85 - 15.647 = 0.212 g

Mass of oxygen = 0.239 - 0.212 = 0.032 g

Elements	Pb	O
Percentage composition	0.207	0.032
Number of moles	$\frac{0.207}{207}$	$\frac{0.032}{16}$
Simplest ratio	$\frac{0.001}{0.001}$	$\frac{0.002}{0.002}$
	1	2

Empirical formula is PbO<sub>2</sub>

(c) Write a balanced equation of the reaction that took place in the tube.



(d) Why was hydrogen burnt at the jet after it has passed over lead oxide?

**This is to prevent excess hydrogen from reacting with atmospheric oxygen since this reaction can cause an explosion in the tube.**

(e) Why was hydrogen allowed to pass over lead until it was cold?

**To prevent lead from being deoxidized to lead oxide**

## Vapour density

Vapour density of a gas expresses the mass of a certain volume of the gas compared with mass of the same volume of hydrogen at the same temperature and pressure.

$$\text{Vapour density} = \frac{\text{Mass of 1 volume of gas}}{\text{Mass of 1 volume of hydrogen gas}}$$

Vapour density can be determined experimentally since it involves weighing equal volume of hydrogen gas and the gas.

By applying **Avogadro's law** which states that equal volumes of all gases at the same temperature and pressure contain the same number of molecules.

$$\text{Vapour density} = \frac{\text{Mass of 1 molecule of gas}}{\text{mass of 1 molecule of hydrogen gas}}$$

$$\text{Vapour density} = \frac{\text{Mass of 1 molecule of gas}}{\text{mass of 2 atoms of hydrogen}}$$

Therefore mass of 1 molecule of gas = vapour density  $\times$  2 = relative molecular mass

## Examples

- 1) A hydrocarbon has a vapour density of 15 and contains 70% hydrogen. Determine the empirical formula and hence the molecular formula of the hydrocarbon.
- 2) Calculate the molecular formula of the compound which has the following composition.  
(4.8%Mg, 13%S, 26%O, 51.2%water of crystallisation)

## Calculating percentage composition of an element given the molecular formula of the compound

In this calculation, one is required to work out the molar mass hence the percentage composition can be determined.

### Examples

- 1) Determine the percentage composition of elements present in sodium carbonate.

$$\text{Molar mass of Na}_2\text{CO}_3 = (23 \times 2) + (12 \times 1) + (16 \times 3) = 106 \text{ g/mol}$$

$$\text{Mass of Na} = (23 \times 2) = 46 \text{ g}$$

$$\text{Percentage of Na} = \frac{46}{106} \times 100\% = 43.4\%$$

$$\text{Percentage of C} = \frac{12}{106} \times 100\% = 11.32\%$$

$$\text{Percentage of O} = \frac{(16 \times 3)}{106} \times 100\% = 45.3\%$$

- 2) Calculate the percentage of carbon in ethene.

$$\text{Molar mass of CH}_2 = \text{CH}_2 = 12 + 2 + 12 + 2 = 28 \text{ g}$$

$$\text{Mass of carbon} = (12 \times 2) = 24 \text{ g}$$

$$\text{Percentage of C} = \frac{24}{28} \times 100\% = 85.7\%$$

- 3) Determine the percentage of water of crystallization in copper (II) sulphate pentahydrate.

$$\text{Molar mass of CuSO}_4 \cdot 5\text{H}_2\text{O} = (64 \times 1) + (32 \times 1) + (16 \times 4) + 5(2 + 16) = 250 \text{ g}$$

$$\text{Mass of water of crystallisation} = 5(2 + 16) = 90 \text{ g}$$

$$\text{Percentage of water of crystallisation} = \frac{90}{250} \times 100\% = 36\%$$

- 4) Calculate the percentage of nitrogen in lead (II) nitrate.

- 5) Determine the percentage of sulphate in sulphuric acid.

### Mole concept

A mole is the amount of substance that contain as many elementally particles as they are in carbon-12 isotope. These particles can be atoms, molecules, electrons or ions.

A mole can be calculated from the formula;

$$\text{Number of moles} = \frac{\text{given mass}}{\text{molar mass}}$$

In order to calculate the number of moles of a given substance whose mass is known, the molar mass of the substance must first be worked out.

### Examples

- 1) Find the number of moles of sodium hydroxide contained in 10 g. ( $\text{Na} = 23, \text{O} = 16, \text{H} = 1$ )

$$\text{Molar mass of NaOH} = 23 + 16 + 1 = 40 \text{ g}$$

$$\text{Number of moles} = \frac{10}{40} = 0.25 \text{ moles}$$

- 2) Find the number of moles of sulphuric acid contained in 9.8 g. ( $\text{H} = 1, \text{S} = 32, \text{O} = 16$ )

- 3) Calculate the number of moles of calcium carbonate contained in 10 g. ( $Ca = 40, C = 12, O = 16$ )
- 4) Determine the number of moles of sodium carbonate contained in 10.6 g. ( $Na = 23, O = 16, C = 12$ )

One mole of any compound of a substance contains  $6.03 \times 10^{23}$  particles. The number  $6.03 \times 10^{23}$  is referred to as the **Avogadro's constant** denoted by letter L.

### Molar mass

This is the mass of one mole of any given substance e.g.

- 1 mole of  $H_2O = 18$  g
- 1 mole of  $NaOH = 40$  g
- 1 mole of  $CaCO_3 = 100$  g
- 1 mole of  $H_2SO_4 = 98$  g
- 1 mole of  $Mg = 24$  g

### Gas laws

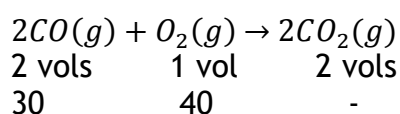
These are laws that explain the behavior of gases;

- 1) **Gay-Lussac's law of combining volumes.** It states that when gases react, they do so in volumes which bear a simple ratio to one another and the volume of gaseous products, provided all volumes are measured at the same temperature and pressure. eg
  - 1 volume of nitrogen gas combines with 3 volumes of hydrogen gas to give two volumes of ammonia gas.
 
$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$
  - 2 volumes of nitrogendioxide break up to form 1 volume of nitrogen gas and 2 volumes of oxygen gas.
 
$$2NO_2(g) \rightarrow N_2(g) + 2O_2(g)$$
  - 1 volume of hydrogen gas combines with 1 volume of chlorine gas to give 2 volumes of hydrogen chloride gas.
 
$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

### Examples

- (1) 40 cm<sup>3</sup> of oxygen gas was to 30 cm<sup>3</sup> of carbonmonoxide and the mixture ignited to produce carbondioxide. Calculate the volume and composition of the reacting mixture.

#### Solution



Volume of mixture = volume of unreacted oxygen gas + volume of carbondioxide formed

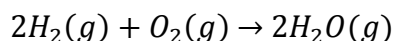
Volume of unreacted oxygen gas =  $(40 - 15) = 25 \text{ cm}^3$

Volume of carbondioxide formed =  $30 \text{ cm}^3$

Total volume =  $30 + 25 = 55 \text{ cm}^3$

(2)  $100 \text{ cm}^3$  of hydrogen gas was ignited with  $30 \text{ cm}^3$  of oxygen to form steam.

a) Write equation for the reaction.



b) Calculate the volume composition of the resulting mixture.

### Solution

Volume of mixture = volume of unreacted hydrogen gas + volume of steam formed

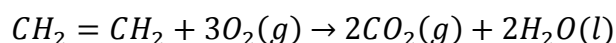
Volume of unreacted hydrogen gas =  $(100 - 60) = 40 \text{ cm}^3$

Volume of steam formed =  $60 \text{ cm}^3$

Total volume =  $40 + 60 = 100 \text{ cm}^3$

### Exercise

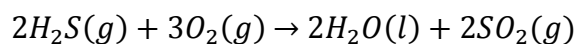
(1) Ethene reacts with oxygen according to the equation below.



If  $15 \text{ cm}^3$  of ethene were mixed with  $60 \text{ cm}^3$  of oxygen and the mixture ignited.

Find the composition of the mixture after the reaction. (Ans  $75 \text{ cm}^3$ )

(2) Hydrogen sulphide reacts with oxygen according to the equation.



If  $30 \text{ cm}^3$  of hydrogen sulphide and  $30 \text{ cm}^3$  of oxygen gas were burnt in a closed container, calculate the volume of sulphurdioxide gas produced. (Ans  $50 \text{ cm}^3$ )

2) **Boyle's law.** It states that the volume of a given mass of a gas is inversely proportional to its pressure at a constant temperature i.e.  $V \propto \frac{1}{P}$ . Therefore  $V = \frac{K}{P}$  where K is the constant of proportionality. When the pressure is increased, the volume of the gas decreases likewise when the pressure is decreased, the volume of the gas increases.

Assuming  $V_1$  is the volume of the gas measured at pressure  $P_1$ . This volume  $V_1$  can be converted to volume  $V_2$  at any other pressure  $P_2$  provided temperature of the system is kept constant.

$$P_1V_1 = P_2V_2$$

3) **Charles's law.** The law states that the volume of a given mass of a gas is directly proportional to absolute temperature provided pressure is kept constant i.e.  $V \propto T$ . Therefore  $V = KT$

or  $\frac{V}{T} = K$  Where K is the constant of proportionality.

This means that if the volume of the gas is doubled, temperature also doubles provided its pressure is constant. Assuming  $V_1$  is the volume measured at  $T_1$ , this volume can be converted to  $V_2$  at any other absolute temperature  $T_2$ .

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

By combining Charles's and Boyle's laws we obtain the ideal gas equation;

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

### Standard temperature and pressure

The standard temperature and pressure are chosen at  $0^\circ\text{C} = 273\text{K}$ ,  $760\text{ mmHg} = 1\text{ atm} = 101325\text{ pa}$ .

### Examples

- 1) A certain mass of gas occupies  $211\text{ cm}^3$  at  $18^\circ\text{C}$  and  $740\text{ mmHg}$ . What volume will this gas occupy at  $20^\circ\text{C}$  and  $770\text{ mmHg}$ .

#### Solution

$P_1 = 740\text{ mmHg}$ ,  $V_1 = 211\text{ cm}^3$ ,  $T_1 = 18^\circ\text{C} = 291\text{ K}$

$P_2 = 770\text{ mmHg}$ ,  $V_2 = ?$ ,  $T_2 = 20^\circ\text{C} = 293\text{ K}$

By applying the ideal gas equation,

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$\frac{740 \times 211}{291} = \frac{770 \times V_2}{293} \text{ Therefore } V_2 = 204\text{ cm}^3$$

- 2) A certain mass of gas occupies  $146\text{ cm}^3$  at  $18^\circ\text{C}$  and  $738\text{ mmHg}$ . Calculate the volume of this gas at s.t.p.

#### Solution

$P_1 = 738\text{ mmHg}$ ,  $V_1 = 146\text{ cm}^3$ ,  $T_1 = 18^\circ\text{C} = 291\text{ K}$

$P_2 = 760\text{ mmHg}$ ,  $V_2 = ?$ ,  $T_2 = 273\text{ K}$

By applying the ideal gas equation,

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$\frac{738 \times 146}{291} = \frac{760 \times V_2}{273} \text{ Therefore } V_2 = 370.3\text{ cm}^3$$

- 3) A certain gas occupies  $360\text{ cm}^3$  at  $0^\circ\text{C}$  and  $700\text{ mmHg}$ . Find the volume of the gas at s.t.p.

#### Solution

$P_1 = 700\text{ mmHg}$ ,  $V_1 = 360\text{ cm}^3$ ,  $T_1 = 0^\circ\text{C} = 273\text{ K}$

$P_2 = 760\text{ mmHg}$ ,  $V_2 = ?$ ,  $T_2 = 273\text{ K}$

By applying the ideal gas equation,

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$\frac{700 \times 360}{273} = \frac{760 \times V_2}{273} \text{ Therefore } V_2 = 280.25\text{ cm}^3$$

## Calculations of gas volume from equation

### Molar gas volume

This refers to the volume occupied by one mole of a gas at standard temperature and pressure. This volume is 22.4 litres/22.4 dm<sup>3</sup>/22400 cm<sup>3</sup>. This is true for all gases so long only one mole is considered.

1 mole of oxygen gas occupies 22.4 litres

1 mole of nitrogen gas occupies 22.4 litres

1 mole of carbon dioxide gas occupies 22.4 litres

1 mole chlorine gas occupies 22.4 litres

1 mole of hydrogen gas occupies 22.4 litres

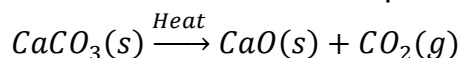
70

### Mass-Volume relationship

In this calculation, one is required to first write a balanced equation if not given. The volume of a gas occupied at s.t.p is given and this is normally 22.4 litres/22.4 dm<sup>3</sup>/22400 cm<sup>3</sup>. The atomic masses of the elements involved in the equation are given. You are therefore required to work out the molar mass or relative molecular mass of the compound and this is normally reactant and relate this to the volume of the gas produced.

### Examples

- 1) Calcium carbonate decomposes according to the equation below.



What volume of carbon dioxide measured at s.t.p will be produced when 75 g of calcium carbonate is completely decomposed. ( $\text{Ca} = 40, \text{C} = 12, \text{O} = 16$ , 1 mole of gas at s.t.p occupies 22400 cm<sup>3</sup>)

#### Solution

Molar mass of  $\text{CaCO}_3 = 40 + 12 + (16 \times 3) = 100 \text{ g/mol}$

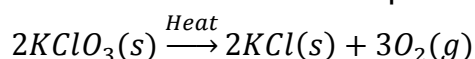
From the reaction equation,

1 mole of  $\text{CaCO}_3$  produces 1 mole of  $\text{CO}_2$  at s.t.p

100 g of  $\text{CaCO}_3$  produces 22400 cm<sup>3</sup> of  $\text{CO}_2$

Therefore 75 g of  $\text{CaCO}_3$  produces  $\frac{75 \times 22400}{100} = 16800 \text{ cm}^3$  of  $\text{CO}_2$

- 2) Potassium chlorate decomposes according to the equation below.



What volume of oxygen will be produced by heating 5 g of potassium chlorate at s.t.p?

( $\text{K} = 39, \text{Cl} = 35.5, \text{O} = 16$ , 1 mole of gas at s.t.p occupies 22.4 dm<sup>3</sup>)

#### Solution

Molar mass of  $\text{KClO}_3 = 39 + 35.5 + (16 \times 3) = 122.5 \text{ g/mol}$

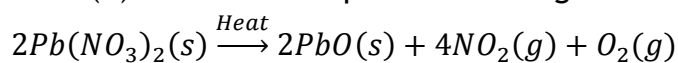
From the reaction equation,

2 moles of  $KClO_3$  produce 3 moles of  $O_2$  at s.t.p

$2 \times 125.5$  g of  $KClO_3$  produce  $3 \times 22.4$  dm<sup>3</sup> of  $O_2$

Therefore 5 g of  $KClO_3$  produces  $\frac{5 \times 3 \times 22.4}{2 \times 125.5} = 1.3714$  dm<sup>3</sup> of  $O_2$

- 3) Lead (II) nitrate decomposes according to the equation.



What volume of nitrogendioxide would be produced at s.t.p when 50 g of lead (II) nitrate is completely decomposed by heating? ( $Pb = 207, O = 16, N = 14$ , 1 mole of gas at s.t.p occupies 22.4 dm<sup>3</sup>)

#### Solution

Molar mass of  $Pb(NO_3)_2 = 207 + (14 \times 2) + (16 \times 3 \times 2) = 331$  g/mol

From the reaction equation,

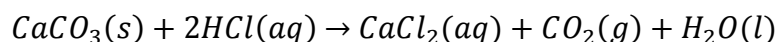
2 moles of  $Pb(NO_3)_2$  produce 4 moles of  $NO_2$  at s.t.p

$2 \times 331$  g of  $Pb(NO_3)_2$  produce  $4 \times 22.4$  dm<sup>3</sup> of  $NO_2$

Therefore 50 g of  $Pb(NO_3)_2$  produces  $\frac{50 \times 4 \times 22.4}{2 \times 331} = 6.7674$  dm<sup>3</sup> of  $NO_2$

- 4) What volume of carbondioxide measured at s.t.p can be obtained by dissolving 50 g of calcium carbonate in dilute hydrochloric acid. ( $Ca = 40, C = 12, O = 16$ , 1 mole of gas at s.t.p occupies 22.4 litres)

#### Solution



Molar mass of  $CaCO_3 = 40 + 12 + (16 \times 3) = 100$  g/mol

From the reaction equation,

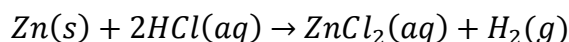
1 mole of  $CaCO_3$  produces 1 mole of  $CO_2$  at s.t.p

100 g of  $CaCO_3$  produces 22.400 litres of  $CO_2$

Therefore 50 g of  $CaCO_3$  produces  $\frac{50 \times 22.4}{100} = 11.2$  litres of  $CO_2$

- 5) Calculate the volume of hydrogen gas produced at s.t.p by reacting 15.2 g of zinc dust with dilute hydrochloric acid. ( $Zn = 65$ , 1 mole of gas occupies 22400 cm<sup>3</sup> at s.t.p)

#### Solution



Molar mass of  $Zn = 65$  g/mol

From the reaction equation,

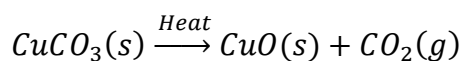
1 mole of  $Zn$  produces 1 mole of  $H_2$  at s.t.p

65 g of  $Zn$  produces 22400 cm<sup>3</sup> of  $H_2$

Therefore 15.2 g of  $Zn$  produces  $\frac{15.2 \times 22400}{65} = 5238.15$  cm<sup>3</sup> of  $H_2$

- 6) 16.0 g of copper (II) carbonate was strongly heated until no more change. Find the volume of carbondioxide produced at room temperature. ( $Cu = 64, C = 12, O = 16$ , 1 mole of gas at r.t occupies 24000  $cm^3$ )

#### Solution



Molar mass of  $CuCO_3 = 64 + 12 + (16 \times 3) = 124 g/mol$

From the reaction equation,

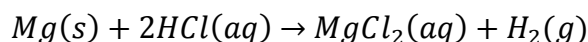
1 mole of  $CuCO_3$  produces 1 mole of  $CO_2$  at r.t

124 g of  $CuCO_3$  produces 24000  $cm^3$  of  $CO_2$

Therefore 16 g of  $CuCO_3$  produces  $\frac{16 \times 24000}{124} = 3096.774 cm^3$  of  $CO_2$

- 7) 26.5 g of dilute hydrochloric acid was reacted with magnesium turnings. Find the volume of hydrogen gas formed at room temperature. ( $H = 1, Cl = 35.5, O$ , 1 mole of gas at r.t occupies 24 litres)

#### Solution



Molar mass of  $HCl = (1 + 35.5) = 36.5 g/mol$

From the reaction equation,

2 moles of  $HCl$  produce 1 mole of  $H_2$  at s.t.p

$2 \times 36.5$  g of  $HCl$  produce 24 litres of  $H_2$

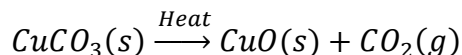
Therefore 26.5 g of  $Zn$  produces  $\frac{26.5 \times 24}{2 \times 36.5} = 8.7123 litres$  of  $H_2$

### Mass-mass relationship

#### Examples

- 1) Calculate the mass of residue formed if 13 g of copper (III) carbonate is strongly heated until no further change. ( $Cu = 64, C = 12, O = 16$ )

#### Solution



Molar mass of  $CuCO_3 = 64 + 12 + (16 \times 3) = 124 g/mol$

Molar mass of  $CuO = 64 + 16 = 80 g/mol$

1 mole of  $CuCO_3$  produces 1 mole of  $CuO$

124 g of  $CuCO_3$  produce 80 g of  $CuO$

13 g of  $CuCO_3$  produce  $\frac{13 \times 80}{124} = 8.3871 g$  of  $CuO$

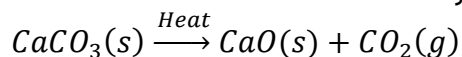
Therefore the mass of the residue is 8.3971 g

- 2) Calculate the loss in mass when 10 g of calcium carbonate is strongly heated. ( $Ca = 40, C = 12, O = 16$ )

#### Solution



**NB//.** The loss in mass is usually the mass of the gas produced.



Molar mass of  $\text{CaCO}_3 = 40 + 12 + (16 \times 3) = 100 \text{ g/mol}$

Molar mass of  $\text{CO}_2 = (12 + 16 \times 2) = 44 \text{ g/mol}$

1 mole of  $\text{CaCO}_3$  produces 1 mole of  $\text{CO}_2$

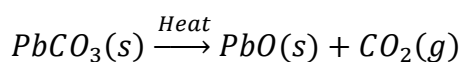
100 g of  $\text{CaCO}_3$  produce 44 g of  $\text{CO}_2$

10 g of  $\text{CaCO}_3$  produce  $\frac{10 \times 44}{100} = 4.4 \text{ g of CO}_2$

Therefore the loss in mass is 4.4 g

- 3) Calculate the mass of lead (II) oxide formed when 60 g of lead (II) carbonate is strongly heated. ( $\text{Pb} = 207, \text{C} = 12, \text{O} = 16$ )

#### Solution



Molar mass of  $\text{PbCO}_3 = 207 + 12 + (16 \times 3) = 267 \text{ g/mol}$

Molar mass of  $\text{PbO} = 207 + 16 = 223 \text{ g/mol}$

1 mole of  $\text{PbCO}_3$  produces 1 mole of  $\text{PbO}$

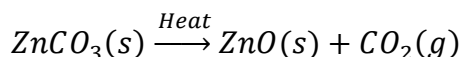
267 g of  $\text{PbCO}_3$  produce 223 g of  $\text{PbO}$

60 g of  $\text{PbCO}_3$  produce  $\frac{60 \times 223}{267} = 50.1124 \text{ g of PbO}$

Therefore the mass of  $\text{PbO}$  is 50.1124 g

- 4) Calculate the mass of zinc carbonate required to produce 75 g of zinc oxide when zinc carbonate is heated strongly. ( $\text{Zn} = 65, \text{C} = 12, \text{O} = 16$ )

#### Solution



Molar mass of  $\text{ZnCO}_3 = 65 + 12 + (16 \times 3) = 125 \text{ g/mol}$

Molar mass of  $\text{ZnO} = 65 + 16 = 81 \text{ g/mol}$

1 mole of  $\text{ZnCO}_3$  produces 1 mole of  $\text{ZnO}$

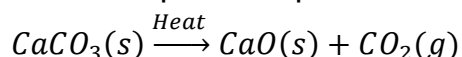
125 g of  $\text{ZnCO}_3$  produce 81 g of  $\text{ZnO}$

$\frac{75 \times 125}{81}$  g of  $\text{ZnCO}_3$  produce 75 g of  $\text{ZnO}$

Therefore the mass of  $\text{ZnCO}_3$  is 115.7407 g

#### Exercise

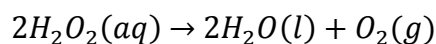
- 1) Marble chips decompose according to the equation.



(a) What volume of carbondioxide is produced at s.t.p when 10 g of marble chips are completely decomposed? (Ans 2.24 dm<sup>3</sup>)

(b) What is the mass of quick lime formed? ( $\text{Ca} = 40, \text{C} = 12, \text{O} = 16$ ) (Ans 5.6 g)

- 2) Hydrogen peroxide decomposes according to the equation below.



- (a) Calculate the volume of oxygen gas formed when 20 g of hydrogen peroxide decomposes at s.t.p. (Ans 6.5882 dm<sup>3</sup>)
- (b) Calculate the mass of water formed when 60 g of hydrogen peroxide decomposes.  
( $H = 1, O = 16$ ) (Ans 31.76 g)
- 3) The action of heat on copper (II) carbonate is represented by the equation below;  

$$CuCO_3(s) \xrightarrow{\text{Heat}} CuO(s) + CO_2(g)$$
 If 24.8 g of copper (II) carbonate is decomposed, what will be the;  
 (a) mass of copper (II) oxide formed. (Ans 16 g)  
 (b) volume of carbondioxide measured at room temperature. (Ans 4.48 dm<sup>3</sup>)  
 ( $Cu = 64, C = 12, O = 16$ , 1 mole of gas at r.t occupies 24000 cm<sup>3</sup>)
- 4) Magnesium reacts with hydrochloric acid according to the equation below;  

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$
 Calculate the mass of magnesium chloride formed when 2.4 g of magnesium is reacted completely with hydrochloric acid. ( $Mg = 24, Cl = 35.5$ )
- 5) (a) Write a balanced equation for the reaction between marble chips and dilute hydrochloric acid.  
 (b) Calculate the mass of calcium chloride produced when 60 g of marble chips is completely reacted with hydrochloric acid. ( $Ca = 40, C = 12, O = 16, Cl = 35.5$ )

## ELECTROLYSIS

This is the decomposition of an electrolyte by passing an electric current through it. It can also be defined as the decomposition of substances into either aqueous state or molten state by passing electric current through it.

An electric current can be carried by either electrons or free ions.

In metals electric current is considered to be the flow of electrons while in a solution, electric current is carried by free ions.

When an ionic solid is dissolved in water or heated, the ions in it become free and therefore can easily carry electric current. Give some examples;

Qn **Solid sodium chloride does not conduct electricity while molten sodium chloride conducts electricity. Explain.**

### Difference between metallic conduction and electrolysis

Metallic conduction	Electrolysis
<ul style="list-style-type: none"> <li>✓ Current is carried by flow of electrons</li> <li>✓ Accompanied by a physical change e.g. heating of the metal</li> <li>✓ Conduction of electricity decreases with increase in temperature</li> </ul>	<ul style="list-style-type: none"> <li>✓ Current is carried by flow of ions</li> <li>✓ Accompanied by a chemical change e.g. liberation of gas, deposition of metals.</li> <li>✓ Conduction increases with increase in temperature</li> </ul>

### Electrolytes

This is a substance which when in molten state or aqueous state conducts electric current and decomposed by it. The electrolytes are ionic compounds.

Examples of electrolytes include;

Electrolytes	Formula	ions
<ul style="list-style-type: none"> <li>✓ Sulphuric acid</li> <li>✓ Hydrochloric acid</li> <li>✓ Sodium hydroxide</li> <li>✓ Copper (II) sulphate</li> <li>✓ Lead (II) bromide</li> <li>✓ Copper (II) chloride</li> <li>✓ Lead (II) nitrate</li> <li>✓ Water</li> <li>✓ Sodium chloride</li> </ul>		

They cannot conduct electric current when in the solid state because the ions in them are joined firmly by strong electrostatic force of attraction. These ions can be set free and become mobile when the solid is dissolved in water or by simply heating the solid.

Covalent compounds do not conduct electric current because they do not contain ions hence they are made up of molecules which cannot conduct electric current. For this reason, they are called non electrolytes for example;

- ✓ Carbon dioxide
- ✓ Starch
- ✓ Ammonia
- ✓ Benzene
- ✓ Carbon tetrachloride

### Types of electrolytes

There are two types of electrolytes;

- (i) **Strong electrolytes.** These electrolytes ionize completely when in solution to giving many free ions hence making it highly conducting.

Examples include

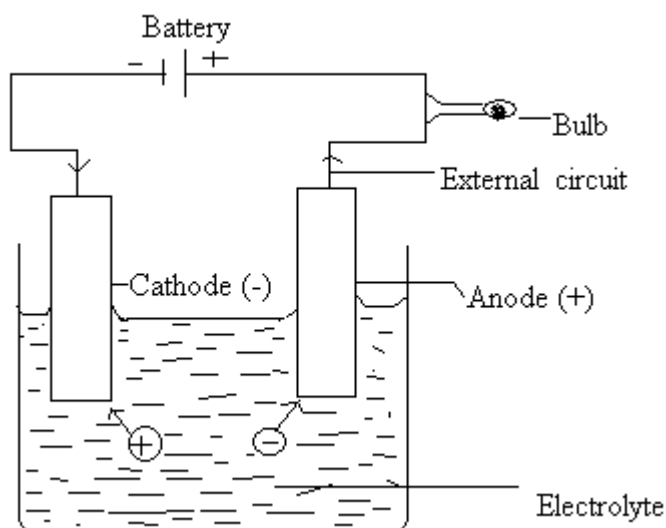
- ✓ Sulphuric acid
- ✓ Nitric acid
- ✓ Copper (II) sulphate
- ✓ Sodium chloride e.t.c

- (ii) **Weak electrolytes.** These are electrolytes which ionize partially to produce very few ions in solution resulting into low conductivity in solution.

