

PREFACE:

This book is intended for use by students of S.1 – S.4 preparing for Uganda Certificate of Education (UCE).

The book covers the entire syllabus recommended by National Curriculum Development Centre (NCDC) in conjunction with Ministry of Education and Sports.

A lot of theoretical treatment of the mole concept especially calculations have been treated well; mathematical concepts required in this topic have been simplified.

In Electrolysis, key terms involved and explanations have been given utmost treatment to enhance learning and understanding.

This book has been carefully planned to enable students understand the subject and be able to make conclusions. At the end of every topic, enough exercises have been included to keep students busy and testing the understanding of the topic.

With this book's *'essentials of Chemistry'* hardworking and serious students are destined to excel. It is hoped that the book will be useful to both teachers and students to enhance the mastery of the subject.

Use this book always and success will automatically be yours.

ACKNOWLEDGEMENT

The author wishes to record his gratitude to the following people for the assistance, encouragement, cooperation and wise suggestions they rendered.

I specifically want to thank the deputy Head teacher Kyebambe Girls SS, Mr. Balinda Richardson for his mentorship, guidance, support and encouragement.

I wish to extend my sincere gratitude to my wife Win Enyangu for typesetting the book, the encouragement and the support.

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DEDICATION:

This book is dedicated to my two children – Apio Tracy and Acen Tricia and all with beautiful brains.

I thank all of you for being supportive, kind and generous towards the production of this book.
May God bless you all.

Enyangu Charles
0782/ 0704/ 0713- 451196

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CHEMISTRY

Chemistry is the study of matter, its composition, properties and behavioural changes that form the environment.

Chemistry is the basis of chemical industry and many chemists work in industries involved in making useful products from raw materials such as soap, detergents, cosmetics, textiles etc.

IMPORTANCE OF LEARNING CHEMISTRY

- The chemical knowledge, concepts and skills enhance improvement in the quality of life of the people.
- To help the learner to develop interest in and care for the surrounding environment.
- It enables people to appreciate the numerous chemical products and how to use them properly.
- It enables people to make progress in related careers e.g. Doctors, teachers, Nurses etc.
- To enable people to discover knowledge using a practical approach and to apply the discovered knowledge in everyday life. The world today relies on chemistry as one key component in every aspect of life.
- Enable man to produce enormous energies which are important in society e.g. Petrol.

LABORATORY RULES AND REGULATIONS

The chemistry laboratory is stocked with different types of chemicals and apparatus. Most of these chemicals in the chemistry laboratory are harmful to the human body and the apparatus breakable if not handled carefully. It is important for beginners in form one to be acquainted with laboratory rules and safety measures before they can confidently use the chemistry laboratory. These rules and regulations include;

1. Never run or play around in a laboratory.
2. Never eat or drink in the laboratory
3. When heating, make sure that the test tube mouth is not facing any one.
4. Ensure that any loose clothing is carefully tucked in
5. Do not enter the laboratory unless the teacher gives you permission.
6. Report all accidents and breakages to the teacher
7. Clean and clear up all apparatus before leaving the laboratory.
8. Do not throw solids in the sink

9. Follow the instructions carefully before you do any experiment.
10. Put all bags, jackets and stools out of your way before you do any experiment.
11. Do not stand on a stool while carrying out an experiment.

LABORATORY APPARATUS

These are pieces of equipment used in the laboratory for experiments.

Common Apparatus in a chemistry laboratory

GLASS WARE

These are pieces of equipment made of glass. Most of the apparatus in the chemistry laboratory are made of glassware for two major reasons;

1. Glass is transparent and enables the experimenter to observe easily the reactions occurring.
2. Glass does not react with most common laboratory chemicals.

1. Test tube

This is used for mixing and heating small amounts of liquids or solids.

2. Boiling tube

A boiling tube is wider than a test tube and it's longer. It's used for heating substances strongly or boiling liquids.

3. Measuring cylinder

This is used for measuring approximate volumes of liquid substances.

4. Burette

This is used for very accurate measurements of volumes. It is used in titration experiments.

5. Pipette

This is used to measure fixed volumes of liquids accurately.

6. Volumetric flask

It's used to measure precise volumes of liquids.

7. Conical flask

This is used for mixing liquids as on shaking the liquid does not splash out. It also holds liquids or solutions for use.

.