Examples include;

- ✓ Water
- ✓ Ethanoic acid
- ✓ Ammonia e.t.c

### Electrolytic cell



An electrolytic cell is a setup used to carry out the process of electrolysis in the laboratory. It is made up of the following;

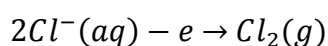
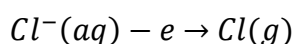
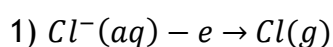
- (i) **Electrodes.** An electrode is a terminal or a pole through which electrons enter or leave the electrolyte. The electrode can be made of carbon (graphite) or a metal. There are two types of electrodes.
  - **Cathode.** This is a negative electrode through which electrons enter an electrolyte. In the electrolytic cell, it is found on the part of the battery with a short line to mean negative.
  - **Anode.** This is the positive electrode through which electrons leave the electrolyte. It is on the side represented by a long line as source of current to mean positive.
- (ii) **Electrolyte.** This is the solution through which electric current passes through during electrolysis. It can either be in the aqueous state or molten state.
- (iii) **External wire.** It connects the anode and the cathode and electrodes produced at the cathode move to the cathode via the external wire. Connected to the external wire is the bulb which is used to indicate whether the process of electrolysis is taking place. This is shown when the bulb gives out light and when there is no light; it means that electrolysis is not taking place.
- (iv) **Battery (direct current source).** The D.C provides a force which propels the electrons from the anode to cathode via the external wire.

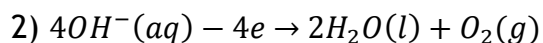
### Ionic theory

This theory was put forward to by Arrhenius in 1880 to explain how chemical compounds conduct electric current in solution.

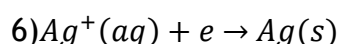
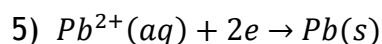
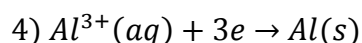
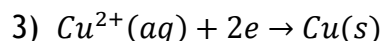
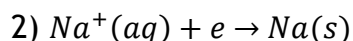
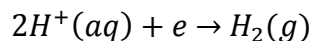
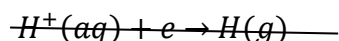
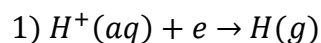
It states that electrolytes are made of ions which are negatively and positively atoms or radicals.

- The negative ions are called anions and during the process of electrolysis, they move to the anode where they are discharged by loss of electrons (oxidation).





- The positive ions are called cations and they migrate to the cathode where they are discharged by gain of electrons (reduction).



**NB//.** Atomicity is the number of atoms present in one molecule of an element. For metals, their atomicity is one while for non metals they have varying atomicity.

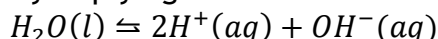
Molecule	Atomicity
✓ Oxygen	✓ 2
✓ Nitrogen	✓ 2
✓ Chlorine	✓ 2
✓ Hydrogen	✓ 2
✓ Sulphur	✓ 1
✓ Bromine	✓ 2

### Role of water in electrolysis

Water is a universal solvent since it can dissolve many chemical compounds. It acts as a universal solvent because;

- It is polar and has high dielectric constant (insulating capacity of solvent)

Water is regarded as a weak electrolyte it poorly conducts electric current since it ionizes partially implying that the ions can recombine to form molecular water.



**NB//.** During electrolysis, the ions that come from water must always be indicated because they also participate in the conduction electricity. For example in the electrolysis of dilute sodium hydroxide solution; the ions present are  $H^+$ ,  $OH^-$ ,  $Na^+$ . But the ions by water are very few because water is a weak electrolyte while the ions produced sodium hydroxide are many because sodium hydroxide is a strong electrolyte. Therefore the conductivity of electric current is majorly due to the sodium hydroxide.

### Factors that determine the discharge of a particular ion during electrolysis

- Position of the ion in the electrochemical series.** This is applied when dealing with dilute solutions. The ion selected for discharge depends on its position in the electrochemical series. The lower the ion in the electrochemical series, the more readily it is discharged since the energy required for discharge decreases down the electrochemical series.

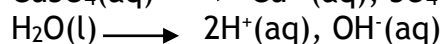
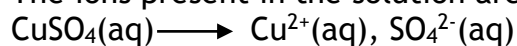
#### Electrochemical series for ions

Cations	Anions
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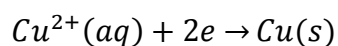
<ul style="list-style-type: none"> <li>▪ <math>K^+</math></li> <li>▪ <math>Na^+</math></li> <li>▪ <math>Ca^{2+}</math></li> <li>▪ <math>Mg^{2+}</math></li> <li>▪ <math>Al^{3+}</math></li> <li>▪ <math>Fe^{2+}</math></li> <li>▪ <math>Sn^{2+}</math></li> <li>▪ <math>Pb^{2+}</math></li> <li>▪ <math>H^+</math></li> <li>▪ <math>Cu^{2+}</math></li> <li>▪ <math>Hg^+</math></li> <li>▪ <math>Ag^+</math></li> <li>▪ <math>Au^+</math></li> </ul>	<p>Ease of discharge ions increases downwards</p>	<ul style="list-style-type: none"> <li>▪ <math>NO_3^-</math></li> <li>▪ <math>SO_4^{2-}</math></li> <li>▪ <math>Cl^-</math></li> <li>▪ <math>Br^-</math></li> <li>▪ <math>I^-</math></li> <li>▪ <math>OH^-</math></li> </ul>
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Consider the electrolysis of solution of copper (II) sulphate using platinum electrodes.

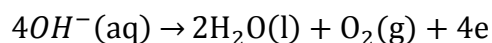
The ions present in the solution are;



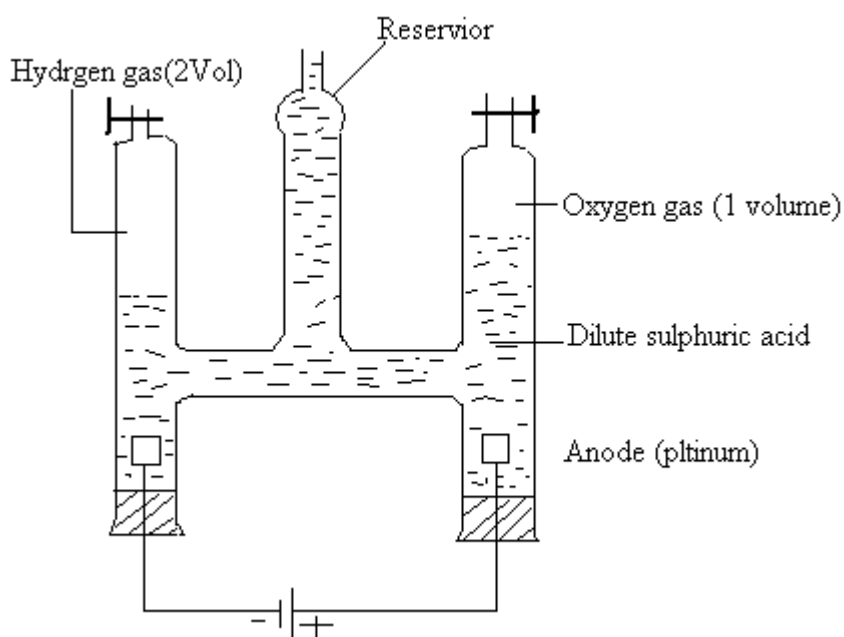
- **Cathode.** Both  $Cu^{2+}(aq)$  and  $2H^+(aq)$  migrate to the cathode but  $Cu^{2+}(aq)$  is selected for discharge because it is lower than  $H^+(aq)$  in the electrochemical series.



- **Anode.** Both  $SO_4^{2-}$  and  $OH^-$  migrate to the anode but is the  $OH^-$  that is selected for discharge since it is lower than  $SO_4^{2-}$  in the electrochemical series.



### Electrolysis of dilute sulphuric acid

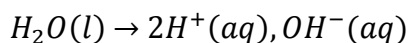
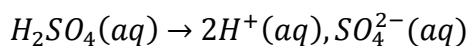


### Procedure

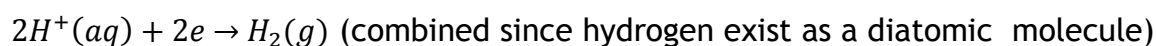
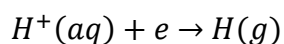
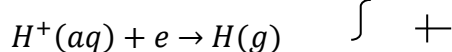
- The taps of the burettes are opened and filled with dilute sulphuric acid through the reservoir until the burettes are full. Immediately the taps are then closed.
- The circuit is then completed by connecting the cathode and the anode via an external wire as shown above. Immediately the process of electrolysis begins.

## Observations

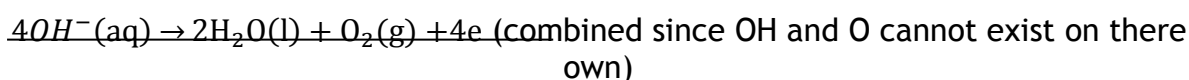
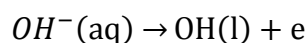
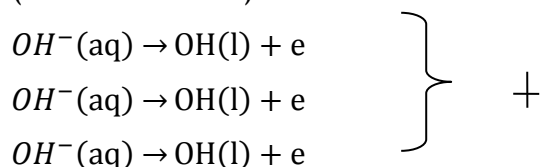
The ions present in solution are produced by sulphuric acid and water;



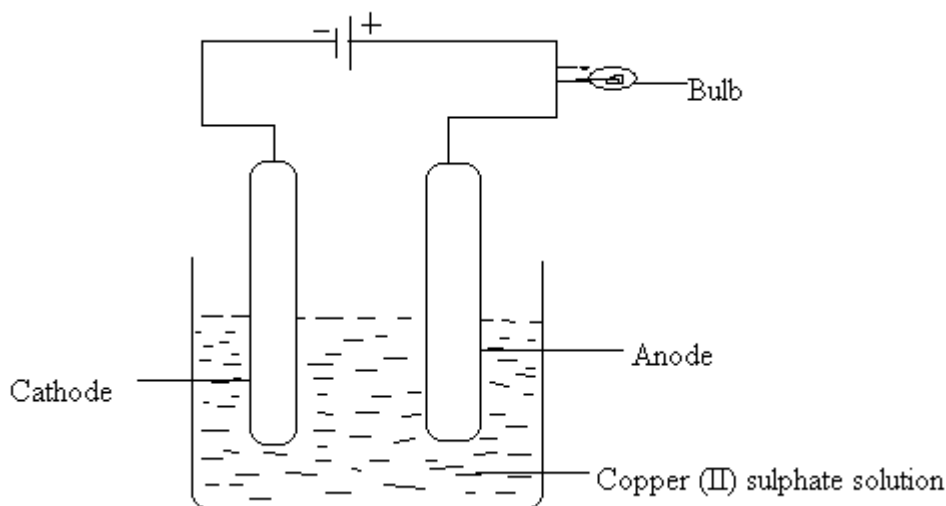
- **At the cathode.** Bubbles of a colourless gas which explode with a pop sound on exposure to a burning splint are evolved. This is because all the hydrogen ions from the sulphuric acid and water migrate to the cathode from where they are discharged by gain of electrons forming hydrogen gas which is the colourless gas. Therefore the acidity of the solution at the cathode decreases (alkalinity increases) since the hydrogen ions are discharged.



- **At the anode.** Bubbles of a colourless gas which relights a lighted splint are given out. This is because hydroxide ions are discharged in preference to the sulphate ions giving oxygen gas which is the colourless gas. This is so since it is lower than sulphate in the electrochemical series. Therefore alkalinity of the anode decreases (becomes acidic)

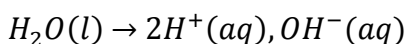
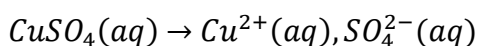


## Electrolysis of copper (II) sulphate solution using carbon electrodes

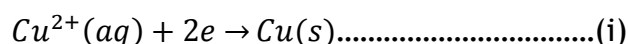


### Observations

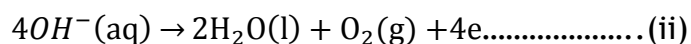
The ions present in solution are;



- **At the cathode.** Brown solid of copper deposited at the cathode and its weight increases. This is because both  $\text{Cu}^{2+}$  and  $\text{H}^+$  ions migrate to the cathode but  $\text{Cu}^{2+}$  is preferentially discharged since it is lower than hydrogen ion in the electrochemical series.

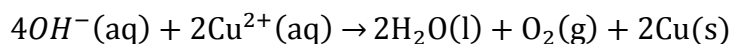


- **At the anode.** Bubbles of a colourless gas that relights a glowing splint are given out. This because hydroxide ion is discharged since it is lower than sulphate in the electrochemical series.



#### Overall equation

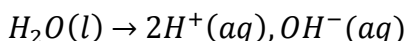
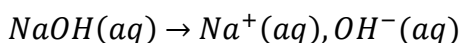
Multiplying equation (i) by 2 and adding it to equation (ii) gives;



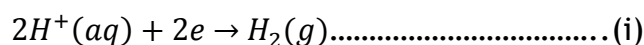
**NB//.** The blue colour of the solution fades to colourless due to the discharge of copper (II) ions responsible for the blue colour of the solution.

### Electrolysis of dilute sodium hydroxide solution using platinum electrodes

The ions present are;



- **At the cathode.** Both  $\text{Na}^+$  and  $\text{H}^+$  migrate to the cathode but  $\text{H}^+$  being lower in the electrochemical series is discharged in preference to  $\text{Na}^+$ .

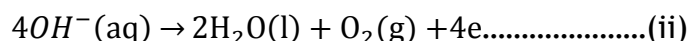


#### Observation

- ✓ Bubbles of a colourless that produce a pop sound when exposed to a burning splint are given out.



- **At the anode.** Hydroxide ions migrate to the anode and are discharged giving oxygen gas.

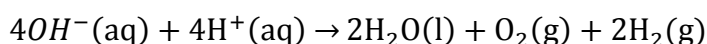


**Observation**

- ✓ A colourless which relights a glowing splint is evolved.

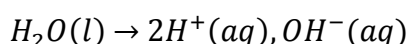
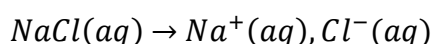
**Overall equation**

Multiplying equation (i) by 2 and adding it to equation (ii) gives;

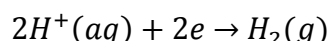


**Electrolysis of dilute sodium chloride solution using graphite as the anode and platinum as the cathode**

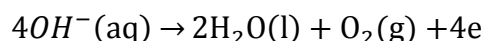
The ions present are;



- **At the cathode.** Both  $\text{Na}^+$  and  $\text{H}^+$  migrate there but its  $\text{H}^+$  selected for discharge since it is lower than  $\text{Na}^+$  in the electrochemical series.



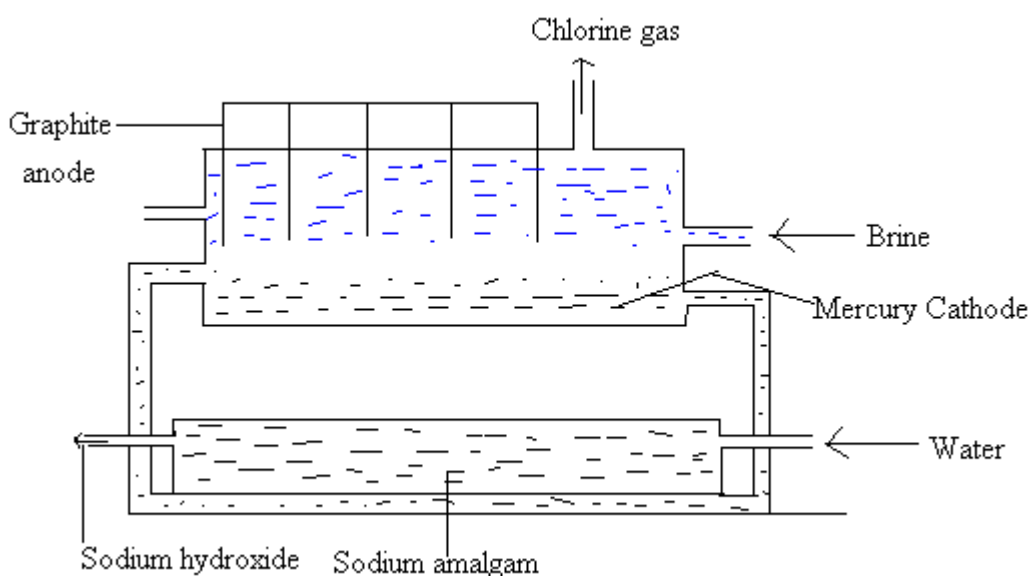
- **At the anode.** Both  $\text{OH}^-$  and  $\text{Cl}^-$  migrate there but the  $\text{OH}^-$  is discharged since it lower in electrochemical series.



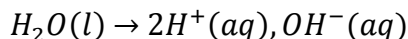
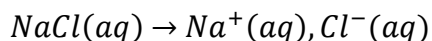
- 2) **Concentration of ions.** This applied for concentrated solutions and molten electrolyte. During electrolysis, the ions from the solute are discharged and not from the water. Do not use electrochemical series to select the ions for discharge. Consider the following examples;

**NB//.** A concentrated solution has a lot of solute with little amount of water.

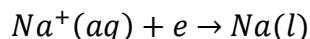
**Electrolysis of concentrated sodium chloride solution using carbon electrode as he anode and mercury as the cathode (manufacture of sodium hydroxide solution)**



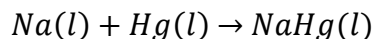
The ions present are;



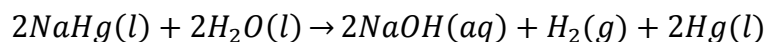
- **At the cathode.** Sodium ions are discharged at the cathode its concentration is higher than that of hydroxide ions.



The sodium formed reacts with mercury giving sodium amalgam.

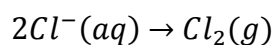


The amalgam is removed and treated with water to form sodium hydroxide solution.



The mercury liberated is returned to the cell.

- **At the anode.** The chloride ions are discharged at the giving chlorine gas (yellowish green gas that bleaches moist litmus paper)

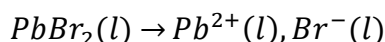


### Uses of sodium hydroxide

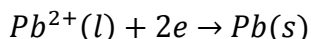
- To manufacture soap in the saponification process.
- Used as a laboratory agent.

### Electrolysis of molten lead (II) bromide using carbon electrodes

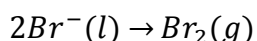
The present are;



- **At the cathode.** Lead (II) ions are discharged giving grey deposits of lead and weight of the cathode increases.



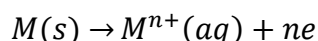
- **At the anode.** Reddish brown vapour is evolved because bromide ions are discharged giving bromine gas (reddish brown vapour)



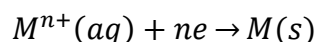
3) **Nature of the electrode.** This factor is applicable when dealing with electrodes of the same metal dipped into solution having the cation of the metal electrodes. One of such electrode is a thin sheet of pure metal that constitutes the cathode and the other is the impure metal which made as the anode. During electrolysis, the anode dissolves giving a solution having the ion of the metal that migrates to the cathode and discharged by gain of electrons. In the course, the size of the cathode increases in size while the anode decreases in size.

**NB//.** This reaction is used in the purification of metals

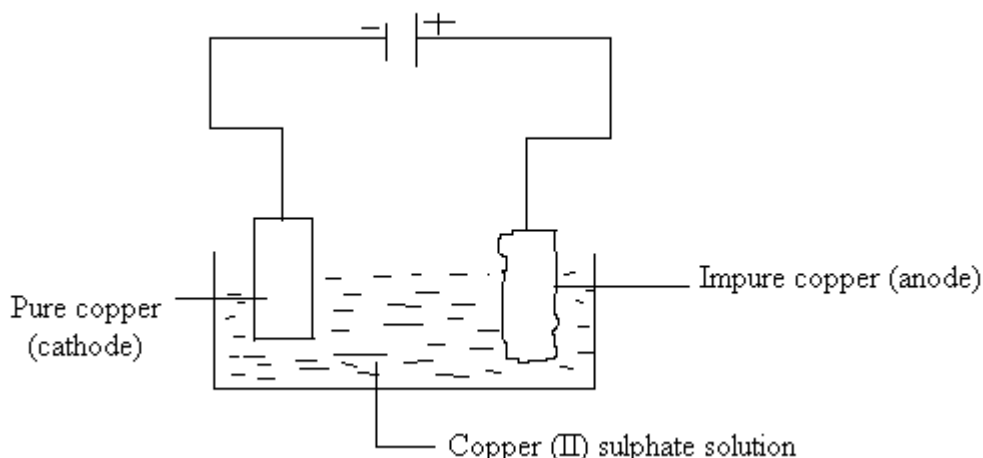
**At the anode;**



**At the cathode;**



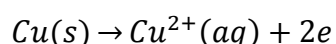
### Electrolysis of copper (II) sulphate solution using copper electrodes



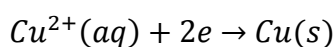
This reaction is used in the purification of copper. During the reaction, the impure copper dissolves to form a solution containing copper (II) ions. These ions migrate to the cathode where they are deposited as copper metal.

#### observations

- **At the anode.** The weight of the anode decreases because the copper atoms dissolve from it and enter the solution as copper (II) ion.



- **At the cathode.** The weight of the cathode increases because copper (II) ions are deposited as brown solid copper. The colour of the solution remains blue.



#### Laws of electrolysis

There are two laws of electrolysis called **Faraday's law** of electrolysis.

- (i) **Faraday's first law of electrolysis.** It states that, during electrolysis, the mass of substances liberated at the electrodes is directly proportional to the quantity of electricity passed through the electrolyte i.e.  $M \propto It$  where  $M$  is the mass liberated in grams (g)

$t$  is time in seconds (s)

$I$  is current in amperes (A)

#### Worked examples

- 1) A current of 2A was passed through copper (II) sulphate solution for 20 minutes. Find the quantity of electricity used.

#### Solution

$$Q = It = 2 \times 30 \times 60 = 3600C$$

- 2) A current of 4A was passed through lead (II) bromide solution for one and half hours. Find the quantity electricity passed through the solution.

## Solution

$$Q = It = 4 \times 60 \times 60 \times 1 \frac{1}{2} = 21600C$$

- (ii) **Faraday's second law of electrolysis.** It states that the quantity of electricity required to liberate one mole of any substance is directly proportional to the charge number on its ion i.e.  $Q \propto Z$  or  $It \propto Z$ . Where  $Z$  is the charge number on the ion. The charge number of an ion is the number of charge carried by one mole of an ion. E.g.

Ion	Charge number
✓ Na <sup>+</sup>	1
✓ Mg <sup>2+</sup>	2
✓ Al <sup>3+</sup>	3
✓ Cu <sup>2+</sup>	2
✓ OH <sup>-</sup>	1
✓ Br <sup>-</sup>	1

## Faraday's constant (F)

This is the quantity of electricity carried by one mole of electrons. It is numerically equal to 96500 coulombs (C). Sometimes this value is written as Faraday as summarized in the examples below.

Ion	Quantity of electricity (C) Required for discharge
✓ Na <sup>+</sup>	✓ 96500 (1 F)
✓ Mg <sup>2+</sup>	✓ 2 × 96500 (2F)
✓ Al <sup>3+</sup>	✓ 3 × 96500 (3F)
✓ Cu <sup>2+</sup>	✓ 2 × 96500 (2F)
✓ OH <sup>-</sup>	✓ 96500 (1F)
✓ Br <sup>-</sup>	✓ 96500 (1F)

In general, the liberation/discharge of one mole of ion,  $M^{n+}/X^{n-}$  may be written as;

$M^{n+}(aq) + ne \rightarrow M(s)$  Therefore to form 1 mole of metal M n Faraday's (nF) of electricity is required.

$X^{n-}(aq) \rightarrow X(g) + ne$  To form one mole of X(g), nF is needed.

Equations for discharge	Quantity of electricity required
$Pb^{2+}(l) + 2e \rightarrow Pb(s)$	✓ 2 × 96500 (2F)
$2Br^{-}(l) \rightarrow Br_2(g)$	✓ 2 × 96500 (2F)
$Na^{+}(aq) + e \rightarrow Na(l)$	✓ 96500 (1F)
$4OH^{-}(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e$	✓ 4 × 96500 (4F)

## Worked examples

- 1) A current of 2A was passed through a solution of copper (II) sulphate using carbon electrodes for 30 minutes. Calculate;
  - (i) the quantity of electricity used.
  - (ii) the mass of copper formed at the cathode.
  - (ii) the volume of oxygen gas liberated at s.t.p at the anode.
- 2) A current of 2A was passed through molten lead (II) bromide for one and half hours.
  - (a) Write equation for reaction at both the anode and cathode.
  - (b) Calculate the mass of substances formed at both the cathode and anode.
- 3) A current of 1.5 A was passed through dilute sulphuric acid for 4 minutes at s.t.p.
  - (a) Write equations for reaction that took place at the electrodes.
  - (b) Calculate the volume produced at the;
    - (i) cathode
    - (ii) anode.
- 4) Dilute sulphuric was electrolyzed between platinum electrodes and 112 cm<sup>3</sup> of oxygen gas and 224 cm<sup>3</sup> of hydrogen gas was given out at s.t.p.
  - (a) Explain why the volume of oxygen gas liberated is half that of hydrogen gas.
  - (b) Calculate the quantity of electricity required to liberate 112cm<sup>3</sup> of oxygen gas at s.t.p.
- 5) An aqueous solution of copper (II) sulphate and silver nitrate were connected in series. An inert electrodes were used and a current of 2 A passed through the solution for 4 minutes.
  - (a) Draw a diagram to show the circuit and the apparatus needed to carryout the electrolysis.
  - (b) A metal coating was deposited on one of the electrodes in solution. Write equation to show what happens to the metal as they are deposited.
  - (c) Calculate the moles of each metal deposited and hence mass.

### Applications of electrolysis

- ✓ Extraction of metals e.g. aluminum, sodium e.t.c.
- ✓ Purification of metals e.g. copper
- ✓ Manufacture of sodium hydroxide and chlorine
- ✓ Electroplating

### Electroplating

This is the coating of iron objects with a thin layer another metal so that it can be protected from corrosion and rusting. The process of electroplating also makes iron objects shining and attractive. The process requires the following;

- The material to be electroplated (iron).
- The metal to be used for electroplating.
- Electrolyte.

### Procedure

- The object to be electroplated is made the as the cathode.
- The metal for electroplating is made as the anode.
- The electrolyte is normally a soluble salt of a metal.

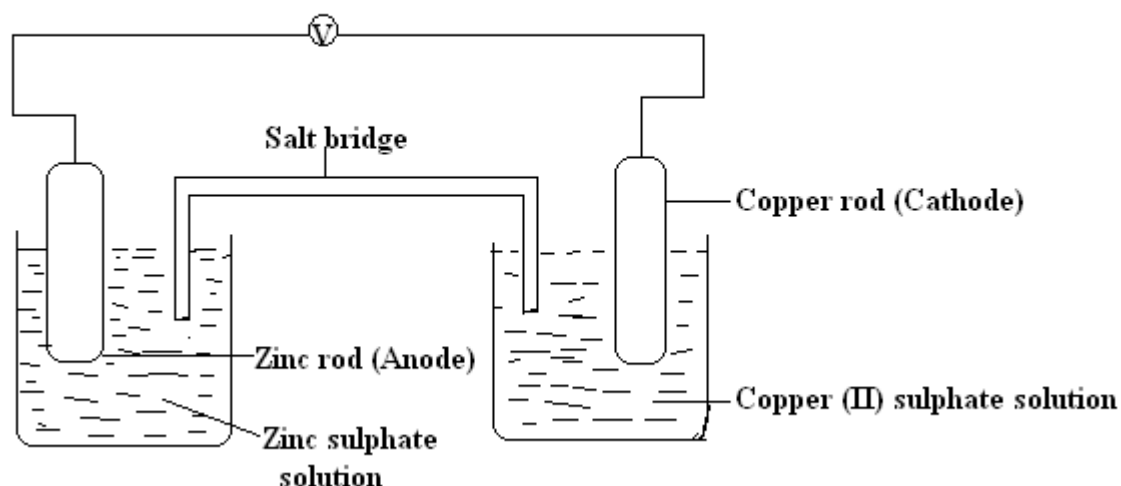
## Electrochemical cells

This is a cell which produces electro energy as a result of chemical reactions taking place. A common example is voltaic cell. A good example of a voltaic cell is the Daniel's cell. Daniel's cell is made of the following;

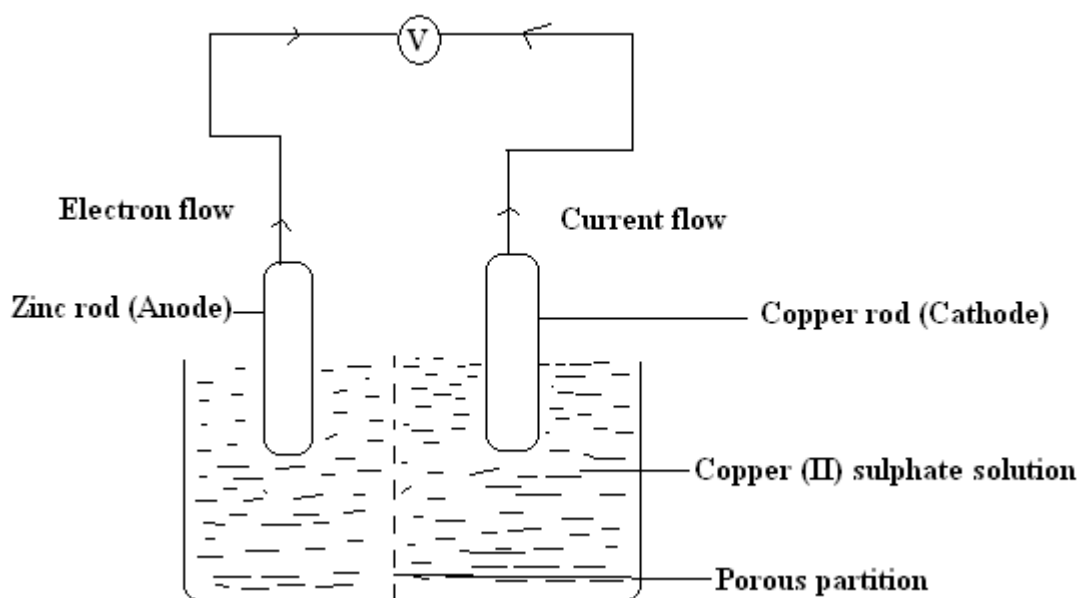
- Zinc rod which acts as the anode.
- Copper rod which acts as the cathode.
- Zinc sulphate solution which acts as the electrolyte at the anode.
- Copper (II) sulphate solution which acts as the electrolyte at the cathode.
- Porous partition which separates the electrolytes in the cathode and anode.

NB//. Sometimes instead of the porous partition, a salt bridge is used. This is a U-tube containing potassium nitrate/ammonium nitrate.

### Daniels's cell



OR



A salt bridge is an inverted U-tube made of glass and contains potassium chloride or potassium nitrate solution. It has the following functions;

- It keeps the ionic charges balanced from both sides of the cell.
- It provides electrical contact between the two electrodes by allowing ions migrate from one electrolyte to another.

The Daniell's cell works at a constant voltage measured by a voltmeter. However, the Daniell's cell cannot produce energy continuously because the chemicals get used up and hence the e.m.f of the cell becomes zero. The amount of energy produced by the Daniell's cell is calculated from;

Electrical energy,  $E = -nFE$

Where  $n$  = Number of electrons

$F$  = Faraday's constant

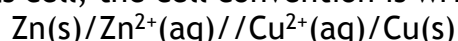
$E$  = Electrode potential

**Qn. Calculate the electrical energy produced by the Daniell's cell when the e.m.f of the cell is 1.5V.**

### Cell convention/notation

Given two metals, one higher in the electrochemical series becomes the anode (left hand electrode) while the one lower in the activity series becomes the cathode (right hand electrode).

For the Daniell cell, the cell convention is written as;



### Revision questions

## CARBON AND ITS COMPOUNDS

Electronic configuration: 2:4

Valency 4

Group 4

Period 2

Carbon is an element which belongs to group 4 and period 2 in the periodic table

It's a non metal and mainly forms compounds which are covalent such as carbondioxide, carbonmonoxide, hydrogencarbonates etc

### Natural occurrence

In nature, carbon occurs in compounds such as carbonates and hydrogen carbonate s and in plants and animal tissues or cells. It also occurs in gases like carbondioxide and carbon monoxide.

### Forms of carbon

Carbon occurs in two major forms and these two major forms are referred to as allotropes of carbon which include diamond and graphite.

Allotrope refers to the different forms of an element in the same physical state. Other elements which exhibit this characteristic include:

ELEMENT	ALLOTROPES
sulphur	-Rhombic sulphur -monoclinic sulphur
phosphorus	-red phosphorus -white phosphorus

### Allotropy

This is the existence of an element in more than one form without the change of physical state.

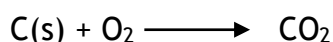
Another form of carbon is sometimes called amorphous carbon e.g. charcoal.

NB. Diamond and graphite are sometimes referred to as crystalline allotropes of carbon.

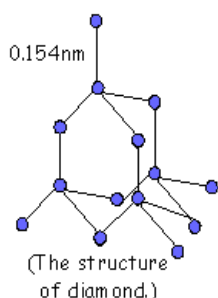
### Proof to show that diamond and graphite are allotropes of carbon.

Equal amounts of diamond and graphite are separately burnt in oxygen under the same conditions.

The same volume of carbondioxide is produced by burning both substances. This therefore shows that diamond and graphite are allotropes of carbon.



### Structure of diamond



Diamond has a structure called giant atomic structure which is made up of carbon atoms covalently bonded together. These bonds are very strong and are many. This gives diamond a three dimensional lattice structure where each carbon atom uses all the four valency electrons to form four covalent bonds with four other carbon atoms to give carbon atom shape. Therefore the coordination number of each carbon atom is four meaning that each carbon atom is surrounded by four other carbon atoms.

### Properties of diamond.

- Its a very hard substance.
- It has a very high meltingpoint.
- It is resistant to chemical reactons.



**NB.** All these properties are due to very strong covalent covalent bonds in the structure of diamond.

- It has a very high density due to close packing of the carbon atoms in the structure.
- It is a non conductor of electricity due the absence of mobile electrons (delocalize electrons) in the structure.

### Uses of diamond

- It is used in the manufacture of glass cutters.
- It is use for the manufacture of drilling machines.

**NB.** The above uses are because diamond is the hardest known substance.

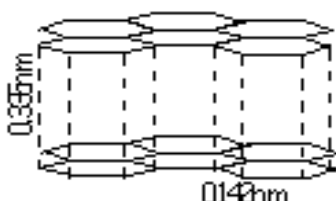
- It is use as valuables such as jewellery, ornaments because it is transparent and when polished it can glitters, shine, and above all, it has a very high refractive index.
- It is use as an insulator because it is non-conductor of electricity.

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### Graphite

It is an allotrope of carbon and exists as black soft slippery hexagonal crystals. Naturally, it is found under the ground as an impure form called plumbago.

### Structure of graphite



The structure of graphite consists of parallel layers and each layer consists of hexagonal rings with each carbon covalently bonded to three other carbon atoms.

In graphite every carbon uses only three electrons for bonding. This leaves every carbon atom with one free electron not used for bonding therefore this unused electron moves freely within the parallel layers. This electron is called delocalized electrons, for this reason graphite is responsible for conduction of electricity.

### Properties of graphite.

- It is fairly a short substance with low melting point. This is due to the weak vanderwaals forces holding the parallel layers.

- It has a lower density compared to diamond because the parallel layers do not pack closely together.
- It is a good conductor of electricity and heat due to the many free and delocalized electrons that are able to move along the layers.
- The parallel layers can slide over one another due to the weak forces holding them.

### Uses of graphite

- It is used in the manufacture of pencil lids. This is made out of graphite and clay because it is black and opaque.
- It is used as a lubricant for small bearings like those dynamos and vacuum cleaner motors which requires little but regular lubrication. This is because graphite is slippery and soft.
- It is used in the manufacture of electrode because graphite is a good conductor of electricity due to the presence of delocalized electrons.
- It is used as a protective coating of iron to prevent it from rusting.

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### Amorphous carbon.

It is non crystalline allotrope of carbon and it exists in many forms like:

- **Wood charcoal:** it is made by burning wood out of contact with air. This is called destructive distillation of wood.  
-it is used as fuel for domestic use  
for making of gas masks absorbs poisonous gas

- **Animal charcoal;** it is made by heating bones out of air. It is also called destructive distillation of bones.

It is used in the purification of sugar especially it removes the brown colour of crude sugar therefore making it white.

- **Lamp black/soot:** it is made when petroleum burns in limited supply of air. It is used to make black shoe polish, carbon papers.
- **Coke:** it is formed by heating coal out of contact with air. it is used as fuel in boilers,ovens,etc

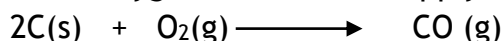
### Chemical reactions of carbon

#### 1) Combustion:

When burnt in plentiful supply of oxygen, carbon dioxide is formed. The reaction produces a lot of heat. For this reason carbon is used as a source of heat.



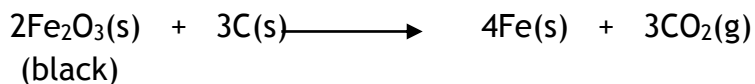
When oxygen is in limited supply carbon monoxide is formed.



This gas is very poisonous.

#### 2) Reducing property:

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### 3) Reaction with concentrated sulphuric acid.



## 1. CARBONDIOXIDE.

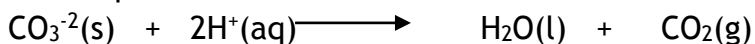
## Laboratory preparation of carbon dioxide

In the laboratory carbondioxide is prepared by reacting a carbonate with dilute acid. Usually calcium carbonate (marble chips) and dilute hydrochloric acid are used to prepare carbondioxide.

### Equation



ionic equation.

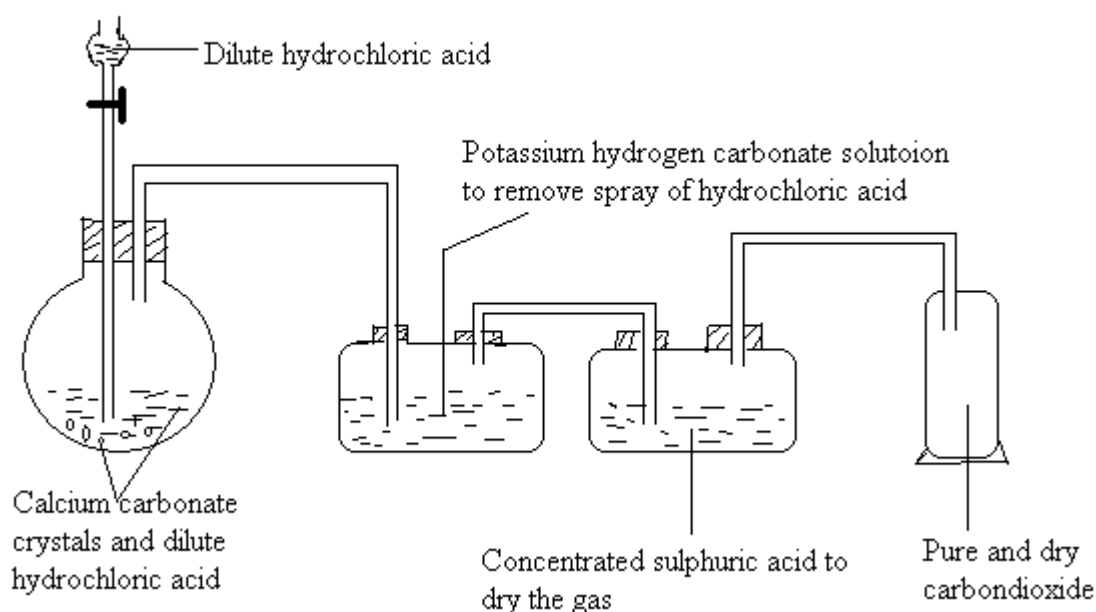


## Procedure

- Calcium carbonate is put in to a round bottomed flask.
- Then dilute hydrochloric acid is added to the Calcium carbonate. immediately carbondioxide gas is evolved which is seen as colourless bubbles.
- The gas is passed through potassium hydrogen carbonate solution put in a wash bottle. The function of this solution is to remove sprays of hydrochloric acid from carbondioxide given.

- The gas is again passed through another wash bottle containing concentrated sulphuric acid to dry it. Or it can also be dried by passing the gas through a U-tube containing anhydrous calcium chloride.
- The pure and dry carbondioxide can then be collected by downward delivery or upward displacement of air since it is denser than air.

### Diagram.

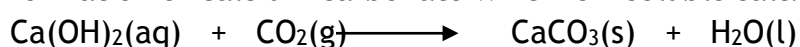


### Chemical test for carbondioxide

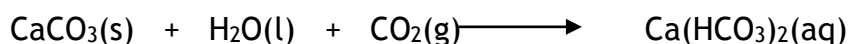
Reagent:lime water (calcium hydroxide solution)

#### Procedure

When carbondioxide is passed through lime water, the limewater turns milky. This reaction confirms the presence of carbondioxide. The milky solution is due to the formation of calcium carbonate which is insoluble salt.



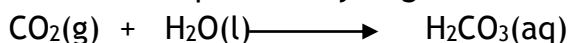
NB// when excess carbondioxide is bubbled, the milky solution dissolves to form a colourless solution due to the formation of soluble calcium hydrogen carbonate.



### Physical properties of carbondioxide

- It is a colourless gas.
- It is odourless.
- It is denser than air.

- It is slightly an acidic gas because it turns damp blue litmus paper to pink or red. This is because it dissolves in water to form a weak carbonic acid which partially ionizes to produce hydrogen ions which make the solution acidic.

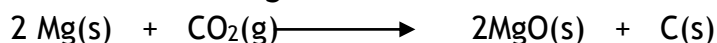


- It does not support burning otherwise it extinguishes burning splint.

### Chemical properties of carbondioxide

#### (a) Burning magnesium;

When burning magnesium is lowered into a gas jar of carbondioxide it continues to burn less vigorously to form white solid called magnesium oxide and also black solid called carbon is formed. This reaction shows that carbondioxide is made out of carbon and oxygen. The reaction is possible because the heat supplied by burning magnesium first decomposes carbon dioxide to carbon and oxygen which then reacts with magnesium to form magnesium oxide.



#### (b) Sodium hydroxide.

Carbon dioxide reacts with sodium hydroxide because it is an acidic gas e.g. when bubbled through concentrated solution, sodium hydroxide reacts with it to form sodium carbonate and water which is seen as colourless solution. When more carbon dioxide is bubbled through this solution sodium hydrogen carbonate is formed which is seen as white precipitate.



**NB.1** In nature carbondioxide can be produced by the following processes.

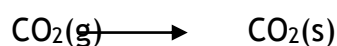
- Combustion of burning organic matter, petroleum products.
- Respiration. This takes place in animal and plant tissues.

2. Process that consumes carbondioxide is photosynthesis.

3. In nature there must be a balance between nature processes that produce carbondioxide and nature processes which reduce the amount of carbondioxide. an imbalance leads to increased production of carbondioxide in the atmosphere. This leads to green house effect which causes global warming.

### Uses of carbondioxide

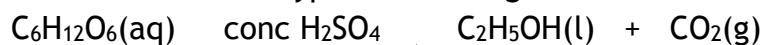
- It is used in fire extinguisher because it does not support burning.
- It is used in the manufacture of aerated drinks such as lemonado, beer, soda because it gives these drinks their flavour and tastes.
- It is used in refrigerators as dry ice ( solid carbondioxide).



- It can be used to make sodium hydrogen carbonate and sodium carbonate by solvay process.
- It can be used in making rain especially in cloud setting. In this process dry ice is scattered in the atmosphere and it facilitates rapid cooling of the cloud.

### Industrial preparation of carbondioxide

- Burning of carbon (coke).
- Liquefaction of air.
- It is obtained as a byproduct during the manufacture of alcohol.



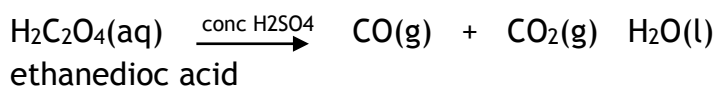
- Decomposition of limestone (calcium carbonate).



## 2. CARBONMONOXIDE GAS.

### Laboratory preparation.

(a) Dehydration of oxalic acid using concentrated sulphuric acid.

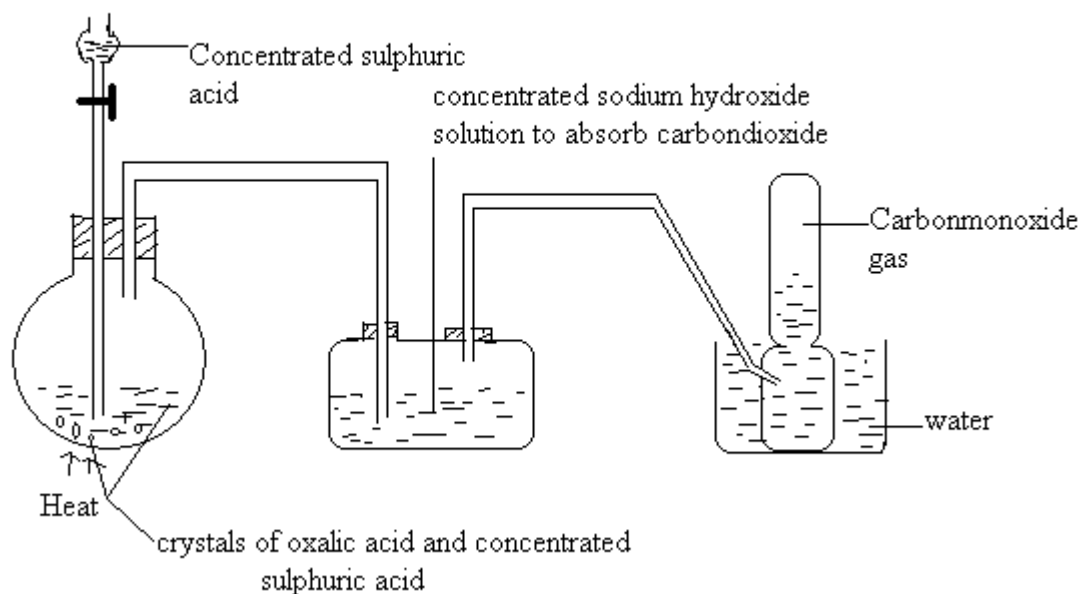


### Procedure

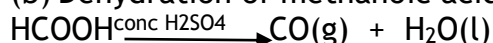
- Crystals of oxalic acid are put in a flask. Concentrated sulphuric acid is then added to the crystals through a thistle funnel and the mixture is heated.
- Both carbonmonoxide and carbondioxide are given out or evolved. The gas mixture is then passed through a concentrated solution of potassium hydroxide or sodium hydroxide which absorbs carbondioxide from the mixture leaving carbonmonoxide which can be dried by passing it through anhydrous calcium chloride.
- The gas can be collected over water by upward delivery if wanted dry.

### Explanation

the reaction between concentrated sulphuric acid and oxalic acid is called dehydration reaction because elements of water is removed from the oxalic acid by the concentrated sulphuric acid therefore sulphuric acid is referred to as a drying agent.

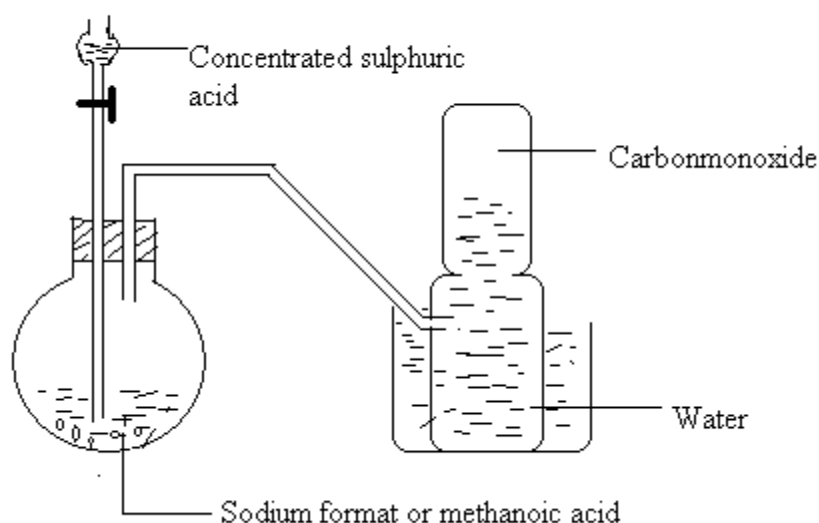


(b) Dehydration of methanoic acid by concentrated sulphuric acid.



This reaction does not take place in presence of heat. it is called dehydration reaction because water has been removed from methanoic acid therefore sulphuric acid is said to be a dehydrating agent.

**Diagram**



## CHEMICAL TEST FOR CARBONMONOXIDE

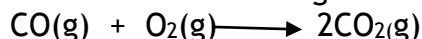
This gas is neutral to litmus paper and hence it has no effect on it. However it has the following physical properties which can be used for its identification.

- It is colourless
- it is slightly soluble in water.

- it is slightly less dense than air.]
- it is odourless (tasteless)

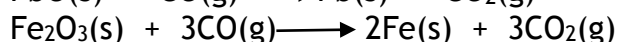
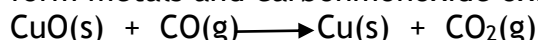
## CHEMICAL PROPERTIES

**a) Combustion.** it burns in air with a blue flame giving carbondioxide. For this reason it is not used in fire extinguisher because it supports burning.



This reaction can also be use to identify carbonmonoxide.

**b) Reduction.** Carbonmonoxide is a very powerful reducing agent because it removes oxygen from compounds like metal oxides. In this reaction, the oxides are reduced to form metals and carbonmonoxide oxidized to carbondioxide.



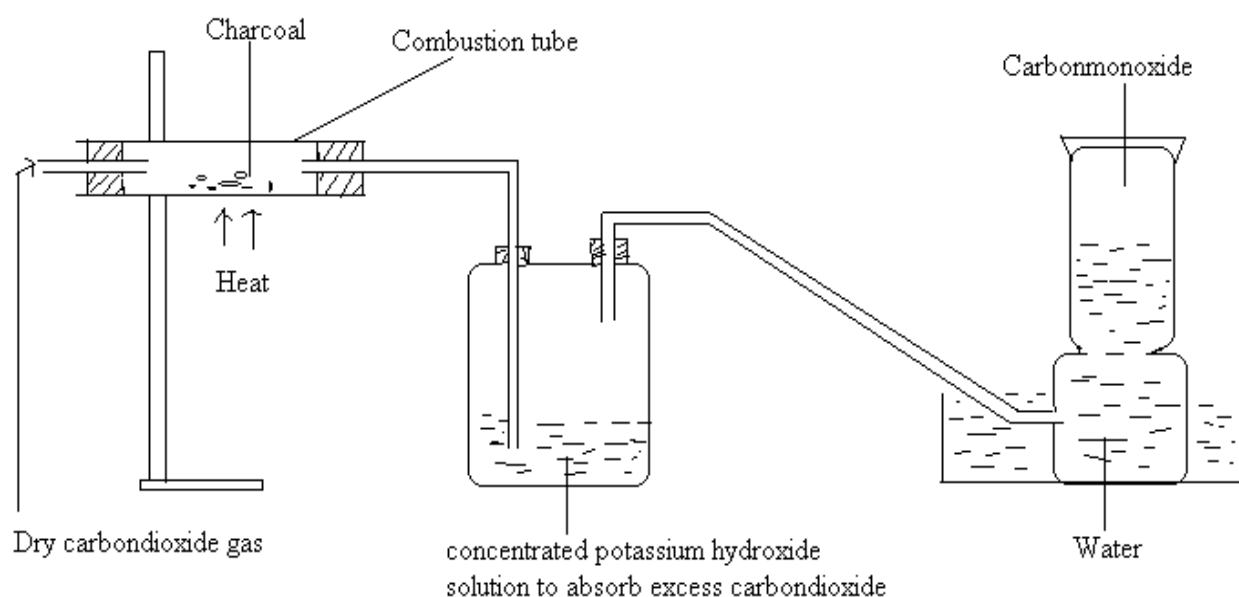
This reaction is use to extract metals with their oxides.

Carbonmonoxide is a poisonous gas. For this reason it is prepared in a fume cupboard.

**(c) Preparation of carbonmonoxide from carbondioxide.**

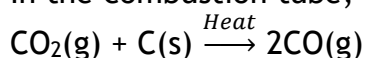
### Procedure

- Dry carbpndioxide gas is passed over heated charcoal in a combustion tube.
- The excess carbondioxide gas is passed through a wash bottle containing concentrated potassium hydroxide which absorbs it.

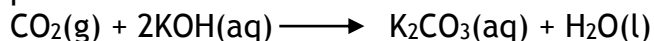


### Reaction equations;

In the combustion tube, carbondioxide gas is reduced to carbonmonoxide.



In the wash bottle, carbondioxide is absorbed by the potassium hydroxide solution giving potassium carbonate and water.





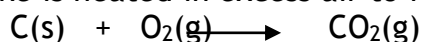
## Differences between carbonmonoxide and carbondioxide

Carbonmonoxide	Carbondioxide
✓ Insoluble in water and alkalis	✓ Soluble in water and alkalis
✓ Has no action on litmus paper	✓ Turns blue litmus paper pink or red
✓ Burns with a blue flame	✓ Does not burn
✓ It is reducing agent	✓ It is non reducing agent
✓ Les dense than air	✓ It is denser than air
✓ Does not react heated charcoal	✓ Reacts with heated charcoal

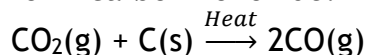
## Producer gas

This is fuel gas consisting of carbonmonoxide gas and nitrogen gas in the ratio 1 : 2. This gas is made into two stages;

- Coke is heated in excess air to form carbondioxide gas.



- Carbondioxide produced is reacted with more soot at elevated temperature to form carbonmonoxide.



- The carbonmonoxide and unreacted nitrogen obtained by fractional distillation of air forms a producer.

## Uses of producer gas

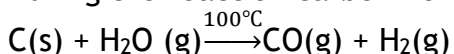
- It is used as a source of fuel when burnt.
- It acts as a source of nitrogen during the manufacture of ammonia in the Haber process.

## Water gas

It is fuel gas consisting of equal volume of carbonmonoxide and hydrogen.

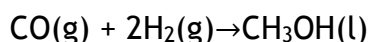
The gas is prepared by passing steam over coke heated to a very high temperature.

During the reaction carbonmonoxide and hydrogen gases are formed.



## Uses of water gas

- It is used as a source of fuel gas for heating.
- It acts as an industrial source for hydrogen gas.
- It is used for making methanol b y passing more hydrogen gas through the mixture.



**NB//.** Water gas is a better source of fuel than producer gas because it produces a lot of heat far more than that of producer gas. The hydrogen gas and carbonmonoxide gas of water both burn to produce heat but for producer gas, it is only the carbonmonoxide gas that burns to produce heat.