8. Glass rod

For stirring to dissolve a solid material in a solvent. It is also used to detect formation of a crystal during evaporation of aqueous solutions.

9. Flat bottomed flask.

This is used for mixing of a solid with a liquid.

10. Round bottomed flask

This is used for heating liquids for a long period. It's used in distillation experiments.

11. Beaker.

Beakers are used for heating, holding and measuring liquids.

12. Thistle funnel

This is used for putting liquids into reaction vessels

13. Separating funnel

This is used for separating immiscible liquids e.g water and Kerosene.

OTHER COMMON APPARATUS (NON GLASSWARE)

1. Tripod stand.

This is used for supporting apparatus while it is heated.

2. Test tube rack

For holding test tubes during an experiment especially qualitative analysis.

3. Wire gauze

For uniform distribution of heat while heating.

4. Spatula

For scooping solid chemicals from their containers.

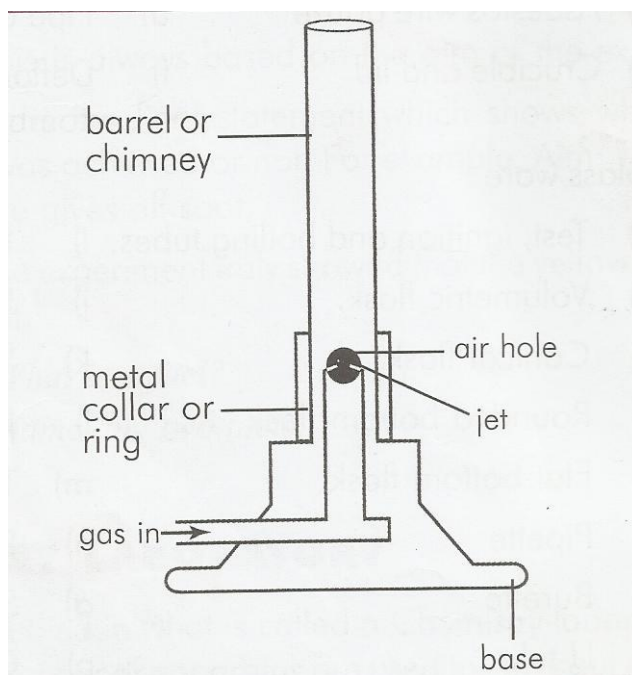
5. Filter paper

For filtration experiments to separate liquid of solution from undissolved solid.

THE HEATING APPARATUS USED IN CHEMISTRY LABORATORY

THE BUNSEN BURNER

This is a simple piece of apparatus used in the laboratory for heating substances and it uses gas as a source of fuel.



The Bunsen burner mixes the gas with air in the correct proportions to give a very hot flame. The gas enters the burner through the jet at the bottom and burns at the top of the chimney. The collar has a hole on each side of it and air can pass through the hole and mix with the gas as it goes up the tube. The amount of air entering can be adjusted by turning the collar. The flames of a candle and burning oil are smoky and luminous (yellow) and are not used for heating in the laboratory for the following reasons:

- (i) They are not hot enough.
- (ii) They cover apparatus with black soot (carbon).

FUNCTIONS OF PARTS OF A BUNSEN BURNER

1. Barrel (Chimney)

It allows gas to pass through to be used to create a flame. It transmits a mixture of gas and air up to the top of the chimney where it is lit.

2. Collar (Metal ring)

When adjusted it increases or decreases the size of air hole by making it smaller or bigger.

3. Air hole

It's an opening that allows air to enter the Bunsen burner so that when it mixes with the gas from the jet, it burns when lit.

4. Jet.

This is where the gas enters the burner and it provides the necessary pressure to push the gas through the Chimney

5. Rubber tubing

It joins the Bunsen burner to the gas cylinder and transmits the gas from the gas cylinder to the jet of the Bunsen burner.

6. Base

This is to ensure that the Bunsen burner is stable. It is where the burner sits and gets its support.