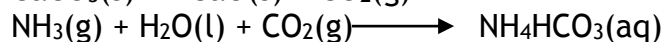
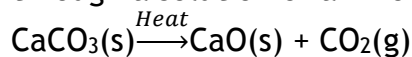
## Sodium carbonate (soda ash)

This is obtained on a large scale by the **Solvay process**. The raw materials used in the Solvay process are;

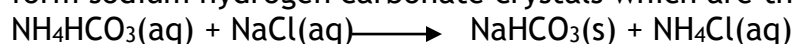
- ✓ Concentrated sodium chloride solution (brine)
- ✓ Ammonia solution
- ✓ Calcium carbonate (marble chips)

The process can be summarized as follows;

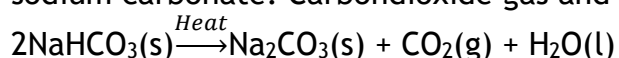
- Carbondioxide gas obtained by the strong heating of calcium carbonate is passed through a solution of ammonia in tower forming ammonium carbonate solution.



- The ammonium carbonate solution trickles down the tower and combines with concentrated sodium chloride solution placed in the lower part of the tower to form sodium hydrogen carbonate crystals which are then filtered off.



- The sodium hydrogen carbonate crystals are then heated strongly giving anhydrous sodium carbonate. Carbondioxide gas and water are also formed.



**NB//.** Hydrated sodium carbonate ( $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ ) is not soda ash

#### Uses of soda ash (anhydrous sodium carbonate)

- For making of glasses.
- For softening hard water.

#### Uses of Washing soda (sodium hydrogen carbonate)

- For making baking powder
- Used for making healthy salts to neutralize excess hydrochloric acid in the stomach.
- Used in fire extinguishers.

#### Carbonates and hydrogen carbonates

These compounds of carbon derived from carbonic acid. Naturally Carbonates occur as;

- ✓ Chalk
  - ✓ Limestone
  - ✓ Marble
  - ✓ Coral reef
  - ✓ Egg shells
- } Carbonates

#### Chemistry of carbonates

##### Chemical properties of carbonates

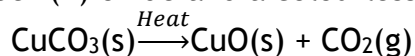
- 1) **Solubility.** All carbonates are insoluble in water except the carbonates of potassium, sodium and ammonium.
- 2) **Action of heat.** All carbonates decompose completely when heated to metal oxide and carbondioxide except the carbonates of potassium and sodium which do not decompose on heating. And ammonium carbonate which decomposes to ammonia gas, carbondioxide and water.



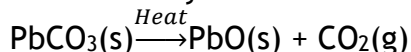
Where M is a metal

Consider the action heat on the following carbonates.

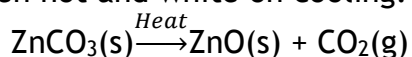
(i) **Copper (II) carbonate.** It is green solid which decomposes to give a black solid of copper (II) oxide and a colourless which turns lime water milky.



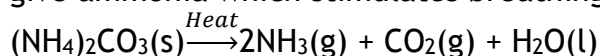
(ii) **Lead (II) carbonate.** It is a white solid and decomposes to a reddish brown solid when and on cooling turns yellow and a colourless which turns lime water milky evolved.



(iii) **Zinc (II) carbonate.** It is white solid that decomposes to zinc oxide which is yellow when hot and white on cooling.



(iv) **Ammonium carbonate.** It is used for smelling of salts because it decomposes to give ammonia which stimulates breathing.

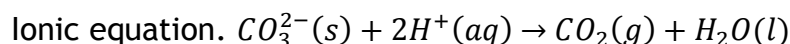
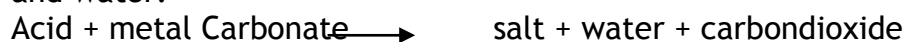


**Observations** can be stated as follow;

- ✓ Bubbles of colourless gas that turns moist red litmus paper blue evolved.
- ✓ Bubbles of colourless gas that turns lime water milky given off.
- ✓ Colourless liquid that turns anhydrous copper (II) sulphate blue or anhydrous cobalt (II) chloride paper pink.

**NB//. Aluminum carbonate does not exist.**

3) **Reaction with acids.** All carbonates react with acids to form salt, carbondioxide gas and water.



**NB//.** Dilute sulphuric acid reacts very slowly with calcium carbonate and lead (II) carbonate and the reaction normally stop before completion.

**Reason**

This is because of the formation of insoluble salt (calcium sulphate and lead (II) sulphate) that covers the remaining calcium carbonate or lead (II) carbonate which prevents the acid from reacting further. Hence the reaction produces very little volume of carbondioxide and it cannot be used to produce carbondioxide.



Lead (II) carbonate also reacts with dilute hydrochloric acid forming an insoluble lead (II) chloride which stops further from taking place.

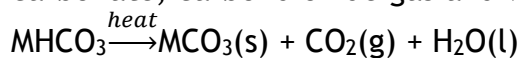


**NB//.** Nitric acid reacts with any carbonate without the formation insoluble salt because all nitrates are soluble.

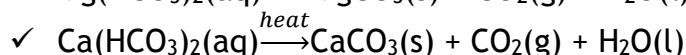
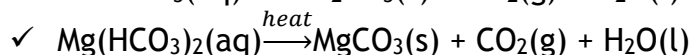
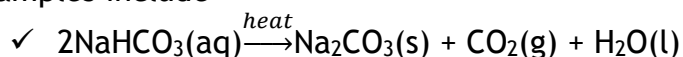
## Chemistry of hydrogen carbonates

1) **Solubility.** All hydrogen carbonates are soluble in water. The hydrogen carbonates which do not exist in the solid state are magnesium hydrogen carbonate and calcium hydrogen carbonate. The rest of the hydrogen carbonates exist in the liquid state.

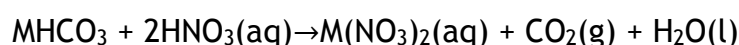
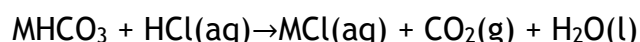
2) **Action of heat.** When a hydrogen carbonate is heated either in solution or solid state, carbonate, carbondioxide gas and water are formed.



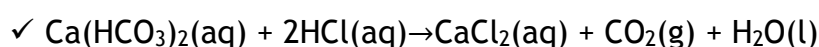
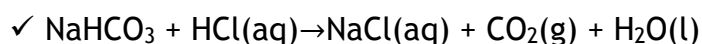
Examples include



3) **Reaction with acids.** All hydrogen carbonates liberate carbondioxide gas, water and salt when reacted with any acid.



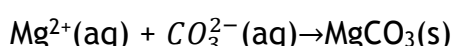
Examples include



### Difference between carbonates and hydrogen carbonates

Magnesium (II) sulphate solution is used to distinguish between a carbonate and a hydrogen carbonate.

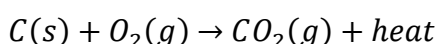
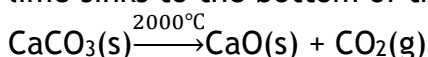
**With carbonate,** a white precipitate of magnesium carbonate is formed.



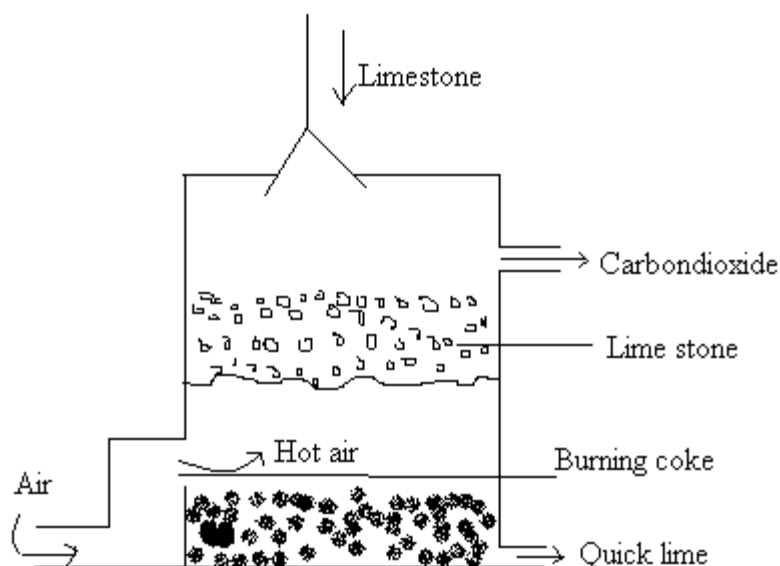
**With hydrogen carbonate,** the solution remains colourless (no observable change)

### Quick lime (calcium oxide)

It is obtained on a large scale by strongly heating limestone (calcium carbonate) in a lime kiln. During the reaction, limestone decomposes completely to quick lime and carbondioxide. The heat required for the reaction is supplied by burning coke. The quick lime sinks to the bottom of the lime kiln where it is removed.



### Diagram of lime kiln

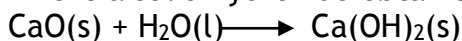


### Uses of quick lime

- Used for drying of ammonia gas since it is hygroscopic.
- Used in the manufacture of cement
- Used as a fertilizer since it contains calcium ions as a plant nutrient.
- Used in the manufacture of slaked lime (calcium hydroxide).

### Slaked lime (calcium hydroxide)

This is a solid hydroxide obtained by adding cold water to quick lime.



When water is added to quick lime, the following observations can be made;

- It swells with a hissing sound liberating a lot of heat.
- A white solid is formed which cracks and crumbles to form a white powder called slaked lime.

### Uses of slaked lime

- Used to make white wash (quick suspension of slaked lime in water).
- Used for smoothening walls of buildings.
- Used for softening hard water
- Used to make mortar (mixture of slaked lime, sand, and water) and sticking bricks as well as plastering walls.
- Used as fertilizer.

### Concrete

This is a mixture of cement, sand, gravel and water.

### Cement

It is mixture of quick lime, clay and sand.

Stop for now. Carbon chemistry ends here.

### Revision questions

#### SULPHUR AND ITS COMPOUNDS

Sulphur is a non metal with atomic number of 16 and mass number of 32. It belongs to group 6 and period 3 in the periodic table.

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#### Natural occurrence of sulphur

There are many sources of sulphur but the most important ones are;

##### 1) Sulphides

In sulphides sulphur is found combined with metals like iron, zinc, lead and copper. Sulphides are generally black solids found under ground. When extracted from the ground, sulphur can be obtained from them.

##### 2) Sulphur in free state

This is the major source of sulphur. The free sulphur is found underground normally extracted by a process called the Frasch's process. This process was discovered by Frasch in USA and in order to sulphur using this method,

- (i) three concentric pipes are lowered in sulphur deposits in the ground.
- (ii) Super heated water is forced through the outer pipe to melt the sulphur since the melting point of sulphur is higher than that of water.
- (iii) Hot compressed air is then blown through the inner most pipe forcing sulphur to rise from the ground through the central pipe.
- (iv) Hot air also keeps the sulphur in a hot state as it rises to the ground surface.
- (v) Molten sulphur is then allowed to solidify as it reaches the ground surface.

#### Physical properties of sulphur

- It is yellow solid at room temperature and each molecule of sulphur consists of eight sulphur atoms joined by covalent bonds. The sulphur molecules are joined by a weak intermolecular force called vanderwaals forces.
- Sulphur is insoluble in water because it is made up of covalent bonds. It is soluble in non polar solvents like benzene.
- The melting point of sulphur is  $119^{\circ}\text{C}$  and boiling point is  $444^{\circ}\text{C}$  with a density of  $2.08\text{ gm}^{-3}$ .

#### Allotropes of sulphur

Allotropes are different forms of the same element in the same physical state. Sulphur has two forms ie crystalline allotropes and amorphous allotropes.

Crystalline allotropes. There are two crystalline allotropes of sulphur.

(a) Rhombic sulphur (alpha)

(b) Monoclinic sulphur (beta)

They are crystalline because their crystals have regular shapes while amorphous sulphur has no regular shape.

#### Differences between rhombic and monoclinic sulphur

Rhombic sulphur	Monoclinic sulphur
<ul style="list-style-type: none"><li>• Stable below 96.5°C</li><li>• Bright yellow</li><li>• Density is 2.08 g/cm<sup>3</sup></li><li>• Crystals are octahedral in shape</li></ul>	<ul style="list-style-type: none"><li>• Stable above 96.5°C</li><li>• Pale yellow</li><li>• Density is 1.98 g/cm<sup>3</sup></li><li>• Crystals are needle shaped</li></ul>

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#### Action of heat on sulphur

In the absence of air or oxygen, when sulphur is heated;

- It melts at 115°C to give a dark brown mobile liquid which easily flows. This is because the forces of attraction between sulphur molecules are broken and the ring of sulphur atoms slide over one another allowing easy flow of the liquid.
- On further heating, it becomes darker and very viscous i.e. does not easily flow. This is because the sulphur rings are converted into long chains of sulphur atoms that twist with one another thereby preventing easy flow.
- Gradually on further heating, the viscous liquid becomes mobile again i.e. easily flows. This is because the long sulphur chains are broken into shorter chains which do not twist with one another and becomes reddish brown in colour.
- Finally at 444.4°C, the sulphur boils to give brown sulphur vapour which condenses on colder surface to form a yellow sublimate called sulphur flowers.

#### Experiment to show that monoclinic and rhombic sulphur are allotropes of sulphur

When equal amounts of monoclinic and rhombic sulphur are separately burnt in air they produce the same volume of sulphur dioxide.

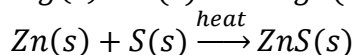
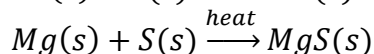
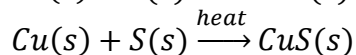
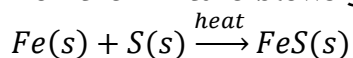
#### Uses of sulphur

- Used in the vulcanization of rubber. Here raw rubber is mixed with sulphur and heated. This makes rubber hard, strong and elastic. The rubber can be used for making of car tyres and shoe soles.
- Used in the manufacture of drugs, ointments and fungicides.
- Used in the manufacture of sulphuric acid in the contact process.
- Used in the manufacture of gun powder, matches, and fire works.
- Used in the manufacture of calcium hydrogen sulphide, a bleaching agent for wood pulp.
- Used for making paper.

## Chemical properties of sulphur

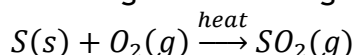
### 1) Reaction with metals

Sulphur combines directly with heated metals (iron, copper, magnesium and zinc) to form metal sulphide. All the metal sulphides formed are black in colour. With iron the mixture blows giving a red hot light with a black solid of iron (II) sulphide.



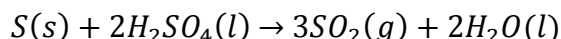
### 2) Reaction with oxygen/air

Sulphur burns with a blue flame forming sulphurdioxide gas. This gas has a choking and irritating smell.

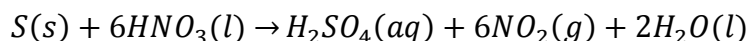


### 3) Reaction with acids

(a) with **concentrated sulphuric acid**, it oxidizes sulphur to sulphurdioxide and acid is reduced to water.

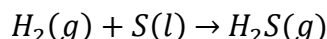


(b) **Concentrated nitric acid**. When heated with sulphur, reddish brown fumes of nitrogendioxide gas is formed. In this reaction sulphur is oxidized to sulphuric acid.



### 4) Reaction with hydrogen

When hydrogen gas is bubbled thorough molten sulphur in a boiling tube. Hydrogen sulphide gas is given off which is confirmed by holding filter paper soaked in lead acetate of lead (II) nitrate near the mouth of the tube. This forms black solid of lead (II) sulphide.

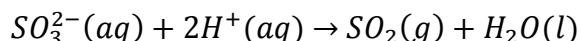
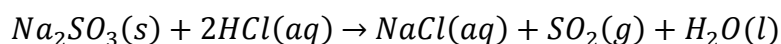


## Compounds of sulphur

### Sulphurdioxide

#### Laboratory preparation

(a) By reacting sodium sulphite with concentrated hydrochloric acid

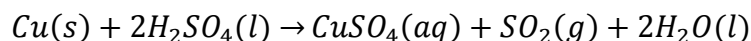


#### Procedure

- Concentrated hydrochloric acid is added through a thistle funnel to crystals of sodium sulphite.
- Immediately sulphurdioxide is evolved and is driven out of the flask by the presence of heat through a delivery tube to a wash bottle containing sulphuric acid which dries it.
- The dry gas is collected by downward delivery since it is denser than air.

(b) By reacting concentrated sulphuric acid with copper turnings

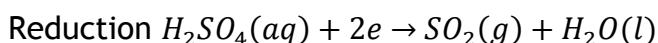
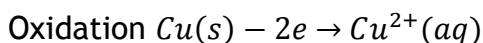




#### Procedure

- Concentrated sulphuric is added to copper turnings in a flask and the mixture is heated.
- Sulphurdioxide gas is produced and passed through wash bottle containing sulphuric acid to dry it.
- The gas is finally collected by downward delivery since it is denser than air.

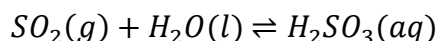
**NB//.** In the reaction concentrated sulphuric acid is an oxidizing agent because it takes electrons from copper hence reduced to sulphurdioxide and water while copper behaves as a reducing agent by giving out electrons thus oxidized to copper (II) sulphate.



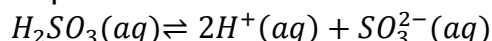
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#### Physical properties of sulphurdioxide gas

- It is a colourless gas.
- It has a choking irritating smell.
- It is denser than air.
- It is soluble in water forming sulphurous acid.



This acid is unstable therefore ionizes immediately to form hydrogen ions and sulphite ions.



For this reason, sulphurdioxide is said to be an acid anhydride (an oxide of a non-metal that reacts with water to form an acid). Other acid anhydrides include nitrogendioxide, carbondioxide, sulphurtrioxide and phosphorous pentaoxide.

#### Test for sulphurdioxide in the laboratory

Reagents	Procedure	Observation
Acidified potassium permanganate solution	• Bubble sulphurdioxide gas through a solution of acidified potassium permanganate.	• The purple colour of the solution changes to colourless.
Acidified potassium dichromate solution	• Bubble sulphurdioxide gas through a solution of acidified potassium dichromate.	• The orange colour of the solution changes to green.

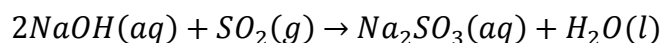
#### Chemical properties of sulphurdioxide gas

##### 1) Acidic property.

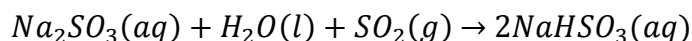
Sulphurdioxide is acidic because it turns damp blue litmus paper red and forms an acid when dissolved in water.

Due to its acidic property, it reacts with sodium hydroxide to form salts.

With limited sodium hydroxide solution, sodium sulphite (normal salt) is formed.

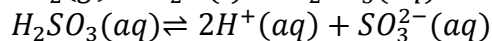
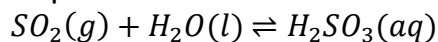


With excess of the gas, sodium hydrogen sulphite (acidic salt)

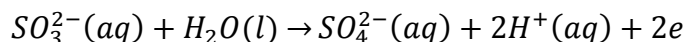


## 2) Reducing property

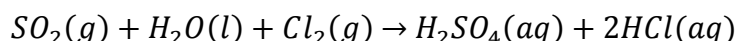
Dry sulphurdioxide does not behave as a reducing agent but in the presence of water it behaves as a powerful reducing agent. This is because of the formation of sulphurous acid which ionizes to sulphite ions believed to be a reducing agent.



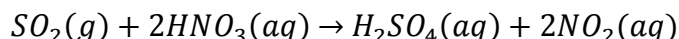
On oxidation, it forms a sulphate



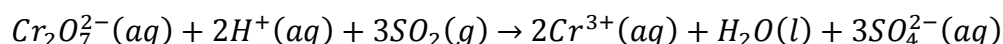
- (a) **With chlorine.** When bubble through the solution, the greenish colour of chlorine turns colourless. In the reaction, chlorine is reduced to hydrochloric acid and sulphurdioxide oxidized to sulphuric acid.



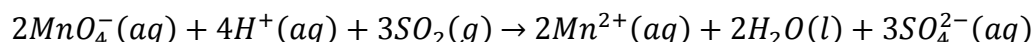
- (b) **With concentrated nitric acid.** Sulphurdioxide reduces concentrated nitric acid to brown fumes of nitrogendioxide and itself oxidized to sulphuric acid.



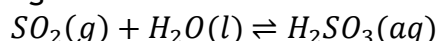
- (c) **With acidified potassium dichromate solution.** It reduces acidified potassium dichromate to chromium (III) ions which is green colour.



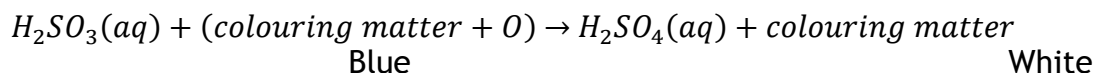
- (d) **With acidified potassium permanganate solution.** It reduces acidified potassium permanganate to manganese (II) ion which is colourless and itself oxidized to sulphate ions.



- 3) **Bleaching action.** Sulphurdioxide gas acts as a bleaching agent when dissolved in water. This is shown by dropping few moist blue flowers into a gas jar of sulphurdioxide. The flowers would loose their colour within a few minutes. Sulphurdioxide reacts with water to form sulphurous acid which is the bleaching agent.



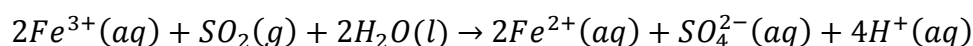
This acid takes up oxygen atom from the colouring matter and in the process, the blue flowers become white (bleached).



If the bleached flowers are exposed to air, they regain their colour by absorbing oxygen from air.

**NB//.** Sulphurdioxide bleaches by reduction (takes oxygen away from dyes) while chlorine bleaches by oxidation (gives oxygen to dyes)

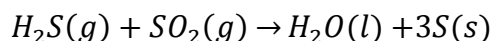
- 4) **Reaction with iron (III) sulphite.** When sulphurdioxide gas is bubbled through iron (III) sulphate solution, its brown colour changes to green due to the formation of iron (II) sulphate solution.



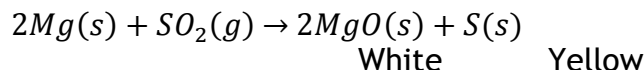
**NB//.** In this reaction iron (III) is reduced to iron (II) by gain of an electron.

- 5) **Oxidizing property.** It behaves as an oxidizing agent by adding oxygen to compounds.eg

(a) **With hydrogen sulphite.** Sulphurdioxide is reduced to sulphur which is seen as a yellow solid and hydrogen sulphite oxidized to water.



(b) **With magnesium.** When burning magnesium is lowered into a gas jar of sulphurdioxide, a yellow solid of sulphur and white solid of magnesium oxide are formed.



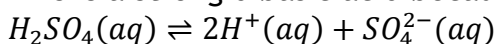
### Uses of sulphurdioxide gas

- Used in the contact process for the manufacture of sulphuric acid.
- Used to bleach wool, straw and sponges which are damaged by chlorine.
- Used in the manufacture of calcium hydrogen sulphide which is a bleaching agent for wood pulp.
- Used for formigation.

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### Sulphuric acid

This is a strong dibasic acid because it ionizes completely to produce many hydrogen ions.



### Contact process for the manufacture of sulphuric acid

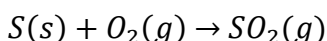
This is the process which leads to a large scale manufacture of sulphuric acid. The raw materials used include;

- Vanadium pentaoxide catalyst ( $V_2O_5$ )
- Concentrated sulphuric acid
- Water

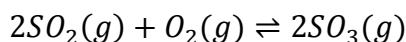
In the contact process the following steps are involved;

- Formation of sulphurdioxide gas
- Formation of sulphurtrioxide gas
- Formation of oleum ( $H_2S_2O_7$ )
- Formation of sulphuric acid from oleum.

(a) **Formation of sulphurdioxide.** This is obtained by burning sulphur in the presence of oxygen gas. Before burning, sulphur and oxygen are purified by removing dust from them which would otherwise poison the catalyst.



(b) **Conversion of sulphurdioxide to sulphurtrioxide.** The sulphurdioxide is reacted with excess oxygen gas and the mixture passed over finely divided vanadium pentaoxide which acts as a catalyst. This reaction is reversible and a lot of heat is given out (exothermic)

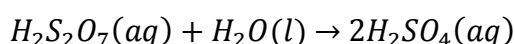
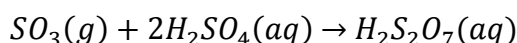


In order to obtain maximum yield of sulphurtrioxide the following conditions are applied;

- Use of vanadium pentaoxide catalyst to increase the rate of reaction.
- High pressure since the forward reaction proceeds with a reduction in volume.
- Low temperature because the forward reaction is exothermic. Generally temperature of 400-500°C is used.
- Excess oxygen gas

**NB//.** Platinum can also be used as a catalyst.

- (c) **Conversion of sulphurtrioxide to oleum.** The sulphurtrioxide formed is dissolved in concentrated sulphuric to form oleum. This oleum is removed and carefully diluted with correct amount of water to form ordinary concentrated sulphuric acid which is 98% and 2% water.



**NB//.** Sulphurtrioxide dissolved in water forms sulphuric acid with evolution of a lot of heat resulting into formation of mist of acid fumes and loss of the acid during production. This problem is overcome by dissolving sulphurtrioxide in concentrated sulphuric acid.

### Uses of sulphuric acid

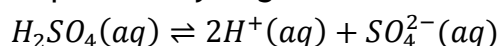
- Used in the manufacture of fertilizers e.g. ammonium sulphate.
- Used in the manufacture of detergents e.g. omo, nomi
- Used in the manufacture of paint and pigments of dyes.
- Used in the manufacture of man made fibers for making clothes e.g. nylon.
- Used in car batteries or accumulators as electrolyte.
- Used in making explosives.
- Used as a drying agent.
- Used in extraction of some metals.

### Physical properties of sulphuric acid

- Ordinary concentrated sulphuric acid is a colourless liquid.
- It is dense acid with density of 1.84 g/cm<sup>3</sup>
- It does not fume.

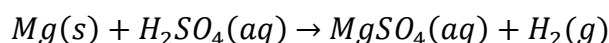
### Chemical properties of sulphuric acid

- 1) **Acidic properties.** It behaves as an acid when dilute because it ionizes completely to produce hydrogen ions.

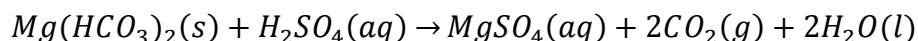
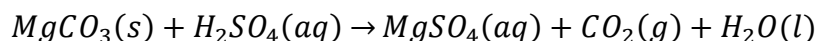


- (a) It turns blue litmus paper red

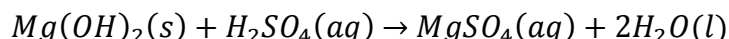
- (b) It reacts with metals above hydrogen in the electrochemical series liberating hydrogen gas with formation of a salt.eg



- (c) It liberates carbon dioxide from carbonates and hydrogen carbonates.eg

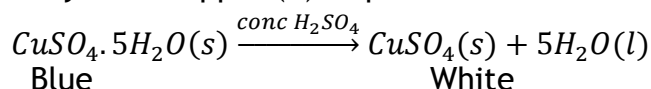


(d) It reacts with bases forming salts and water. eg

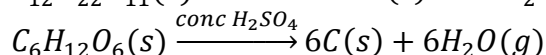
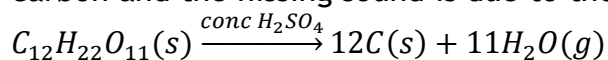


**2) Dehydrating properties.** It behaves as a dehydrating agent only when it is concentrated. This makes it remove water or elements of water from a substance. Examples include;

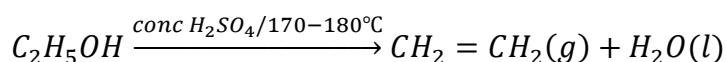
(a) **With hydrated copper (II) sulphate.** When put in beaker containing concentrated sulphuric acid, blue colour of the crystals become white due formation of anhydrous copper (II) sulphate after removal water.



(b) **With sucrose (C<sub>12</sub>H<sub>22</sub>O<sub>11</sub>) and glucose (C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>).** When sucrose and glucose are added to separate beakers of concentrated sulphuric acid, the white crystals turn brown with evolution of a lot of heat and steam is given off with hissing sound. Finally, a black solid that swells and fills up the beaker is formed. This solid is carbon and the hissing sound is due to the liberation of steam and heat.



(c) **With ethanol (C<sub>2</sub>H<sub>5</sub>OH).** When mixed with ethanol and heated to a temperature in the range of 170-180°C. Water is removed from the compound and a gas called ethane is formed.

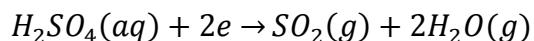


**NB//.** Concentrated sulphuric acid behaves as a dehydrating agent because it has a high affinity for water than any other acid. For this reason when the acid is put in a beaker and left exposed in air sometime, the level of the liquid in the beaker rises because the acid absorbs water vapour from the atmosphere. This can be demonstrated as below;

#### Diagram

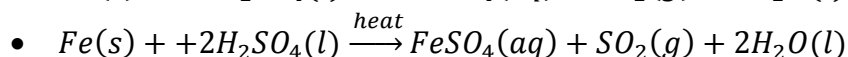
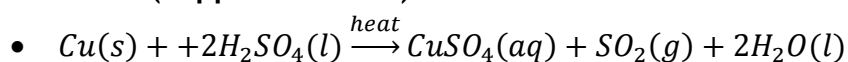
For this reason when diluting the acid, it is safer to add acid to water not water to acid because a lot of heat is evolved during the dilution process. This may cause the acid to splash out in form of vapour. However, if the acid is added to water this won't occur because acid molecules are denser than water molecules and penetrate to the bottom minimizing a lot of heat from being generated in one spot.

**3) Oxidizing properties.** It behaves as an oxidizing agent when concentrated. During the reaction it accepts electrons and itself reduced to Sulphur dioxide and water.



Examples

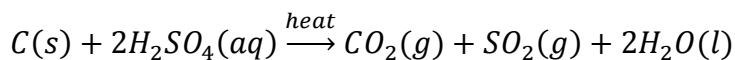
(a) **With metals (copper and iron)**



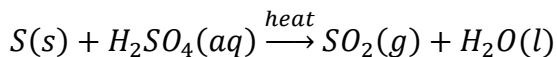
In the above reactions the metals are oxidized to metal sulphates (salts)

(b) **With non metals (carbon, sulphur and hydrogen sulphide).**

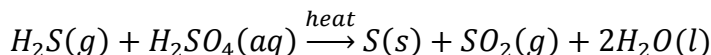
- Carbon. Addition of hot concentrated sulphuric acid to charcoal oxidizes the charcoal to carbon dioxide and itself reduced to sulphur dioxide and water.



- Sulphur



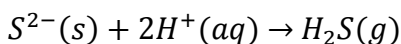
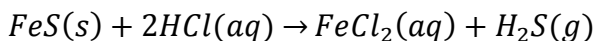
- Hydrogen sulphide. When bubbled through hot concentrated sulphuric acid, the acid is reduced to Sulphur dioxide and water and itself oxidized to sulphur.



## Hydrogen sulphide gas

### Laboratory preparation

It is prepared by reacting dilute hydrochloric acid with iron (II) sulphide.



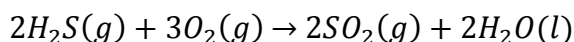
Diagram

### Physical properties of hydrogen sulphide gas

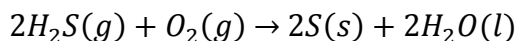
- It is colourless
- It is denser than air
- It has a characteristic smell of rotten egg
- It is fairly soluble in water and its solution in water is acidic i.e. turns blue litmus red.

### Chemical properties of hydrogen sulphide gas

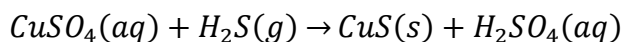
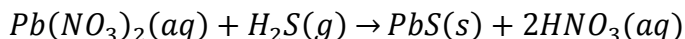
- 1) Reaction with oxygen (air).** It burns in excess air to form sulphur dioxide and water.



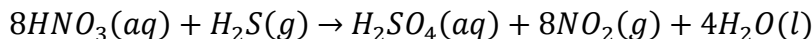
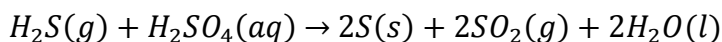
In limited supply of oxygen, yellow solid of sulphur and water are formed.



- 2) Reaction with salts.** When bubbled into a solution containing metal ion, a metal sulphide is formed which is seen as a black solid. eg

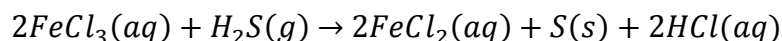


- 3) Reaction concentrated Sulphuric acid.**



- 4) Reaction with potassium permanganate solution.** Purple colour of the solution changes to colourless and a yellow solid surface of sulphur is formed.

- 5) **Reaction with iron (III) chloride solution.** Yellow colour of the solution changes to green due to formation of iron (II) chloride and yellow deposit of sulphur is formed.



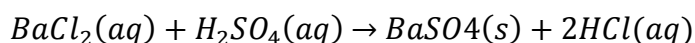
### Testing for hydrogen sulphide gas in the laboratory

Reagent	Procedure	observation
Lead acetate solution or lead (II) nitrate solution	• Filter paper soaked in lead acetate/lead (II) nitrate solution is dropped into a gas jar containing hydrogen sulphide gas	• The filter paper turns from white to black due to the formation of lead (II) sulphide.

### Sulphates

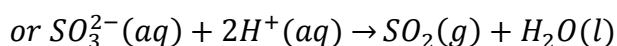
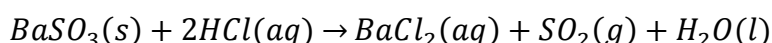
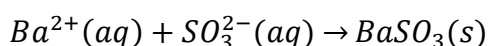
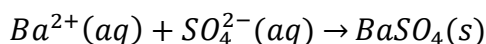
These are salts derived from sulphuric acid. All sulphates are soluble in water except lead (II) sulphate, calcium (II) sulphate and barium (II) sulphate.

Barium and lead (II) sulphates are prepared by double decomposition or precipitation.



### Testing for sulphates and sulphites

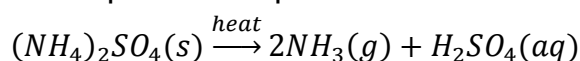
	Reagents	Procedure	Observations
$SO_4^{2-}$	Barium chloride dilute	Add barium (II) chloride followed by dilute hydrochloric acid	• White precipitate insoluble in the acid
$SO_3^{2-}$	hydrochloric acid or barium nitrate and dilute nitric acid	Or barium nitrate followed by dilute nitric acid	• White precipitate that dissolves with effervescence of a colourless gas that turns acidified potassium dichromate from orange to green.



**NB//.** The acid is to decompose any impurity present in the test solution.

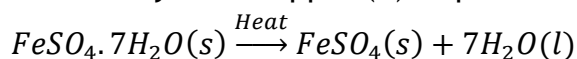
### Properties of sulphates

- 1) **Solubility.** All are soluble in water except lead (II) sulphate, calcium (II) sulphate and barium (II) sulphate.
- 2) **Action of heat.** All sulphates decompose on heating forming a metal oxide and Sulphur trioxide except iron (II) sulphate and ammonium sulphate.
  - Ammonium sulphate decomposes to sulphuric acid and ammonia gas.

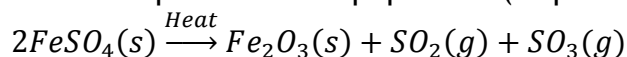




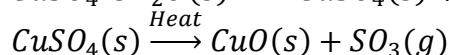
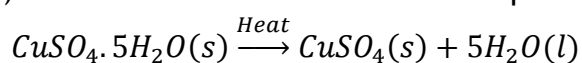
- Hydrated iron (II) sulphate on gentle heating gives colourless vapour which condenses to colourless liquid that turns blue cobalt (II) chloride paper pink or white anhydrous copper (II) sulphate blue.



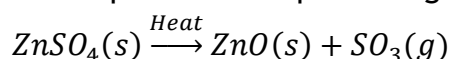
On strong heating, the pale green solid decomposes to reddish brown solid residue of iron (III) oxide, colourless gas that turns damp blue litmus paper red and acidified dichromate solution green (Sulphur dioxide) and white fumes that turns damp blue litmus paper red (sulphur trioxide)



- Hydrated copper (II) sulphate on heating gives;
  - Colourless vapour which condenses to colourless liquid that turns blue cobalt (II) chloride paper pink.
  - White solid which decomposes to black solid residue of copper (II) oxide
  - White fumes which turns damp blue litmus paper red.



- Zinc sulphate decomposes to give yellow solid when hot and white when cold.

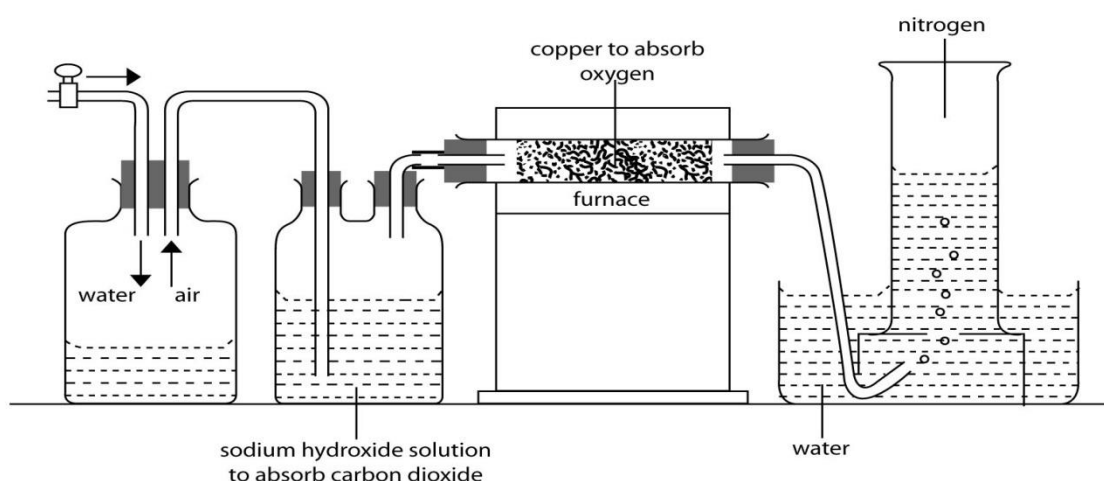


## Revision sample questions

## NITROGEN AND ITS COMPOUNDS

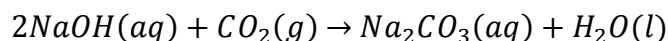
### Laboratory preparation of nitrogen gas

It is prepared from air. Air contains a mixture of carbon dioxide, oxygen, nitrogen and rare gases. To obtain nitrogen other air components are removed by the chemical process given below;

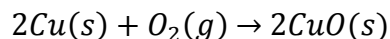


- Dust particles. They are removed from air by passing air through water which dissolves dust particles.
- Carbon dioxide. This is removed by passing air through concentrated sodium or potassium hydroxide.





- (c) Oxygen is removed by passing the mixture over heated copper metal in combustion tube. The brown colour of the metal turns to black due to the formation of copper (II) oxide.



Nitrogen obtained is denser than air as it contains rare gases which cannot be removed by any chemical means as they are unreactive.

Nitrogen is collected over water since it is insoluble in water.

## Industrial preparation of nitrogen gas

Nitrogen is obtained industrially by fractional distillation of liquid air. The process is carried out as below;

- Air (carbondioxide, oxygen, nitrogen gas and rare gases) obtained is purified by removing dust particles.
- The mixture is then compressed at high pressure and allowed to expand by escaping through a small jet. This makes the mixture to cool and liquefy. During the process of expansion and contraction, the air molecules loose there energies, attract each other to become liquid.
- The liquid air contains carbondioxide, oxygen and nitrogen gas. Since these components have different boiling points they can be separated by fractional distillation.
- Nitrogen with a lower boiling point than oxygen evaporates first leaving behind oxygen. The nitrogen obtained is further liquefied and stored.

## Uses of nitrogen

- Used in filling air balloons as it is lighter
- Used in Haber process for the manufacture of ammonia gas.
- Used for storage of semen, sodium metal because it provides an inert atmosphere or environment.

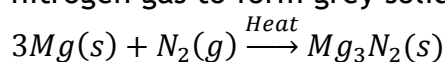
## Physical properties of nitrogen

- It is colourless and odourless.
- It is insoluble in water
- It is slightly denser than air because of presence of rare gases.
- Nitrogen gas does not have any effect on litmus paper.

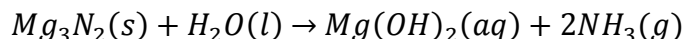
## Reactivity of nitrogen

Nitrogen gas is regarded as inert or unreactive compared to all other gases. This is because the nitrogen molecule has two nitrogen atoms joined by strong triple bonds which require a lot of energy to break during chemical reaction. However, under conditions of high temperature and pressure it reacts with the following;

- 1) **Magnesium metal.** It reacts when burning magnesium metal is lowered into gas jar of nitrogen gas to form grey solid of magnesium nitride.

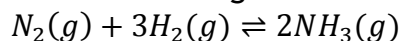


When magnesium nitride is treated with water, it dissolves to form magnesium hydroxide and ammonia gas.



2) **Oxygen/air**. It burns in air/oxygen to form white solid of magnesium oxide and magnesium nitride which is seen as grey solid.

3) **Hydrogen**. It combines with hydrogen at high pressure and moderate temperature to form ammonia gas. The reaction is catalyzed by finely divided iron.



## Compounds of nitrogen

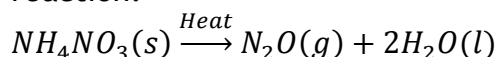
### Oxides

There are three types of oxides of nitrogen. Namely;

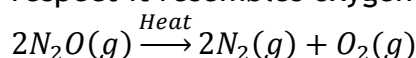
- Dinitrogen oxide (laughing gas).
- Nitrogenmonoxide (nitric oxide)
- Nitrogendioxide

1) **Dinitrogenoxide**. Sometimes it is called laughing gas because it produces temporary insensibility when inhaled.

It is prepared by strongly heating ammonium nitrate. Water is also formed during the reaction.



The gas is tested by using a burning splint i.e. it relights a burning splint and in this respect it resembles oxygen gas.



The burning splint decomposes to the dinitrogenoxide to nitrogen and oxygen of which oxygen relights a burning splint as shown above.

However, this gas can be distinguished from oxygen as it has a faint sweet smell while oxygen has no smell.

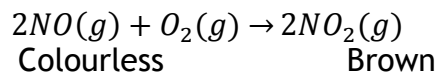
### Physical properties

- It is colourless
- Slightly soluble in water
- Neutral to litmus
- Slightly denser than air
- Has sweet faint smell
- Mainly used as an anesthesia during minor surgical operations because it reduces pain for a small period of time.

2) **Nitrogenmonoxide**. This gas is prepared in the laboratory by reacting moderately concentrated nitric acid (50%  $HNO_3$  and 50%  $H_2O$ ) with copper metal.



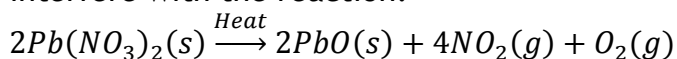
The presence of this gas is detected in the laboratory by exposing it to air i.e. reacts with air to form reddish brown fumes of nitrogendioxide.



## Physical properties

- Colourless gas
- Neutral to litmus paper
- Slightly denser than air and insoluble in water

**3) Nitrogendioxide.** The gas is prepared by strongly heating nitrates of heavy metals e.g. lead (II) nitrate because it does not contain water of crystallization which would interfere with the reaction.



This process can be carried out using the set up below.

## Diagram

When lead (II) nitrate is heated, the following observations are made;

- Decrepitates and melts to give reddish brown fumes containing a mixture of nitrogendioxide and oxygen gas.
- The reddish brown fumes passes through a U-tube put in a freezing mixture. At this point, the nitrogendioxide will cool and condense to form a yellow liquid which can be collected.
- Oxygen proceeds and is collected over water.

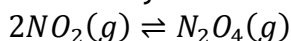
## Physical properties of nitrogendioxide gas

- Reddish brown gas
- Pungent irritating smell
- Very soluble in water forming nitric and nitrous acids. For this reason it is referred to as a mixed acid anhydride.



## Dissociation of nitrogendioxide

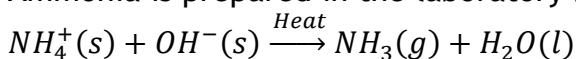
At low temperature, nitrogendioxide exists as yellow liquid. It becomes brown and eventually reddish brown indicating that it has dissociated.



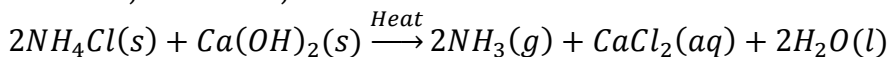
## Ammonia gas

## Laboratory preparation

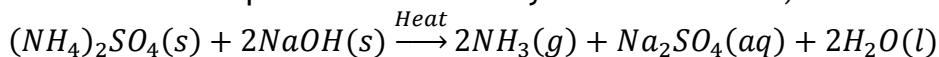
Ammonia is prepared in the laboratory by heating an ammonium salt with a hydroxide.



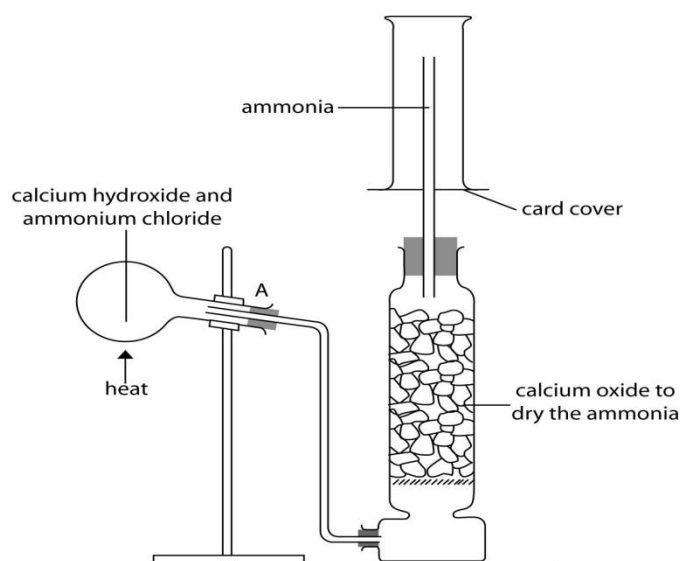
Generally, in the laboratory ammonium chloride and calcium hydroxide are used. In the reaction, ammonia, calcium chloride and water are formed.



If ammonium sulphate and sodium hydroxide are used;



The reactants must be in solid state and not in aqueous state because ammonia is a very soluble gas hence readily dissolves in the solution.



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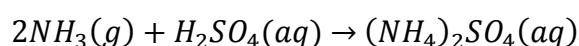
## Procedure

- A known amount of solid calcium hydroxide and solid ammonium chloride are mixed and ground into fine powder and put in a flask as shown above.
- The mixture is then heated strongly and ammonia gas is produced.
- The gas is passed through quick lime (calcium oxide) to dry it and finally collected by upward delivery since it is less dense than air.

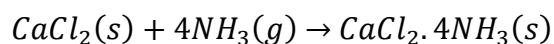
**NB//.** The flask is slanted in order to avoid the condensed water from running back into the reaction mixture which would otherwise dissolve all the ammonia gas produced.

Ammonia is only dried using calcium oxide and not the following substances;

- Concentrated sulphuric acid because ammonia is alkaline therefore reacts with the acid forming ammonium sulphate.



- Anhydrous calcium chloride because it reacts with it forming a complex compound.



## Chemical test for ammonia gas in the laboratory

Reagents	Procedure	Observations
Litmus paper	Dip a moist red litmus paper into a gas jar of ammonia	The red litmus paper turns to blue
Concentrated hydrochloric acid or dry hydrogen chloride gas	Pass ammonia gas through concentrated hydrochloric acid or hydrogen chloride gas	Dense white fumes of ammonium chloride is formed
$\text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s)$		

## Physical properties of ammonia gas

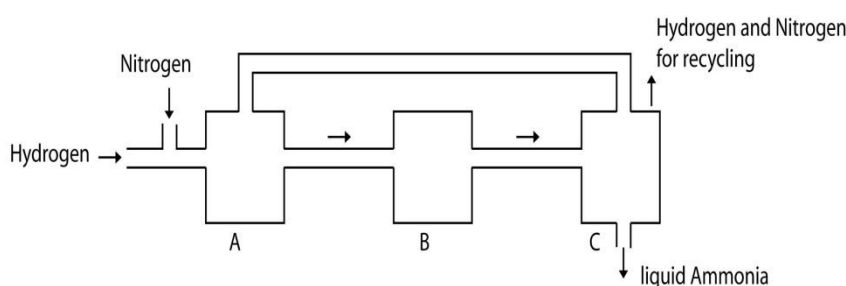
- Colourless
- Pungent choking smell
- Less dense than air hence collected by upward delivery.
- It is very soluble in water and its solution is called ammonium hydroxide.

## Industrial preparation of ammonia gas (Haber process)

### Raw materials

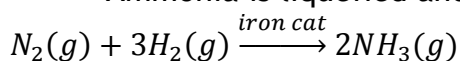
- Nitrogen obtained by fractional distillation of liquid air
- Hydrogen gas obtained by Bosch process
- Finely divided iron catalyst

### Chemical plant



### Procedure

- Nitrogen and hydrogen gas are mixed in the volume ratio of 1 : 3 and the mixture passed over finely divided iron catalyst at 500°C and high pressure. Ammonia gas is formed and a lot of heat is given out.
- The ammonia produced is separated from the unreacted nitrogen gas by cooling the mixture and the unreacted gases are recycled back into the converters where they are converted into ammonia.
- Ammonia is liquefied and stored.



### Conditions for the reaction

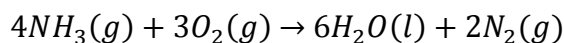
A good yield of ammonia is obtained under the following conditions;

- Finely divided iron catalyst to increase the rate of reaction.
- Since the forward reaction producing ammonia proceeds with a decrease in volume, high pressure will favour a good yield of ammonia.
- Since the reaction producing ammonia is exothermic, generally, low temperature is required (450-500°C).

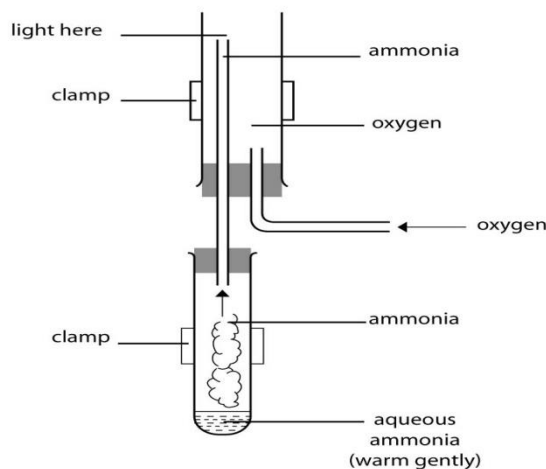
## Chemical properties of ammonia gas

### 1) Reaction with oxygen/air

(i) **In the absence of catalyst.** It burns with a greenish yellow flame to produce water and nitrogen gas.

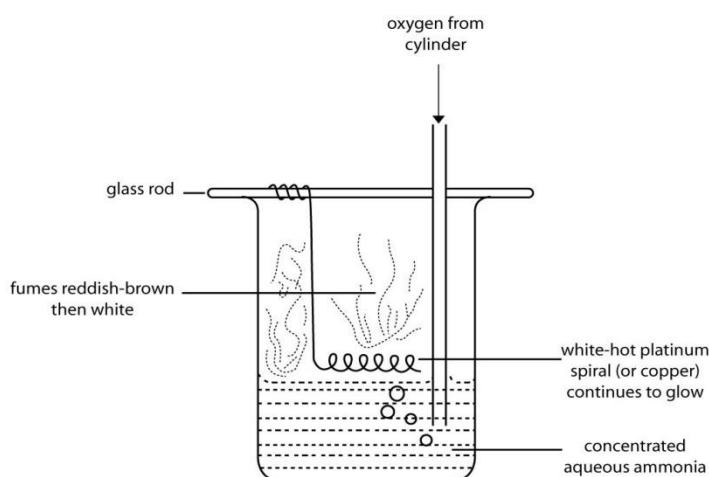


This can be demonstrated by the experiment below;



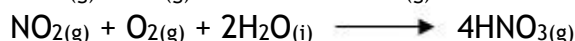
(ii) **In the presence of a catalyst.** The catalyst use is platinum or copper metal.  
**Procedure**

- Concentrated ammonia is put in a beaker. This solution releases ammonia gas.
- Copper or platinum wire is wrapped around glass rod and this wire is bent to hang just above the ammonia solution. Before this reaction, the wire is heated until red hot and place in a beaker.
- Oxygen gas is then passed through a glass tube into the solution and immediately the reaction begins. The reaction produces a lot of heat keeping to copper wire glowing throughout the experiment.



### Observations

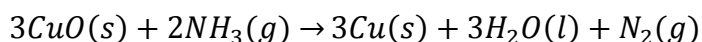
Reddish brown fumes of nitrogendioxide are seen to come out just above the spiral wire and later turned white due to the formation of ammonium nitrate. Spiral wire keeps glowing red hot through this experiment because the reaction exothermic.





2) **Reaction with metal oxides.** It reacts with oxides of lead and copper forming metal, nitrogen gas and water. This reaction shows that ammonia gas is a reducing agent.

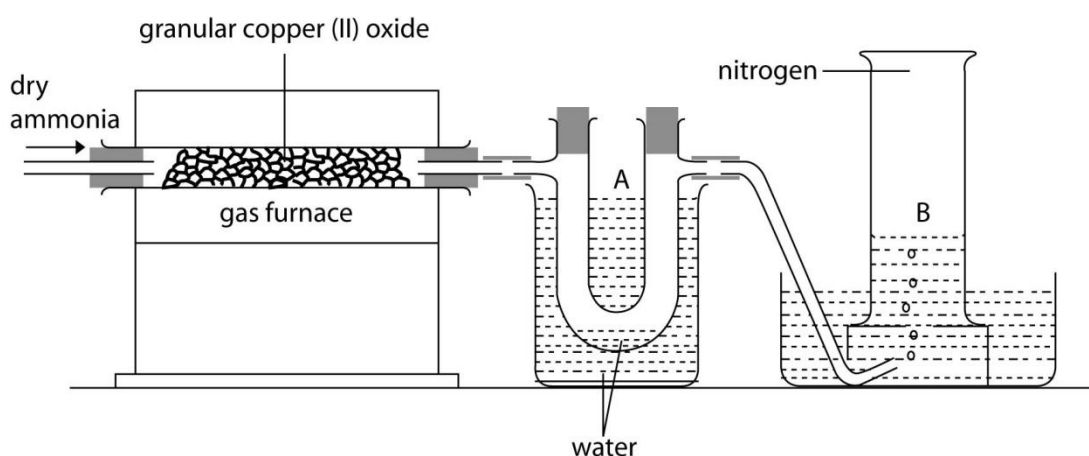
(a) **Copper (II) oxide.** Ammonia is oxidized to water and nitrogen gas and copper (II) oxide reduced to copper metal.



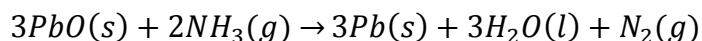
**Observations**

- Black solid of copper (II) oxide turns brown due to formation of copper metal.
- Colourless inert gas with no action on litmus paper formed.
- Colourless liquid condenses on cooler parts of the tube and turns blue cobalt (II) chloride paper pink or anhydrous copper (II) sulphate from white to blue.

This property is demonstrated using the setup below.



(b) **Lead (II) oxide.** Ammonia is oxidized to nitrogen and water and lead (II) oxide reduced to lead metal.

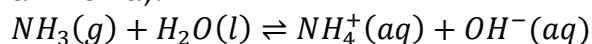


**Observations**

- Reddish brown solid of lead (II) oxide turns grey due to formation of lead metal.
- Colourless inert gas with no action on litmus paper formed.
- Colourless liquid condenses on cooler parts of the tube and turns blue cobalt (II) chloride paper pink or anhydrous copper (II) sulphate from white to blue.

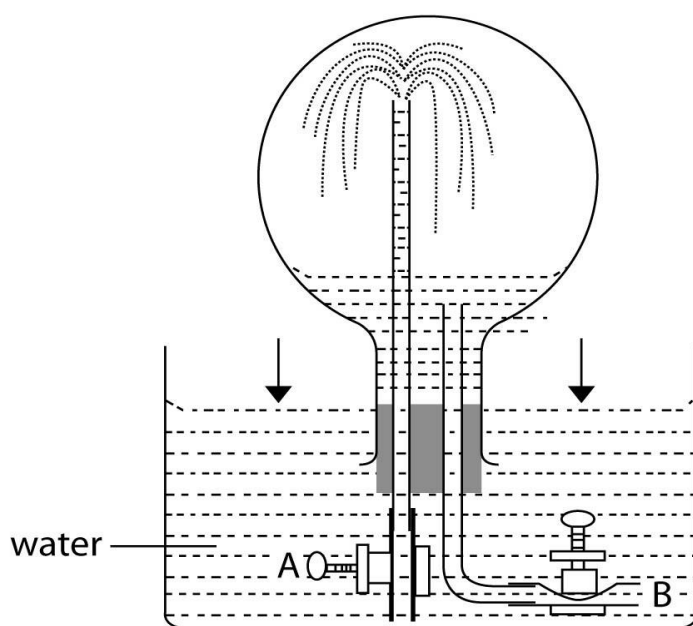
The experiment can be demonstrated using the setup above.

3) **Reaction with water (solubility in water).** Ammonia is considered to be the most soluble in water. Its solution in water is called ammonium hydroxide (aqueous ammonia).



When it reacts, it produces hydroxide ions which makes the solution alkaline ( $\text{pH} > 7$ ). For this reason, ammonia solution in water is regarded as a base. However, it is a weaker base compared to sodium hydroxide because it partially ionizes to produce a few hydroxyl ions.

Solubility of ammonia in water can be demonstrated by the fountain experiment.



### Procedure

- A thick walled glass is filled with ammonia gas.
- A rubber stopper carrying two tubes are fitted with clips A and B are inserted into the flask as shown above.
- The flask is then inverted in a trough of water.
- Clip B is opened to allow some few drops of water in the flask and immediately the clip is closed. Ammonia being very soluble is absorbed by the water droplets.
- This creates a partial vacuum in the flask i.e. gas pressure in the flask is reduced.
- Clip A is opened, a fountain is at once formed at the tip of long tube in form of a jet.

This fountain continues until the pressure in the flask and the outside have been equalized.

**NB//.** Hydrogen chloride gas gives the same observation.

- 4) **Reaction of ammonium hydroxide solution with metal ions.** Ammonium hydroxide is used in qualitative inorganic analysis to identify positive ions. During this reaction, a metal hydroxide which is either soluble or insoluble in excess ammonia is formed.

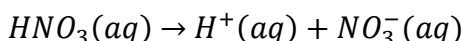
Metal ions	Observations
$Pb^{2+}, Zn^{2+}$	White precipitate soluble in excess forming a colourless solution $Pb^{2+} + 2OH^{-}(aq) \rightarrow Pb(OH)_2(s)$ In excess $Pb(OH)_2(s) + 2OH^{-}(aq) \rightarrow [Pb(OH)_4]^{2-}(aq)$ $Zn^{2+} + 2OH^{-}(aq) \rightarrow Zn(OH)_2(s)$ In excess $Zn(OH)_2(s) + 2OH^{-}(aq) \rightarrow [Zn(OH)_4]^{2-}(aq)$
$Cu^{2+}$	Blue precipitate soluble in excess giving deep blue solution $Cu^{2+} + 2OH^{-}(aq) \rightarrow Cu(OH)_2(s)$ $Cu(OH)_2(s) + 2OH^{-}(aq) \rightarrow [Cu(OH)_4]^{2-}(aq)$
$Fe^{2+}$	Dirty green precipitate insoluble in excess and turns brown on



	standing. $Fe^{2+} + 2OH^{-}(aq) \rightarrow Fe(OH)_2(s)$
$Fe^{3+}$	Brown precipitate insoluble in excess $Fe^{3+} + 3OH^{-}(aq) \rightarrow Fe(OH)_3(s)$
$Al^{3+}$	White precipitate insoluble in excess $Al^{3+} + 3OH^{-}(aq) \rightarrow Al(OH)_3(s)$
$NH_4^{+}$	No observable change

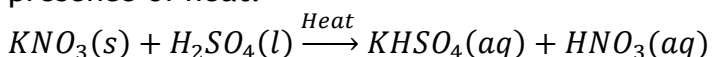
## Nitric acid

This is strong monobasic acid that ionizes completely to produce hydrogen ions and nitrate ions.



### Laboratory preparation

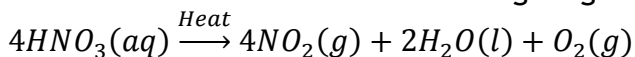
It is prepared by reacting concentrated sulphuric acid with potassium nitrate in the presence of heat.



**NB//.** This preparation is carried out in a special setup called retort glass. It is entirely made of glass because nitric acid is corrosive hence will attack cork, rubber and metals used in the setup.

### Diagram

Nitric acid produced is slightly decomposed by heat to brown fumes of nitrogendioxide which dissolves in the acid formed giving it a yellow colour.



**NB//.** Sodium nitrate is no used to prepare nitric acid because it is hygroscopic.

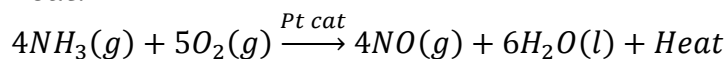
Hydrochloric acid and nitric acid are volatile.

Concentrated sulphuric acid used as it is non-volatile hence displaces volatile acids from their salts.

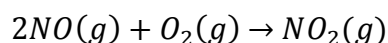
### Industrial preparation of nitric acid

It is obtained on a large scale by catalytic oxidation of ammonia gas. The catalyst used is platinum. The process involves three stages;

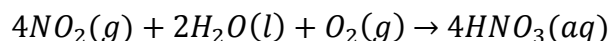
- 1) Ammonia gas is mixed with excess oxygen gas and passed over heated platinum catalyst to produce nitrogenmonoxide and water. The reaction produces a lot of heat.



- 2) The mixture produced is cooled and nitrogenmonoxide is mixed with excess oxygen to give reddish brown fumes of nitrogendioxide.



- 3) Nitrogenmonoxide formed is mixed with more oxygen and dissolved in warm water to produce nitric acid which contains 68% nitric acid and 32% water



## Conditions for reaction

- Presence of platinum catalyst
- High pressure (8 atmospheres)
- Excess oxygen gas

## Uses of nitric acid

- Making of plastics
- Manufacture of fertilizers e.g. ammonium nitrate, urea etc
- Used in the manufacture of explosives e.g. trinitrotoluene.

**NB//.** A mixture of concentrated nitric acid and concentrated hydrochloric acid in equal volume called aquaregia is used to test the presence of gold. All substances dissolve in the above mixture except gold.

## Physical properties of nitric acid

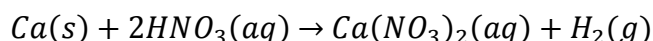
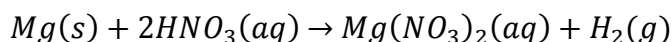
- It is colourless volatile liquid. However, it appears yellow in colour as it contains dissolved nitrogendioxide.
- It is very corrosive acid and this reason it is stored in glass containers not rubber, coke as they will react with them.

## Chemical properties of nitric acid

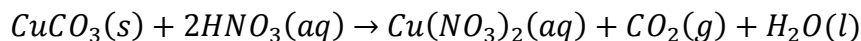
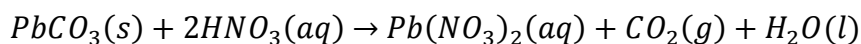
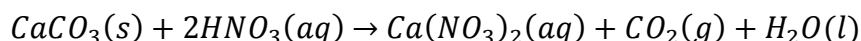
It is divided into two;

**1) Acidic property.** It behaves as an acid when dilute.

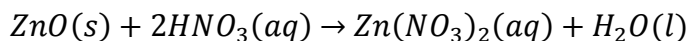
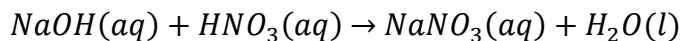
- It turns blue litmus paper red
- Liberates hydrogen gas when reacted with electropositive metals. Only calcium and magnesium react with dilute nitric acid.



- Librates carbondioxide from carbonates and hydrogencarbonates



- Reacts with bases forming salt and water.

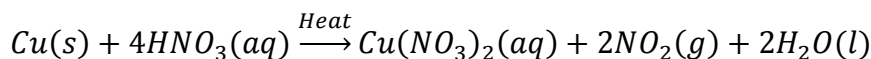


- Liberates hydrogensulphide gas from sulphites

**2) Oxidizing a property.** Nitric acid is a strong oxidizing agent and this takes place when the acid is concentrated. During the reaction, the acid is reduced to nitrogendioxide and water. Examples

**(a) Reaction with copper.**

- With concentrated nitric acid, in the presence of heat, copper is oxidized to copper (II) nitrate and acid reduced to nitrogendioxide and water.



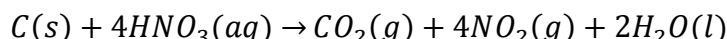
- With moderately concentrated nitric acid, nitrogen monoxide is formed instead.



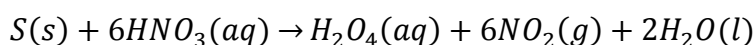
**NB//.** Nitric acid does not react with iron and aluminium during its oxidizing property because it renders the two metals passive by adding more oxide layer on the surface of the metals. The oxides prevent the acid from reacting with the metal.

**(b) Reaction with non metals.**

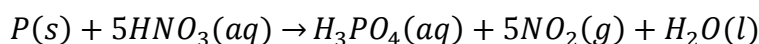
- **With carbon.** Concentrated nitric acid is reduced to water and nitrogendioxide and itself oxidized to carbondioxide.



- **With sulphur.** It oxidizes sulphur to sulphuric acid and itself reduced to nitrogendioxide and water.



- **With phosphorous.** It is oxidized to phosphorous acid.



## Nitrates

These are salts derived from nitric acid and contain the nitrate ion.

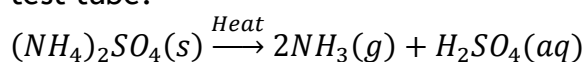
### Properties of nitrates

- 1) Solubility. All nitrates are soluble in water.
- 2) All nitrates exist as crystalline solids and are deliquescent i.e. absorbs water vapour when exposed to air and dissolves forming a solution. Other crystalline salts are;
  - Sulphates
  - Chlorides
  - Sulphites
  - Nitrates
  - Carbonates (powder form)
- 3) Action of heat. Nitrates are divided into four groups as far as action of heat is concerned.

Nitrate of;	Action of heat
$\text{NH}_4$	$\text{N}_2\text{O}$ and $\text{H}_2\text{O}$
$\text{K}$ $\text{Na}$	} <i>Metal nitrite + oxygen gas</i>
$\text{Ca}$ $\text{Mg}$ $\text{Al}$ $\text{Zn}$ $\text{Fe}$ $\text{Pb}$ $\text{H}$ $\text{Cu}$	



When heated, it sublimes and on cooling, forms white sublimate on the cooler part of the test tube.



**NB//.** Generally, when any ammonium salt is heated, ammonia gas is given off.

### Testing for ammonium ions in the laboratory

Reagent	Procedure	Observations
Sodium hydroxide solution	Add sodium hydroxide solution to the test solution and heat the mixture	Colourless gas with choking smell and turns damp red litmus paper blue and dense white fumes with concentrated hydrochloric acid evolved, $NH_4^+(aq) + OH^-(aq) \xrightarrow{\text{Heat}} NH_3(g) + H_2O(l)$ $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$

125

## CHLORINE AND ITS COMPOUNDS

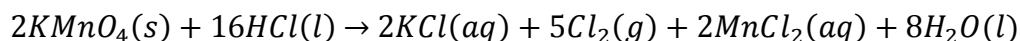
- Chlorine belongs to group 7 in the periodic table.
- It has a relative atomic mass of 35.5 and atomic number of 17 with a valency of 1.
- Other elements belonging to group 7 are fluorine, bromine, iodine and astatine.
- They are generally called halogens meaning salt producers.

### Laboratory preparation of chlorine gas

#### 1) Oxidation of concentrated hydrochloric acid by potassium permanganate.

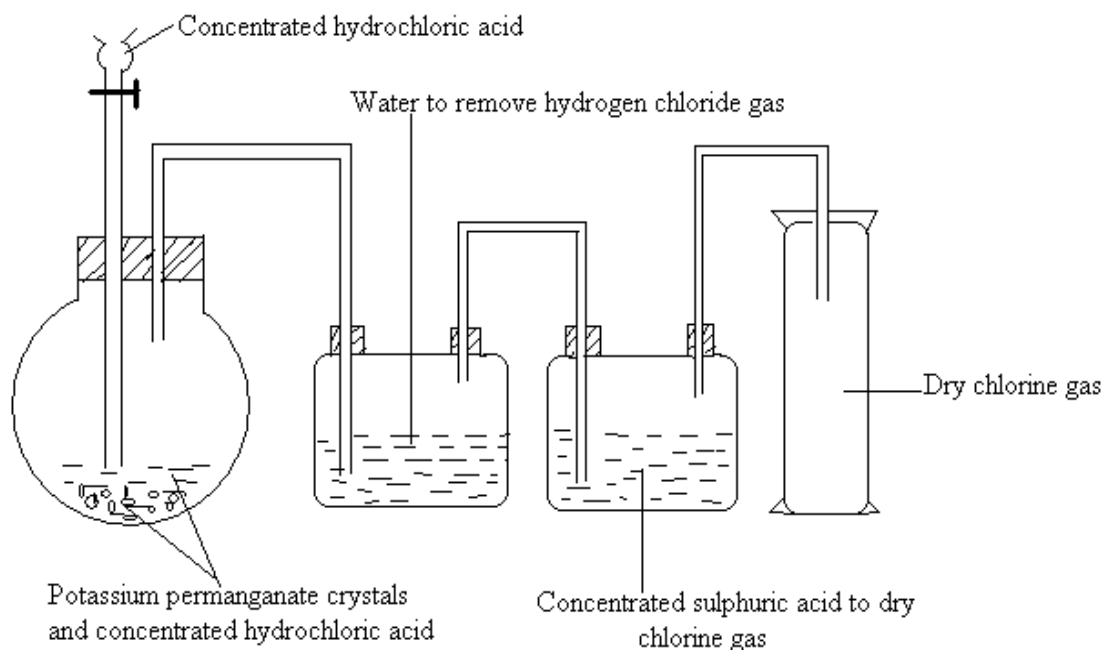
##### Procedure

- Potassium permanganate crystals are put in flask and concentrated hydrochloric acid added to it through a tap funnel.
- Immediately chlorine gas is evolved. This greenish yellow gas is passed through a wash bottle containing water which removes the spray of hydrogen chloride gas.



- Chlorine gas is finally passed through another wash bottle containing concentrated sulphuric acid to dry the gas.
- The gas is collected by downward delivery since it is denser than air.

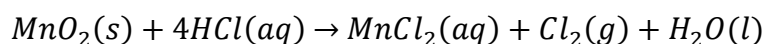
### Experimental setup



## 2) Oxidation of concentrated hydrochloric acid using manganese (IV) oxide.

### Procedure

- Manganese (IV) oxide is put in a flask with concentrated hydrochloric acid.
- The mixture is heated and chlorine is immediately evolved. Other products formed include manganese (II) chloride and water.

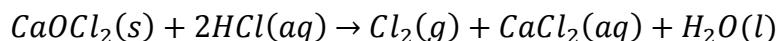
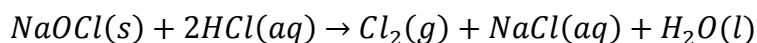


- The chlorine is passed through warm water to remove sprays of hydrogen chloride gas and the gas is dried by passing it through concentrated sulphuric acid.
- Chlorine gas is finally collected by downward delivery since it is denser than air.

**NB//** the diagram above can be used for the same preparation.

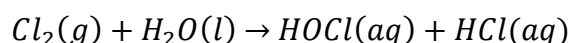
In the above reaction, manganese (IV) oxide is considered as an oxidizing agent and hydrochloric acid as a reducing agent as oxygen is taken from manganese (IV) oxide.

**NB//.** Chlorine can also be prepared by either reacting sodium hypochlorite or calcium hypochlorite with dilute hydrochloric acid.

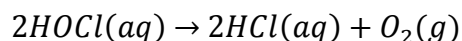


### Physical properties of chlorine gas

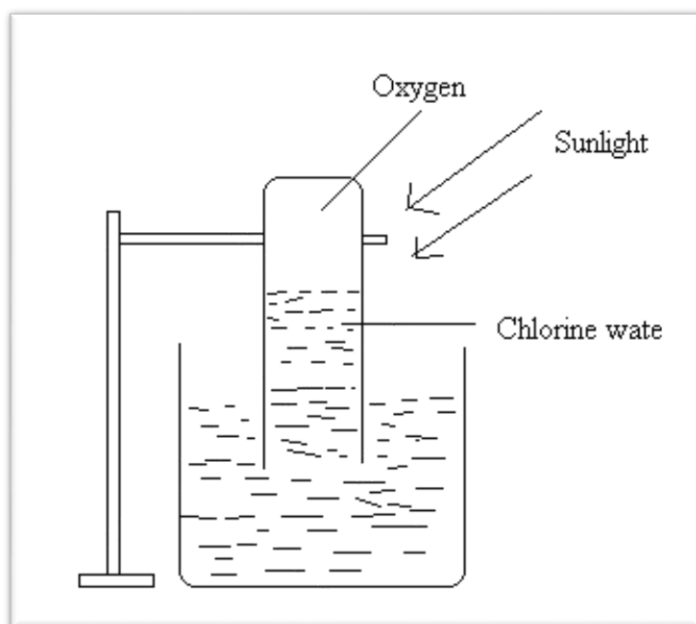
- Chlorine gas is greenish yellow gas
- It is denser than air therefore collected by downward delivery.
- It is fairly soluble in water forming hydrochloric acid and hypochlorous acid.



Hypochlorous acid is unstable and decomposes in the presence of sunlight to hydrochloric acid and oxygen gas.



This can be demonstrated by the setup below.



The above experiment shows that when chlorine water is exposed to sunlight, oxygen is produced. In this experiment, the test tube is inverted in a beaker of cold water and left in sunlight for sometime. A colourless gas collects at the end of the experiment to the close end of the tube above chlorine water. The gas relights a glowing splint.

Explanations

The chlorine and water reacts to form hydrochloric acid and hypochlorous acid. In the presence of sunlight, hypochlorous acid immediately decomposes giving oxygen gas.

### Chemical properties of chlorine gas

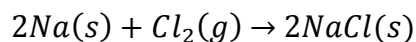
1) **Action on litmus paper.** When moist blue litmus paper is dropped into a gas jar containing chlorine gas, it turns red and finally bleached (the red colour of litmus turns white). In this reaction chlorine shows two properties;

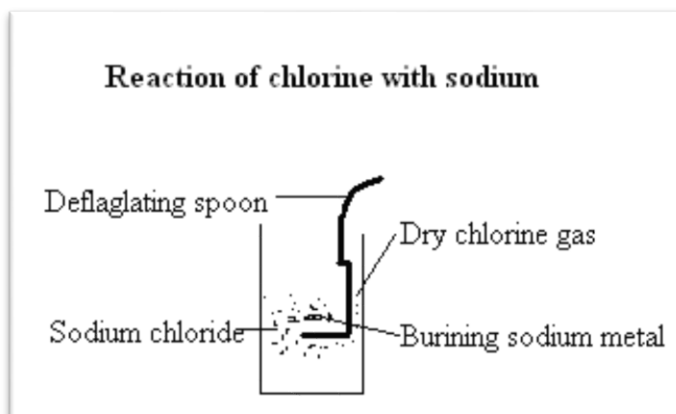
a) **Acidic property.** This is because it turns blue litmus to red.

b) **Bleaching property** since it removed the red colour of the litmus paper.

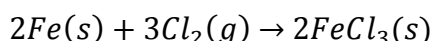
2) **Reaction with metals.**

a) **Sodium metal.** Dry chlorine combines directly with melted sodium forming white fumes which condenses to white solids of sodium chloride.

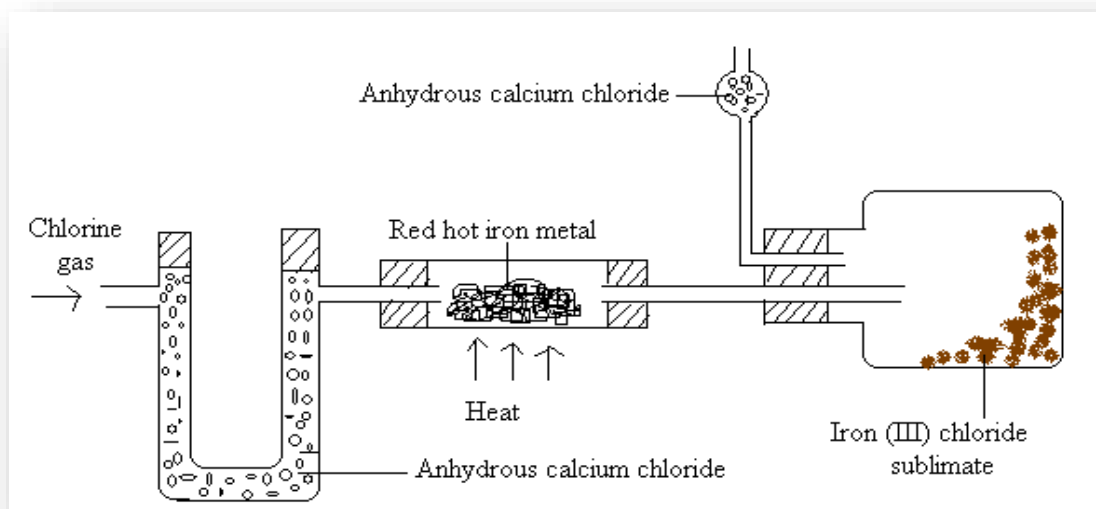




- b) **Iron metal.** Dry chlorine reacts with red hot iron to form iron (III) chloride. This compound is produced as a vapour in the combustion tube and later condensed to form black sublimate in the flask.



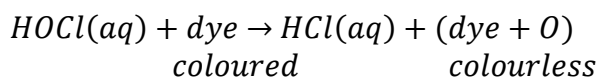
Anhydrous calcium chloride is used to remove water vapour as it enters the combustion tube because water vapour would readily react with iron (III) chloride formed.



Excess chlorine is allowed to leave the setup through the flask connected to a fume cupboard. The anhydrous calcium chloride is dry air which enters the setup through the fume cupboard.

**NB//.** The setup above can also be used to prepare iron (II) chloride by passing dry hydrogen chloride gas over heated iron.

- 3) **Bleaching action of chlorine.** Chlorine is a bleaching agent because it removes colour from compounds and turns them colourless. Dry chlorine gas does not bleach but when dissolved in water, it bleaches. This is due to the formation of hypochlorous acid formed when chlorine is dissolved in water. During bleaching, the hypochlorous acid gives its oxygen to the dye making the compound turn colourless.



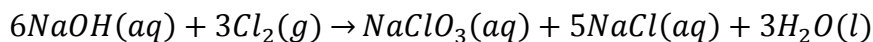
- 4) **Reaction with alkali (sodium hydroxide).** The product formed depends on the concentration of the alkali and condition of reaction. When chlorine gas is bubbled through;



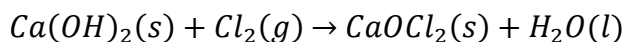
- a) **Cold dilute sodium hydroxide solution.** Yellow solution of sodium hypochloride and water are formed. This makes the resulting solution appears yellow.



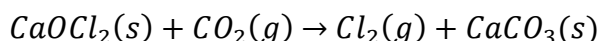
- b) **Hot concentrated sodium hydroxide solution.** Sodium hypochlorate, sodium chloride and water are formed.



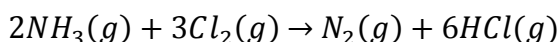
- c) **Calcium hydroxide.** Makes the colour and smell of gas to disappear due to its absorption (formation of calcium hypochloride)



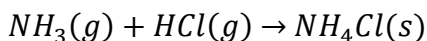
The bleaching powder smells of chlorine because it reacts with carbon dioxide in the air to give chlorine gas.



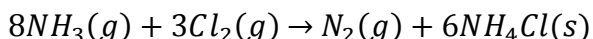
- 5) **Reaction with ammonia.** If a mixture of ammonia and chlorine gas is heated, hydrogen chloride gas and nitrogen gas are first formed.



But in the presence of excess ammonia, the hydrogen chloride gas formed combines to form dense white fumes of ammonium chloride.

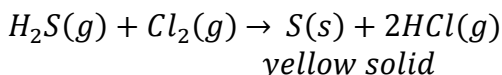


**Combined equation**

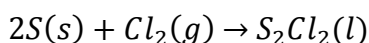


In the above reaction, chlorine acts as an oxidizing agent since it has removed hydrogen from ammonia hence reduced to hydrogen chloride.

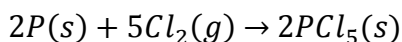
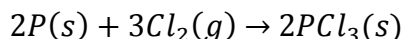
- 6) **Reaction with hydrogen sulphide.** When reacted with chlorine, yellow solid of sulphur is formed and hydrogen chloride gas is evolved. In this reaction, chlorine is an oxidizing agent as it removes hydrogen from hydrogen sulphide.



- 7) **Reaction with molten sulphur.** Passage of dry chlorine gas over molten sulphur results into formation of reddish liquid of disulphur dichloride.



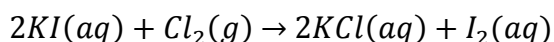
- 8) **Reaction with phosphorous.** Both red and white phosphorous when treated with chlorine gas forms white fumes of phosphorous trichloride and phosphorous pentachloride.



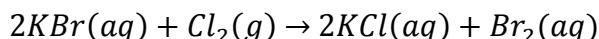
- 9) **Displacement reaction of chlorine**

This is a reaction in which chlorine displaces other halogens below it in the group. These halogens include bromine and iodine. The reaction can take place when chlorine is bubbled through a solution containing the ions of iodine and bromine.

- a) When chlorine is bubbled through potassium iodide solution, iodine is liberated which is seen as brown solution and potassium chloride is also formed.

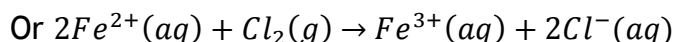
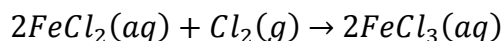


- b) When chlorine is bubbled through potassium bromine solution, the colour turns reddish brown due to the liberation of bromine solution.

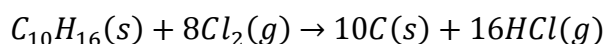


**10) Reaction with iron (II) chloride solution**

When chlorine is bubbled through iron (II) chloride solution, the pale green colour of the solution turns yellow (brown) due to the formation of iron (III) chloride solution

**11) Reaction with turpentine(C<sub>10</sub>H<sub>16</sub>)**

Turpentine is a hydrocarbon since it contains only carbon and hydrogen. On reacting it with chlorine gas forms carbon and hydrogen chloride gas. Carbon is deposited as black solid.



When candle wax is put in a gas jar of chlorine, black solids of carbon and white fumes of hydrogen chloride gas are formed. This shows that candle wax is made of hydrogen and carbon only.

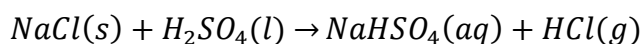
**Uses of chlorine gas**

- Used as disinfectant in the treatment of domestic water and swimming pools. This is because chlorine water contains hydrochloric acid killing germs.
- Used in the manufacture of insecticides eg DDT
- Used in the manufacture of plastics eg polyvinyl chloride
- Used in the manufacture of bleaching agents eg jik
- Chlorine its self is a bleaching agent.

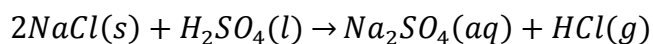
**Hydrogen chloride gas****Laboratory preparation**

It is prepared by the action of concentrated sulphuric acid on sodium chloride (normal salt).

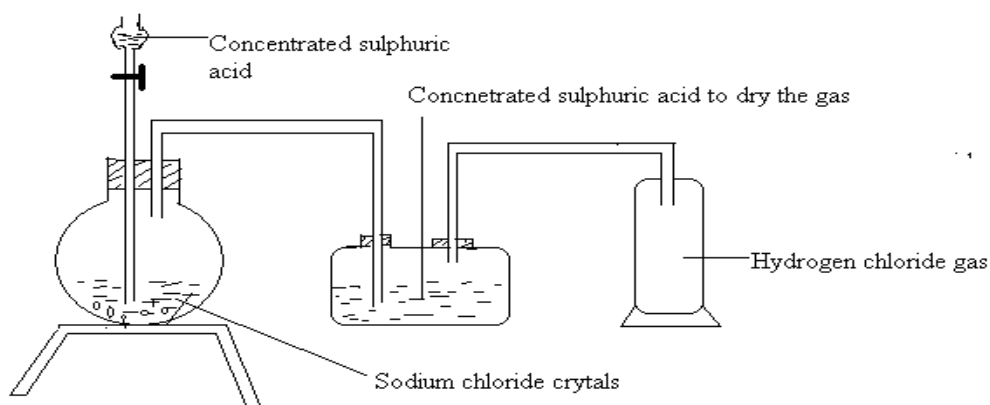
In the absence of heat, the salt formed is an acid salt (sodium hydrogen sulphate)



In the presence of heat, a normal salt (sodium sulphate) is formed.

**Procedure**

- Concentrated sulphuric acid is added to sodium chloride in a flask.
- Hydrogen chloride gas fumes are liberated and passed through a wash bottle containing concentrated sulphuric acid to dry it.
- The gas is then finally collected by downward delivery since it is denser than air.



### Physical properties of hydrogen chloride gas

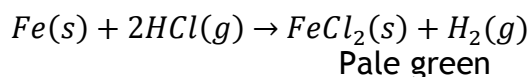
- It is colourless gas when dry but when moist it appears as misty fumes.
- It has a choking and irritating smell.
- It is denser than air.
- It is an acidic gas i.e. turns damp blue litmus paper red.
- It is very soluble in water. For this reason it is not collected over water.
- It does not support combustion.

### Chemical properties of hydrogen chloride gas

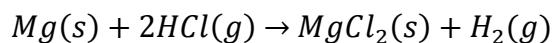
#### Reaction with metals

Dry hydrogen chloride gas reacts with heated metals to produce a metal chloride and hydrogen gas.

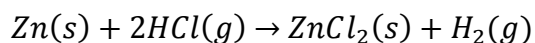
(a) With iron

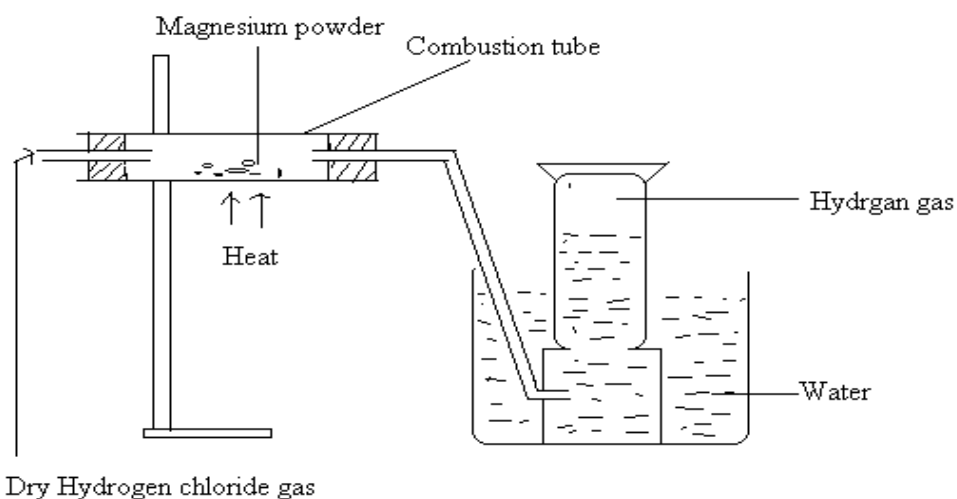


(b) With magnesium



(c) With zinc





### Solubility

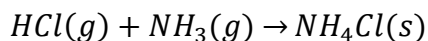
Hydrogen gas is highly soluble in water and a solution of this gas in water is called hydrochloric acid or aqueous hydrogen chloride.

**NB//.** Dry hydrogen chloride gas is a covalent compound and for this reason it does not conduct electricity when dry since it contains no ions.

### Chemical test for hydrogen chloride gas

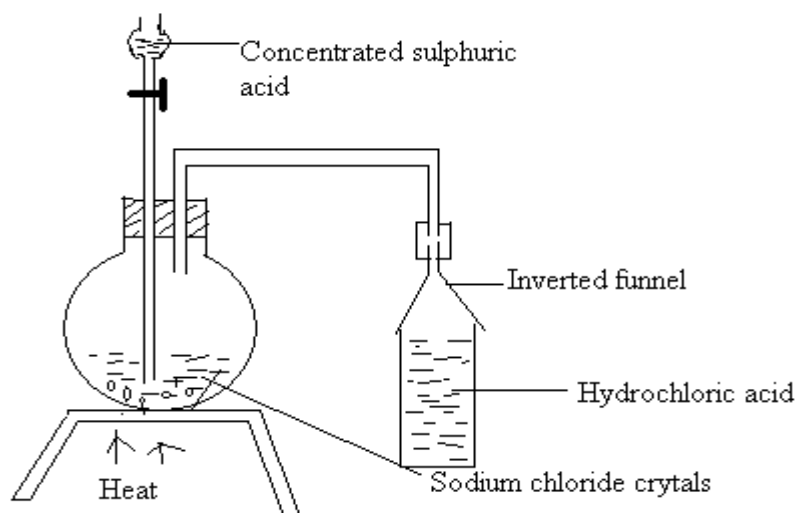
**Reagent.** Concentrated ammonia solution

**Observation.** Dense white fumes of ammonium chloride



### Hydrochloric acid

This acid is formed by bubbling hydrogen chloride gas in water.



Hydrogen chloride is dissolved in a beaker with cold water by means of an inverted funnel to avoid sucking back. Sucking back is caused as follows;

Hydrogen chloride is very soluble in water hence it is absorbed in water more quickly than it is being generated in the flask.

This results into decrease in pressure in the delivery tube and the flask therefore the atmospheric pressure outside forces the water back into the delivery tube and the flask, a process called **sucking back**.

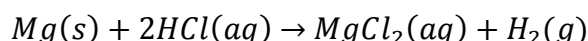
Towards the problem, an inverted funnel is used here as a considerable volume of water is required to fill the funnel before the narrow delivery tube is reached.

For effective preparation, the delivery tube must be in its rim just immersed in water.

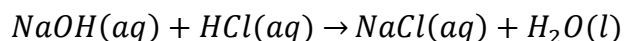
### Properties of hydrochloric acid

Like all other acids, it has the following properties;

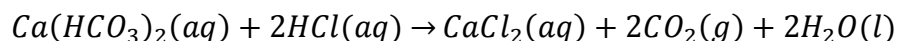
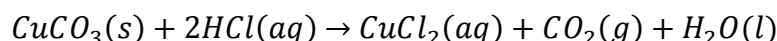
- It has sour taste.
- Turns blue litmus paper red.
- Reacts with electropositive metals to form salt and hydrogen gas i.e. metals above hydrogen in the reactivity series. Eg



- Reacts with bases to form salts and water. Eg



- Reacts with carbonates and hydrogen carbonates to form salt, water and carbondioxide. Eg



### Chlorides

**Solubility.** All chlorides are soluble in water except lead (II) chloride and silver chloride. NB//. Lead (II) chloride is soluble in hot water and insoluble in cold water.

**Action of heat.** All chlorides do not decompose when heated except ammonium chloride, iron (III) chloride which simply sublime when heated.



### Testing for the presence of chloride ions in the laboratory

Reagents	Procedure	Observations
Silver nitrate solution and dilute nitric acid	To a solution containing chloride ions in a test tube add dilute nitric acid followed by few drops of silver nitrate solution	White precipitate insoluble in acid
The white precipitate is silver chloride		
$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$		
NB//. Nitric acid is to decompose impurities which may react with silver nitrate solution.		

## EXTRACTION OF METALS

Metals do not occur in nature but combined with other elements to form compounds like oxides, sulphides, chlorides, sulphates etc. these compounds are the major sources of metals because metals can be extracted from them. They are given a general name ore. An ore is a naturally occurring substance from which a metal can be extracted. Eg

Ore	Metal that can be extracted
Bauxite (aluminum oxide)	Aluminum
Copper glance (copper (II) sulphide)	Copper
Common salt (sodium chloride)	sodium

There are three methods for extraction of metals

- (i) Electrolysis
- (ii) Reduction using carbon, carbonmonoxide and hydrogen
- (iii) Heating

The method chosen for a particular metal depends on its reactivity as given by the reactivity series;

K	}	They are very reactive hence form vey stable compounds and extracted by electrolysis since energy is provided
Na		
Ca		
Mg		
Al	}	Moderately reactive metals extracted by reduction of the ore
Zn		
Fe		
Pb		
H	}	Least reactive metals extracted by heating
Cu		
Hg		
Ag		
Au		

Extraction of metals is chemically considered to be a reduction process because the metal ions within the ore are positively charge. Hence to convert them into neutral atoms, electrons have to be added to the ions.

### Extraction of sodium

Sodium belongs to group one and period three in the periodic table. This metal is extracted from common salt (sodium chloride) and is mainly found in lakes, oceans etc. the process by which this metal is extracted is called **Down's process**.

### Process of extraction

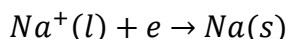
During the extraction, solid sodium chloride is strongly heated to form molten sodium chloride which is then put in the Down's cell.

This cell consists of carbon acting as the anode and iron as the cathode.

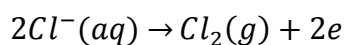
Calcium chloride is added to the molten sodium chloride to reduce or lower the melting point of sodium chloride from 800-600°C.

The ions present in the molten solution are sodium ions and chloride ions. Therefore the sodium ions migrate to the cathode where they are discharged by gain of electrons which

is stored under liquid nitrogen. Liquid nitrogen is used because it provides an inert environment thereby preventing sodium from reacting with air.



Chloride ions migrate to the anode where they are discharged to form chlorine gas as a byproduct.



Carbon is used as the anode because it is inert therefore does not react with chlorine. The major problem encountered during manufacture of sodium is chlorine produced as byproduct will pollute the air.

### Uses of sodium

- Used to make chemical compounds such as sodium cyanide, extraction of gold.
- Used to make sodium hydroxide
- Used to make sodium vapour lamps for street lighting.
- Used to make anti-knock compounds internal combustion engines.
- Used in nuclear reactor as a coolant.

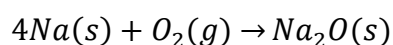
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### Physical properties of sodium

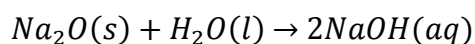
- It is a soft metal with low melting point and boiling point. This is because of the weak metallic bonds between its atoms.
- It has low density.
- It is a good conductor of electricity.

### Chemical properties of sodium

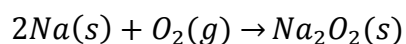
- 1) **Reaction with air.** When left exposed in moist air, it reacts slowly with oxygen to form a thin film of sodium monoxide deposited on the metal surface and for this reason it is kept under oil to prevent this reaction.



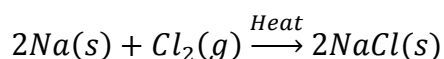
When exposed for a long time in air, the oxide reacts with water forming sodium hydroxide solution which absorbs carbon dioxide forming sodium hydrogen carbonate.



But when burnt in air, it gives out a bright yellow flame and a pale yellow solid of sodium peroxide. This is only seen when oxygen is in excess.



- 2) **Reaction with chlorine.** Sodium continues to burn with a bright golden yellow flame to form a white solid of sodium chloride.



## Manufacture of sodium hydroxide

It is manufactured by electrolysis of concentrated sodium chloride solution. This process is carried out in a special cell called mercury cell where flowing mercury acts as the cathode and the anode is made of graphite. (Refer to electrolysis of concentrated sodium chloride)

The process of producing sodium hydroxide on large scale has a bad effect on the environment as below;

- Chlorine gas given off is poisonous leading to air pollution.
- Mercury is poisonous metal because it leads or may cause cancer if it leaks into the surrounding water bodies.
- Sodium hydroxide may cause change in pH of the surrounding soils which is possible when it leaks to the surrounding areas i.e. pH changes from acid to alkaline.

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## Uses of sodium hydroxide

- Used in the manufacture of soap and detergents.
- Used in analytical practicals to find composition of compounds.
- Used in the manufacture of other compounds.

## Reaction of sodium hydroxide with metal ions

When sodium hydroxide solution is added to a solution containing a metal ion in a test tube drop wise until in excess, a metal hydroxide is precipitated. Some of the metal hydroxides are insoluble in excess while others are soluble in excess.

Observations	Deductions
White precipitate formed which is insoluble in excess sodium hydroxide	$\text{Ca}^{2+}$ or $\text{Mg}^{2+}$ probably present
White precipitate formed which dissolves in excess sodium hydroxide solution forming a colourless solution	Probably $\text{Zn}^{2+}$ , $\text{Al}^{3+}$ or $\text{Pb}^{2+}$ ions present $\text{Pb}^{2+} + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Pb}(\text{OH})_2(\text{s})$ In excess $\text{Pb}(\text{OH})_2(\text{s}) + 2\text{OH}^{-}(\text{aq}) \rightarrow [\text{Pb}(\text{OH})_4]^{2-}(\text{aq})$ $\text{Zn}^{2+} + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Zn}(\text{OH})_2(\text{s})$ In excess $\text{Zn}(\text{OH})_2(\text{s}) + 2\text{OH}^{-}(\text{aq}) \rightarrow [\text{Zn}(\text{OH})_4]^{2-}(\text{aq})$ $\text{Al}^{3+} + 3\text{OH}^{-}(\text{aq}) \rightarrow \text{Al}(\text{OH})_3(\text{s})$ In excess $\text{Al}(\text{OH})_3(\text{s}) + \text{OH}^{-}(\text{aq}) \rightarrow [\text{Al}(\text{OH})_4]^{-}(\text{aq})$
Blue precipitate formed which is insoluble in excess sodium hydroxide solution	$\text{Cu}^{2+}$ ions present $\text{Cu}^{2+} + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s})$
Dirty green precipitate formed which is insoluble in excess sodium hydroxide solution	$\text{Fe}^{2+}$ ions present $\text{Fe}^{2+} + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_2(\text{s})$



Reddish brown precipitate formed which is insoluble in excess sodium hydroxide solution	$\text{Fe}^{3+}$ ions present $\text{Fe}^{3+} + 3\text{OH}^-(\text{aq}) \rightarrow \text{Fe}(\text{OH})_3(\text{s})$
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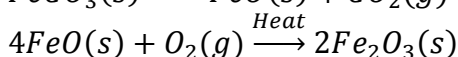
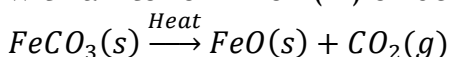
## Extraction of iron

The major ore from which iron can be extracted include the following;

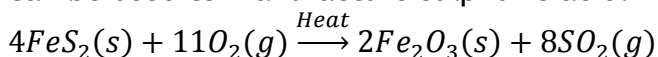
- Haematite (iron (III) oxide),  $\text{Fe}_2\text{O}_3$
- Magnetite (tri-iron tetraoxide),  $\text{Fe}_3\text{O}_4$
- Spathic iron (iron (II) carbonate),  $\text{FeCO}_3$
- Iron pyrites,  $\text{FeS}_2$

If iron (II) carbonate or iron pyrites are used, they are first converted to iron (III) oxide as shown below;

If iron (II) carbonate is used, it is strongly heated to form iron (II) oxide and then reacted with air to form iron (III) oxide which can be used for the extraction.



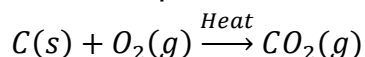
If iron pyrite is used, it is roasted in air to form iron (III) oxide and sulphurdioxide which can be used to manufacture sulphuric acid.



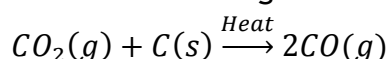
Extraction of iron takes place in a special plant called the blast furnace. This is a tall building approximately 11 meters high. Its wall is made of iron metal and inside the wall is lined with fire bricks.

### Diagram

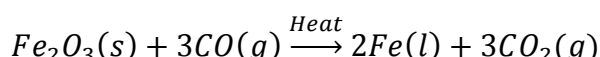
- The raw materials include; Haematite, Limestone and Coke
- The iron ore, limestone and coke are fed into the blast furnace through a pulley system called the cones and hoppers.
- A blast of hot compressed air at  $1600^\circ\text{C}$  is fed into the furnace through the bottom using special pipes called the tuyeres.
- As the hot air rises up the furnace, it meets coke and burns to produce carbondioxide and a lot of heat is liberated. The heat is used to carryout the reduction process in the blast furnace.



- The carbondioxide produced rises and meets hot coke and is reduced to carbonmonoxide gas which is the major reducing agent.

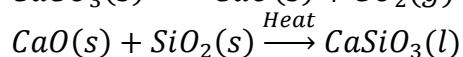
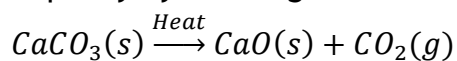


- At  $700^\circ\text{C}$ , within the blast furnace reduction iron (III) oxide takes place to form molten iron and carbondioxide.



- The molten iron drips to the bottom of the furnace where it is removed and cooled to form iron bars. The iron has a lot of impurities. For this reason it is called pig iron.

- The limestone at the same time decomposes due to the high temperature to calcium oxide and carbondioxide. Calcium oxide removes silica (acidic) present as impurity by reacting with it to form calcium silicate.



- The molten slag (calcium silicate) drips down to the bottom of the furnace and since it is less dense than the molten iron, it floats on molten iron thereby preventing iron from being oxidized.

### Uses of pig iron (cast iron)

- Used to make drainage pipes
- For making Bunsen burner bases
- For making of stoves, railway lines

### Wrought iron

This is pure form of iron obtained from pig iron by removing impurities. In this process pig iron is heated with haematite. The wrought iron is ductile and malleable unlike pig iron which is brittle. For this reason it is used for making iron sheets, hoes, coffins, nails.

### Steel

Steel is an alloy of iron and carbon where the carbon content has been carefully controlled. Other metals can be added to steel to make it special such as chromium, nickel, manganese. This form of alloy is called stainless steel and is resistant to corrosion hence used for making cutlery.

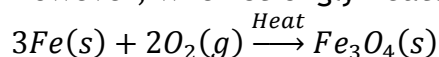
### Perm alloy

- Used to make electromagnets

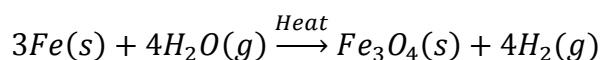
### Chemical properties of iron

- 1) **Reaction with air.** When exposed to moist air, it reacts slowly with oxygen in the presence of water forming rust chemically called hydrated iron (III) oxide reddish brown in colour.

However, when strongly heated in air it forms tri-iron tetraoxide.



- 2) **Reaction with water.** Iron does not react with cold water but when heated it reacts with steam to form tri-iron tetraoxide and hydrogen gas.



- 3) **Reaction with acids.** It liberates hydrogen gas when reacted with dilute acids and iron (II) salt is formed.

**NB//.** Iron does not react with nitric acid as it is rendered passive. Other metals that do not react with nitric acid are aluminium and chromium.

## Extraction of copper

Copper can be extracted from the following ores;

- Copper pyrite,  $\text{CuFeS}_2$
- Copper glance (copper (I) sulphide),  $\text{Cu}_2\text{S}$
- Copper (I) oxide
- Basic copper (II) carbonate,  $\text{CuCO}_3 \cdot \text{Cu(OH)}_2$

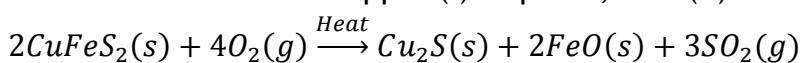
## Extraction process

This takes place in three stages;

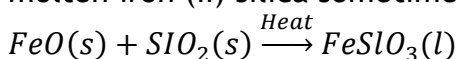
- Concentration of the ore i.e. removing impurities
- Roasting of the concentrated ore to get impure copper
- Refining impure copper (blister copper) using electrolysis.

(a) **Concentration of the ore.** The copper ore is concentrated by a process called froth floatation. In the froth floatation, the ore is ground into fine powder, dissolved in water containing a frothing agent and air is blown through the mixture to agitate the mixture. During this process, the impurities sink to the bottom of the floatation tank while the ore rises to the surface of mixture. The ore is skimmed off from the surface, filtered and dried.

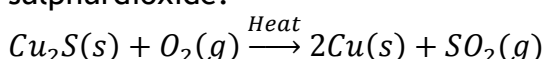
(b) **Roasting of the concentrated ore.** The concentrated ore is roasted in a limited amount of air to form copper (I) sulphide, iron (II) oxide and sulphurdioxide gas.



The mixture is heated and iron (II) oxide reacts with silica (silicon dioxide) to form molten iron (II) silica sometimes called slag.



Slag is removed from the mixture and the remaining copper (I) sulphide is treated by heating it in a controlled amount of air forming blister copper (impure copper) and sulphurdioxide.

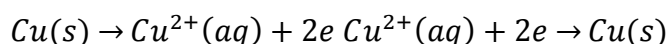


(c) **Refining impure copper (blister copper).** In this process, the blister copper is made the anode, pure copper the cathode and copper (II) sulphate solution as the electrolyte.

Copper atoms from blister copper dissolve and enter the electrolyte as copper (II) ions. The copper (II) ions migrate to the cathode and gain electrons forming copper.

**Anode**

**Cathode**



The anode decreases in weight while the cathode increase. Copper is obtained is pure and can be used for the following purposes;

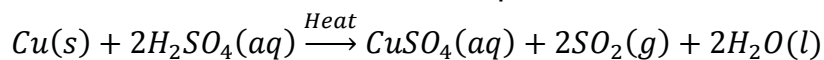
- Making of electric cables or electricity transmission.
- Used to make bullet shells.
- For making coins.
- For making alloys e.g. brass.

## Chemical reactions of copper

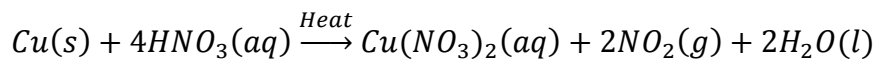
Copper does not react with water and dilute acids because it is lower than hydrogen in the electrochemical series.

However, it reacts with concentrated acids.

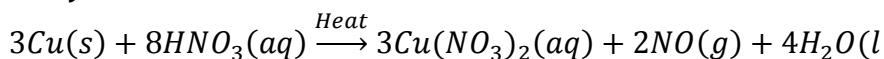
(a) **With concentrated sulphuric acid.** Copper is oxidized to copper (II) sulphate and the acid reduced to water and sulphurdioxide.



(b) **With concentrated nitric acid.**



(c) **With moderately concentrated nitric acid.**



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## THERMOCHEMISTRY

This is the study of heat changes that takes place during a chemical reaction. During a chemical reaction, heat is either absorbed or liberated from or into the environment. The heat content of a given substance is called enthalpy denoted by capital H.

### Enthalpy change/heat change.

This is the heat change during a chemical reaction given by the expression below;

$$\Delta H = \text{Enthalpy of products} - \text{Enthalpy of reactants}$$

$$\Delta H = H_p - H_R$$

### Types of reactions

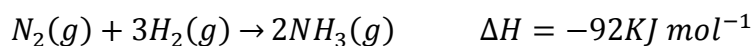
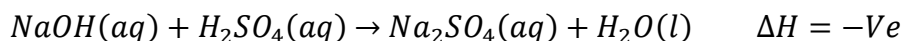
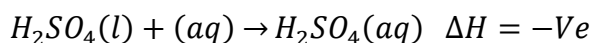
There are two types of reactions depending on the heat changes;

- Endothermic reactions
- Exothermic reactions

(a) **Endothermic reactions.** These reactions proceed by absorption of heat from the surrounding. They are assigned a positive enthalpy change. Eg



(b) **Exothermic reactions.** These reactions proceed by liberation of heat to the surrounding. They are given a negative sign for the enthalpy change.eg



Generally there are two factors that affect the quantity or amount of heat absorbed or liberated during a chemical reaction.

- State of the reactants
- Pressure of the system

### Relationship between heat and temperature

When a substance gains heat, its temperature rises and when it loses heat, its temperature drops. The quantity of heat lost or absorbed is given by the relationship below;

$$\theta = MC\Delta t$$

Where  $\theta$  = Quantity of heat absorbed or lost

$M$  = Mass of the substance

$\Delta t$  = Temperature change, °C/K

$C$  = Specific heat capacity of the substance

## Types of heat changes

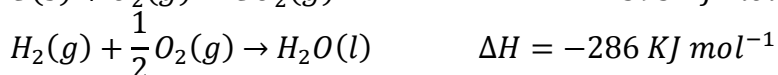
There are four types of heat changes depending on the reaction that takes place.

- Heat/enthalpy of combustion
- Heat/enthalpy of neutralization
- Heat/enthalpy of solution
- Heat/enthalpy of displacement

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## Enthalpy of combustion

This is heat liberated when one mole of a substance is completely burnt in oxygen. Eg



## Experiment to determine enthalpy of combustion of ethanol

Diagram

### Procedure

- Some ethanol is added to a spirit lamp or bottle fitted with a wick.
- Weigh the bottle together with its content and record its mass,  $m_1$ , before burning.
- Put a known volume of water,  $V_1 \text{ cm}^3$ , in a thin metal can as arranged in the diagram above.
- Take note of the initial temperature,  $T_0^\circ\text{C}$ , of water in the can.
- Light the spirit bottle wick and using the flame from the burning ethanol, heat the water in the can while stirring the water gently with the thermometer.
- Allow the temperature of the water to rise and put out the flame after the highest temperature,  $T_1^\circ\text{C}$ , has been achieved.
- Reweigh the specimen bottle plus its contents after burning and record its mass,  $m_2 \text{ g}$ .

### Calculations

Temperature change,  $\Delta t = (T_1 - T_0)^\circ\text{C}$

Mass of ethanol burnt,  $M = (m_1 - m_2) \text{ g}$

Specific heat capacity of water =  $C$

Mass of water, =  $V_1 g$  assuming density of water is  $1 g/cm^3$

Heat given during burning = Heat absorbed by water

$$= MC\Delta t$$

$$= V_1 \times C \times (T_1 - T_o) J$$

Molar mass of ethanol,  $C_2H_5OH = (12 \times 2) + 5 + 16 + 1 = 46 g/mol$

Number of moles of ethanol burnt =  $\frac{(m_1 - m_2)}{46}$  moles

$\frac{(m_1 - m_2)}{46}$  moles of ethanol liberates  $V_1 \times C \times (T_1 - T_o) J$  of heat

1 mole of ethanol burns to produce  $\frac{1 \times V_1 \times C \times (T_1 - T_o)}{(m_1 - m_2)} \times 46 J$  of heat

Therefore the heat of combustion of ethanol =  $\frac{1 \times V_1 \times C \times (T_1 - T_o)}{(m_1 - m_2)} \times 46 J$

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### Worked examples

- 1) In an experiment to determine the enthalpy of combustion of propan-1-ol ( $C_3H_7OH$ ), 0.54 g of propan-1-ol was burnt and the heat evolved caused the temperature of  $150 cm^3$  of water to rise by  $21.5^\circ C$ . Calculate the value of enthalpy of combustion.

#### Solution

Heat given during burning = Heat absorbed by water

$$= MC\Delta t$$

$$= 150 \times 4.2 \times 21.5 = 135.45 J$$

Molar mass of propanol,  $C_3H_7OH = (12 \times 3) + 7 + 16 + 1 = 60 g/mol$

Number of moles of propanol burnt =  $\frac{0.54}{60} = 0.009$  moles

0.009 moles of ethanol liberates 135.45 J of heat

1 mole of propanol burns to produce  $\frac{1 \times 135.45}{0.009} = 1505000 J$  of heat

Therefore the heat of combustion of propanol = 1505 KJ

- 2) When 0.4 g of ethanol was burnt, it raised the temperature of 0.1 g of water by  $20^\circ C$ . Calculate the enthalpy of combustion by taking into consideration that the specific heat capacity of water is  $4.2 KJ/Kg/^\circ C$ .

#### Solution

Heat given during burning = Heat absorbed by water

$$= MC\Delta t$$

$$= 0.1 \times 4.2 \times 20 = 8.4 J$$

Molar mass of ethanol,  $C_2H_5OH = (12 \times 2) + 5 + 16 + 1 = 46 g/mol$

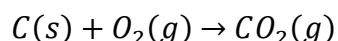
Number of moles of ethanol burnt =  $\frac{0.4}{46} = 0.0086$  moles

0.0086 moles of ethanol liberates 8.4 J of heat

1 mole of ethanol burns to produce  $\frac{1 \times 8.4}{0.0086} = 976.7 J mol^{-1}$  of heat

Therefore the heat of combustion of ethanol =  $976.7 J mol^{-1}$

- 3) Carbon burns in excess oxygen according to the equation below.



$$\Delta H = -393 KJ mol^{-1}$$

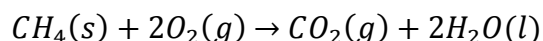
What mass of carbon will produce 750 KJ of energy?

### Solution

393 KJ of heat is produced by 12 g of carbon

750 KJ of heat will be produced by  $\frac{750 \times 12}{393} = 22.90 \text{ g of carbon}$

4) Methane burns according to the equation below.



A litre of methane costs 600F. Calculate the cost of methane required to produce  $1.746 \times 10^3 \text{ J}$  of heat. (1 mole of gas occupies 24 litres at r.t, heat of combustion of methane is -882 KJ mol<sup>-1</sup>)

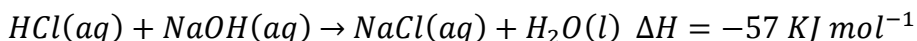
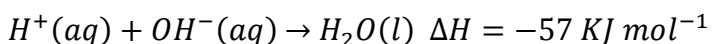
### Solution

## Importance of enthalpy of combustion

- It helps in nutrition when recommending suitable food stuff for individual depending on his or her energy requirements.
- When buying fuel since the different types of fuel differ in their energy values.

## Enthalpy of neutralization

This is the heat liberated when one mole of hydrogen ion completely reacts with one mole of hydroxyl ion to form one mole of water. ie



The heat of neutralization depends on the nature of the acid and the base reacted i.e.

- For strong acids and strong bases, the heat produced is quite high and constant. This is because strong acids and bases are completely ionized giving many free hydrogen ions and hydroxyl ions which readily combine giving a lot of heat.
- For strong acid/weak base or weak acid/strong base or weak acid/weak base, the heat released is small and inconsistent. This is because weak acids and bases are partially ionized giving few hydrogen ions and hydroxyl ions and some heat which should have been released is used to ionize the weak acid and base.

## Experiment to determine the heat of neutralization of hydrochloric acid by sodium hydroxide

### Diagram

### Procedure

- A known volume,  $V_1 \text{ cm}^3$  of hydrochloric acid of known molarity,  $m_1$  is measure into a vacuum flask or a calorimeter and its initial temperature,  $T_1^\circ\text{C}$  noted.
- Similarly, a known volume,  $V_2 \text{ cm}^3$  of sodium hydroxide of known molarity,  $m_2$  is measured and its initial temperature,  $T_2^\circ\text{C}$  recorded.

- Sodium hydroxide is carefully added to the acid in the vacuum flask with constant stirring and the highest temperature,  $T^\circ\text{C}$  of the mixture is recorded.

### Calculation

$$\text{Temperature change} = T_3 - \left(\frac{T_1 + T_2}{2}\right)^\circ\text{C}$$

Mass of the mixture =  $(V_1 + V_2)\text{g}$  assuming the density of the solution is  $1\text{g/cm}^3$

Specific heat capacity of the solution =  $C$

Heat given out = Heat gained by the mixture

$$= \text{mass of mixture} \times \text{shc of solution} \times \text{temperaure rise}$$

$$= (V_1 + V_2) \times C \times \left[T_3 - \left(\frac{T_1 + T_2}{2}\right)\right] \text{ J}$$

$$\text{Number of moles of acid used} = \frac{V_1 \times m_1}{1000} \text{ moles}$$

$$\text{From } H^+(aq) + OH^-(aq) \rightarrow H_2O(l) \text{ moles of hydrochloric acid} = \frac{V_1 \times m_1}{1000} \text{ moles}$$

$$\frac{V_1 \times m_1}{1000} \text{ moles of hydrogen ions liberate } (V_1 + V_2) \times C \times \left[T_3 - \left(\frac{T_1 + T_2}{2}\right)\right] \text{ J of heat}$$

$$1 \text{ mole of hydrogen ions would liberate } \frac{(V_1 + V_2) \times C \times \left[T_3 - \left(\frac{T_1 + T_2}{2}\right)\right] \times 1000}{V_1 \times m_1} \text{ J of heat}$$

$\therefore$  The heat of neutralization hydrochloric acid by sodium hydroxide =

$$\frac{(V_1 + V_2) \times C \times \left[T_3 - \left(\frac{T_1 + T_2}{2}\right)\right] \times 1000}{V_1 \times m_1} \text{ J}$$

### Worked examples

- When  $50\text{ cm}^3$  of  $0.5\text{M}$  hydrochloric acid was added to  $50\text{ cm}^3$  of  $0.5\text{ M}$  sodium hydroxide in a calorimeter. There was a temperature rise from  $27.5$  to  $30.8^\circ\text{C}$ . Calculate the heat of neutralization.

#### Solution

Heat given out = Heat gained by the mixture

$$= \text{mass of mixture} \times \text{shc of solution} \times \text{temperaure rise}$$

$$= (50 + 50) \times 4.2 \times [30.8 - 27.5] = 1386 \text{ J}$$

$$\text{Number of moles of acid used} = \frac{50 \times 0.5}{1000} = 0.025 \text{ moles}$$

$0.025 \text{ moles}$  of hydrogen ions liberate  $1386 \text{ J}$

$$1 \text{ mole of hydrogen ions would liberate } \frac{1386 \times 1}{0.025} = 55440 \text{ J of heat}$$

$\therefore$  The heat of neutralization hydrochloric acid by sodium hydroxide =

$$-55.44 \text{ KJ mol}^{-1}$$

- $50\text{ cm}^3$  of sodium hydroxide of concentration  $0.4\text{M}$  required  $20\text{ cm}^3$  of  $0.5\text{ M}$  sulphuric acid for complete neutralization. The temperature rise was  $3.4^\circ\text{C}$ . If both solutions in the calorimeter were initially at the same temperature. Find the enthalpy of neutralization of the reaction. (Assume the heat capacity of the calorimeter is negligible and s.h.c of water is  $4.2\text{ Jg}^{-1}\text{mol}^{-1}$ )

#### Solution

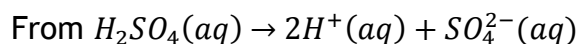
Heat given out = Heat gained by the mixture

$$= \text{mass of mixture} \times \text{shc of solution} \times \text{temperaure rise}$$

$$= (50 + 20) \times 4.2 \times 3.4 = 999.6 \text{ J}$$



$$\text{Number of moles of acid used} = \frac{20 \times 0.4}{1000} = 0.008 \text{ moles}$$



$$\text{Number of moles of hydrogen ions} = 2 \times 0.008 = 0.016 \text{ moles}$$

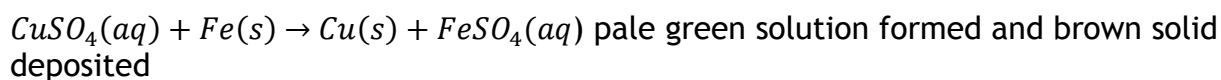
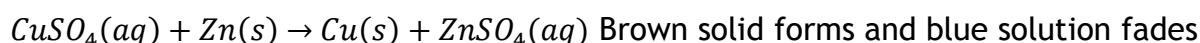
0.016 moles of hydrogen ions liberate 999.6 J

$$1 \text{ mole of hydrogen ions would liberate } \frac{999.6 \times 1}{0.016} = 62475 \text{ J of heat}$$

$$\therefore \text{The heat of neutralization sulphuric acid by sodium hydroxide} = -62.475 \text{ KJ mol}^{-1}$$

## Heat of displacement

This is the heat change which occurs when one mole of a metal ion has been displaced from its solution by another metal higher than it in the electrochemical series. eg



These reactions are normally accompanied by colour changes. In this reaction, a known volume of copper (II) sulphate solution of known molarity is measured into a conical flask and its initial temperature is taken.

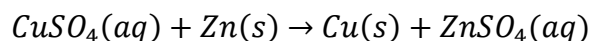
The solution is put in a calorimeter and a known mass of zinc dust is added to the copper (II) sulphate solution with constant stirring until there is a reasonable change in the heat content of the mixture. The highest temperature attained is recorded.

### Diagram

## Worked examples

- 1) 25 cm<sup>3</sup> of copper (II) sulphate solution of molarity 0.5 m was put in a beaker and excess zinc dust (1 g) was added to the solution. The initial temperature of the solution was 25°C and final temperature was 27°C. Find the heat of displacement of copper (II) sulphate.

### Solution



$$\text{Heat change} = MC\Delta t$$

$$= 25 \times 4.2 \times (27 - 25) = 210 \text{ J}$$

$$\text{Number of moles of } CuSO_4(aq) = \frac{25 \times 0.5}{1000} = 0.0125 \text{ moles}$$

0.0125 moles of Cu<sup>2+</sup> displaced gives out 210 J of heat

$$1 \text{ mole of Cu}^{2+} \text{ displaced gives out } \frac{210}{0.0125} = 16800 \text{ J of heat}$$

$$\text{Therefore the heat of displacement} = -16.8 \text{ KJ mol}^{-1}$$

2)

## Heat of solution

This is the heat change which occurs when one mole of a solute is dissolved in a specified amount of solvent to form an infinitely dilute solution.

An infinitely dilute solution is one in which on further dilution does not show any heat change.

**NB//.** Enthalpy of solution can be endothermic or exothermic. If it is exothermic it means that the solute is not very soluble in water and would easily dissolve on heating. If it is endothermic, the solute is said to be very soluble in water e.g. sodium hydroxide, gases such as sulphurdioxide, anhydrous copper (II) sulphate.

### Experiment to determine the heat of solution of sodium chloride

#### Diagram

#### Procedure

- A known mass of water is placed in a calorimeter and its initial temperature noted.
- A known mass of sodium chloride is added to the water with constant stirring.
- The highest or lowest temperature of the mixture attained is recorded.

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#### Worked examples

- 1) 1 g of sodium chloride was added to 1000 cm<sup>3</sup> of water in a calorimeter. The temperature of water changed from 23 to 27°C. Calculate the heat of solution of sodium chloride (shc is 4.2J/g/K and density is 1g/cc)

#### Solution

$$\text{Heat change} = MC\Delta t$$

$$= (100 + 1) \times 4.2 \times (27 - 23) = 1696.8J$$

$$\text{Molar mass of NaCl} = 23 + 35.5 = 58.5 \text{ g}$$

$$\text{Number of moles of NaCl} = \frac{1}{58.5} = 0.017 \text{ moles}$$

$$0.017 \text{ moles of NaCl liberates } 1696.8J$$

$$1 \text{ mole of NaCl will liberate } \frac{1696.8 \times 1}{0.017} = 99811.76KJ \text{ mol}^{-1}$$

$$\text{Therefore the heat of solution} = 99.81176KJ \text{ mol}^{-1}$$

2)

## RATES OF REACTION

### Kinetics

This is a branch of chemistry that deals with the study rates of reaction.

### Rates of reaction

This refers to change in concentration of reactants or products during a chemical reaction. During a chemical reaction, the concentration of reactants decreases while that of the product increases. The concentration of the reactants decreases because they are being used up to form products and the concentration of products increase until a maximum point when it remains constant implying that all reactants have been used up. This can be illustrated as below.

#### Diagram

The study of rates of reaction helps us to;

- Design new drugs which work better
- Design new ways of making other chemicals efficiently.

### Why reactions occur

This can be explained in terms of collision theory of chemical reaction which states that a reaction occurs if the particles involved in the reaction collide together and if they have appropriate energy or right energy. This energy is called activation energy.

Activation energy is the minimum amount of energy required by the reactant molecules to change into products. This energy exists in reactant molecules in form of potential energy. Reactions may not take place if;

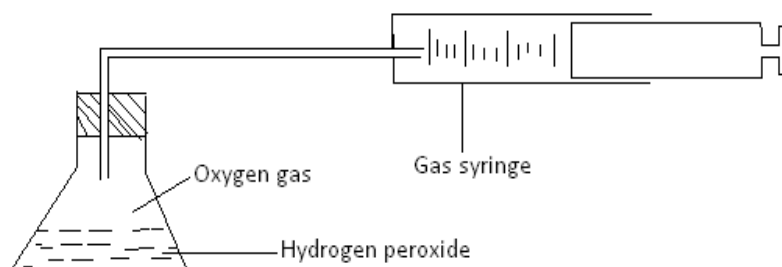
- Reactant molecules have energy less than the activation energy.
- Reactants do not collide in the right geometry.

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### Measuring rate of reaction

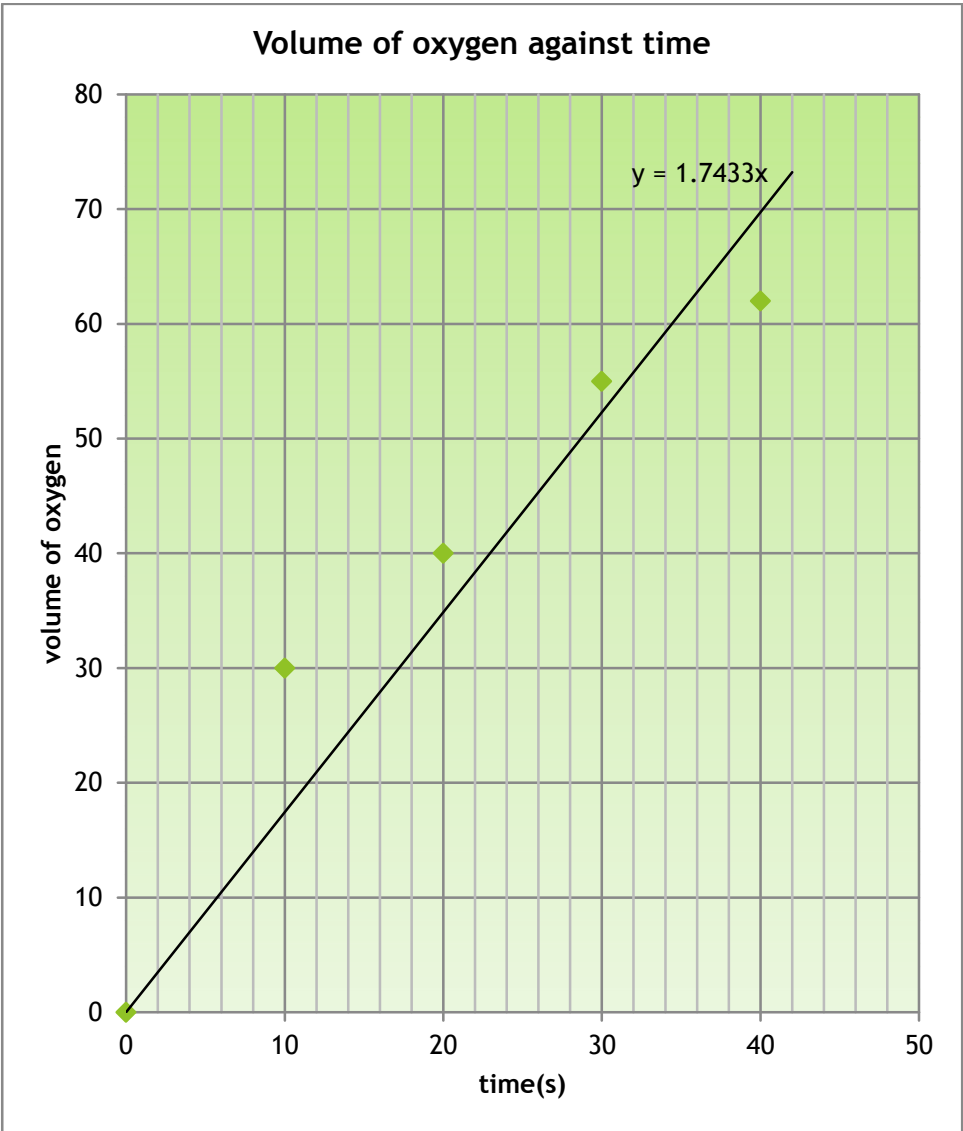
This can be done in the following ways;

- 1) **Measuring the volume of gas given out per unit time.** This method is for reactions where a gas is being produced. eg consider the decomposition of hydrogen peroxide. In this method, the volume of oxygen gas evolved is measured after intervals of time and a graph of volume of oxygen gas produced is plotted against time.

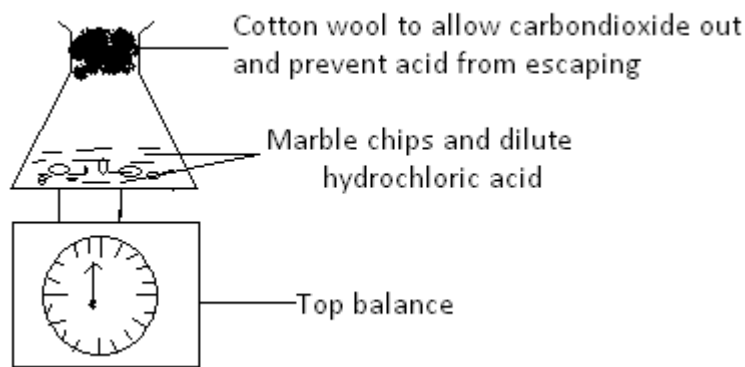
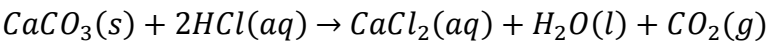


#### Specimen results

Time (s)	0	10	20	30	40
Volume of oxygen (cm <sup>3</sup> )	0	30	40	55	62



2) **Measuring change in mass.** In this method, a weighing balance is used and a change in mass of the reactants is recorded pre unit time. Consider the reaction between calcium carbonate and dilute hydrochloric acid.



**Specimen results**

Mass of calcium carbonate (g)	50	30	35	20
Time (s)	100	150	170	200

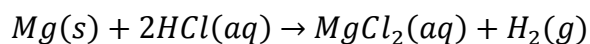
A graph of mass of calcium carbonate against time is plotted. The gradient or slope of the curve at any instant gives the rate of reaction at that point.

### Factors affecting the rate of a chemical reaction

- Concentration of the reactants
- Temperature
- Surface area (size of particles)
- Catalyst
- Pressure for gaseous reactants

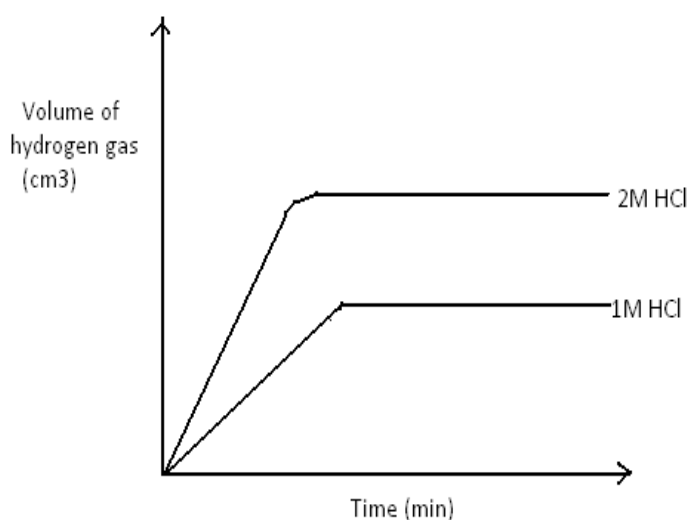
The above factors can either increase or decrease the rate of reaction depending on the type of the reaction.

- 1) **Concentration of the reactants.** Generally, increase in the concentration of the reactants increases the rate of chemical reaction and decrease in concentration decreases the rate of reaction. According to the collision theory, reactants must first collide before products are formed thus the higher the concentration the higher the frequency of collision per unit time and the faster the products are formed. Consequently, a higher rate of reaction and the reverse is true. Eg consider the reaction between dilute hydrochloric acid and magnesium as shown by the equation;



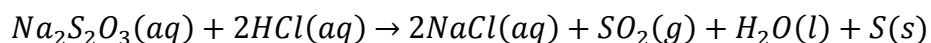
#### Procedure

- The concentration of hydrochloric acid is varied and that of magnesium kept constant.
- In the first experiment, 2M hydrochloric acid is used and the volumes of hydrogen gas produced per unit time are measured or noted.
- In the second experiment, 1M hydrochloric acid is used and the volumes of hydrogen gas produced per unit time are recorded.
- A graph of volume against time is plotted for both 2M hydrochloric acid and 1M hydrochloric acid as shown below.



Effect of concentration can also be demonstrated using the reaction between dilute hydrochloric acid and sodium thiosulphate solution.

Sodium thiosulphate normally reacts with dilute hydrochloric acid to produce sulphur, sulphurdioxide and sodium chloride. Sulphur produced during the reaction makes the solution appears cloudy.

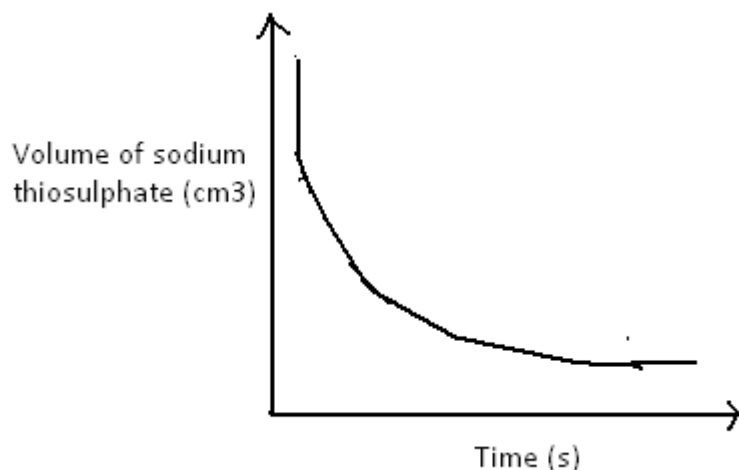


#### Procedure

- Measure 50 cm<sup>3</sup> of sodium thiosulphate solution and pour it into a beaker placed on a white paper marked with a cross.
- Add 5 cm<sup>3</sup> of 2M hydrochloric acid to the solution and at the same time start a stop clock.
- Note the time taken for the cross to disappear
- Repeat the experiment by varying the volume and concentration of sodium thiosulphate i.e. 40, 30. And 10 cm<sup>3</sup>.
- Record the results in the table below.

Volume of Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub> (cm <sup>3</sup> )	60	50	40	30	20	10
Volume of H <sub>2</sub> O(l) (cm <sup>3</sup> )	0	10	20	30	40	50
Volume of HCl (cm <sup>3</sup> )	5	5	5	5	5	5
Time for cross to disappear (s)	27	18	31	35	70	147

- State why the cross disappear in this experiment.
- Write an ionic equation for the reaction.
- Plot a graph of;
  - volume of sodium thiosulphate against time.
  - $\frac{1}{t}$  against volume of sodium thiosulphate.
- Deduce from the graph how the rate of reaction varies with concentration of sodium thiosulphate.
- Explain the relationship between rate of reaction and  $\frac{1}{t}$ .



Rate of reaction is inversely proportional to time.

$$R \propto \frac{1}{t}$$

Therefore a graph of  $\frac{1}{t}$  against concentration gives a straight line which passes through the origin implying that the rate of reaction is directly proportional to  $\frac{1}{t}$ .

Volume of Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub> (cm <sup>3</sup> )	50	40	30	20	10
Volume of H <sub>2</sub> O(l) (cm <sup>3</sup> )	0	10	20	30	40

Volume of HCl (cm <sup>3</sup> )	5	5	5	5	5
Time for cross to disappear (s)	18	26.5	30	46	124

- 2) **Temperature.** Generally, increase in temperature increases the rate of chemical reaction and decrease in temperature decreases the rate of chemical reaction. As temperature increase more reactant molecules acquire activation energy and this result into increase in the number of collisions per unit. It can be demonstrated by considering the reaction between dilute hydrochloric acid and sodium thiosulphate.

#### Procedure

- Measure 50 cm<sup>3</sup> of sodium thiosulphate solution and pour it into a beaker placed on a white paper marked with a cross.
- Add 5 cm<sup>3</sup> of 2M hydrochloric acid to the solution and at the same time start a stop clock.
- Note the time taken for the cross to disappear
- Repeat the experiment by varying the temperature of sodium thiosulphate i.e. 40, 30. And 10 cm<sup>3</sup>.

- 3) **Surface area.** This refers to the total number of particles or molecules available for reaction. Sometimes it is referred to as particle size. Generally, the greater the surface area of reactants, the higher the rate of reaction since greater number of particles exposed collides more frequently with each other and the smaller the surface area, the lower the rate of reaction because few reactant molecules are exposed to collide with each other. For solid reactants surface area can be increases by;

- Grinding or pounding solid lamps to powder form (many small particles).
- Dissolving particles if it is a soluble compound.

The effect of surface area on rate of reaction can be investigated by considering the reaction between calcium carbonate and hydrochloric acid.

In this experiment, dilute hydrochloric acid is added to powdered calcium carbonate and the volumes of carbondioxide gas evolved measured at a give time interval.

The experiment is repeated using the same mass of lamp calcium carbonate and volume of hydrochloric acid used. The volume of carbondioxide gas is also measured at a given time interval as above.

Finally, a graph of volume of carbondioxide gas evolved from both experiments is plotted against time on the same axes.

#### Sketch

From the graph, the rate of reaction at the same instant is high when powdered calcium carbonate is low compared to when lamp size calcium carbonate is used.

#### Question

When a certain volume of 0.1M hydrochloric acid was reacted with excess iron fillings are room temperature, 120 cm<sup>3</sup> of a gas was produced.

- (a) Draw a well labeled diagram to show how the rate of reaction was determined.

- (b) Write an equation for the reaction.
- (c) Calculate the;
- volume of 0.1M hydrochloric acid required to produce 120 cm<sup>3</sup> of a gas.
  - mass of iron fillings that reacted.
- (d) Draw a sketch graph of volume of the gas against time.
- (e) State how the rate of reaction would change if the reaction was carried out at a temperature above room temperature.
- 4) **Catalyst.** Generally, rate of reaction increases in the presence of a catalyst.

Reactions	Catalyst
$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$	MnO <sub>2</sub> , PbO and CuO
$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$	CuSO <sub>4</sub>
$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$	V <sub>2</sub> O <sub>5</sub>
$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$	Fe

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### Characteristics of catalysts

- They remain chemically unchanged after reaction.
- They are used in small quantities.
- Efficiency of catalysts can be improved by other substances e.g. growth promoters.

Promoters increase the efficiency of a catalyst. There are two types of catalysis. It refers to the process of increasing rate of chemical reaction using catalyst.

**Homogeneous catalyst.** This where the catalyst and the reactants are in the same physical state e.g. hydrolysis of sucrose.

**Heterogeneous catalyst.** Catalysts and reactants are in different physical states.