HOW TO LIGHT A BUNSEN BURNER

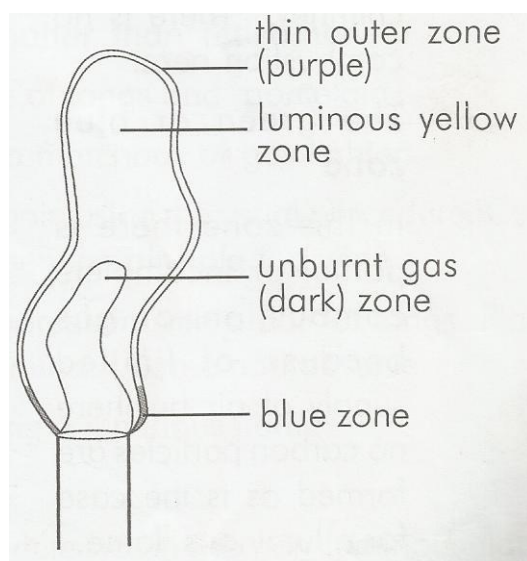
- (i). Connect the Bunsen burner to the gas cylinder
- (ii). Turn the collar until the air holes are closed.
- (iii). Turn on the gas fully.
- (iv). Light the gas at the chimney to produce a flame.
- (v). Open the air holes slowly until the flame is non luminous.

FLAMES

A flame is burning gases that give out heat and light.

1. Luminous flame (Air-holes closed)

This is the flame produced when the air holes are closed. The flame is large and bright. It's luminous because it gives out light.

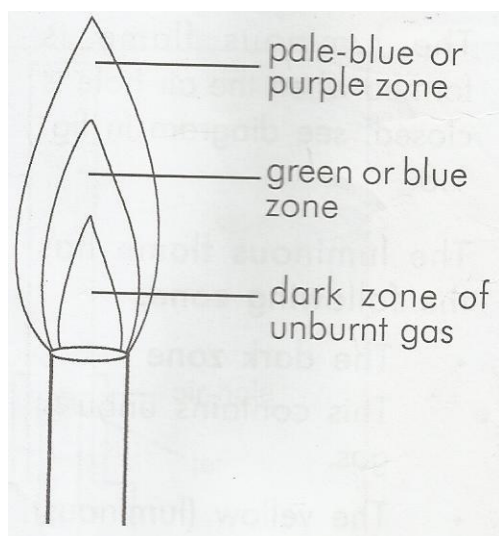


A luminous flame has four parts or zones with different colours.

- (a) A dark inner zone of cool, un burnt gas.
- (b) A luminous yellow zone. The gas burns in this zone but not completely as there is not enough air.
- (c) A thin outer zone in which the gas burns completely as there is plenty of air. This zone does not give out light, it's difficult to see this zone.
- (d) A blue zone at the bottom of the flame. The part receives plenty of air because of rising convection currents. Burning is more complete than in the yellow zone.

2. Non-Luminous flame (Air-holes open)

This is a flame produced when the air hole is open. Air enters the tube and mixes with the gas which burns quickly and completely. The flame becomes smaller and hotter. It gives out only a little light because it contains no carbon. The flame has three zones.



Uses of non luminous flame.

- It's used in gas cookers.
 - It's also used in a gas fires
- (a) An inner zone of cool unburnt gas.
- (b) A green or blue middle zone. Some gas burns in this zone, but not all because there is not enough air.
- (c) A pale blue or purple zone. In this zone the burning of the gas is more complete.

“Burning Back”

This is a phenomenon where the flame moves down the chimney and is seen to burn at the jet. It occurs when the air holes are open and the gas is turned down so that the rate at which the gas is burnt exceeds the rate at which it is supplied from the jet.

Control of Burning Back

When burning back occurs, turn off the gas, close the air hole and light it again using the normal procedure.

Experiment to find out the hottest part of the Bunsen flame

- Get a piece of paper and hold it horizontally.
- Pass the paper quickly through a non luminous flame around two places .
- Brown patches form on the paper. The brown patch is formed at the hottest part of the Bunsen burner located just above the blue or green zone.

Differences between luminous and non luminous flame

Luminous	Non luminous
<ul style="list-style-type: none">• Has four zones• Produces soot• Burns silently• Produces a lot of light• Yellow and unsteady• Large and not hot enough	<ul style="list-style-type: none">• Has three zones• Does not produce soot• Burns with noise• Produces little light• Blue and steady• Small and hot enough

LABORATORY METHODS OF PREPARING GASES

The following steps are involved when preparing a gas:

1. Production

Reagents are mixed in a reaction vessel and the gas is produced.

2. Purification

The gas is passed through another reagent normally in a wash bottle to remove the impurities.

3. Drying

In most cases dried sample of gases may be required. Common drying agents include:

Concentrated sulphuric acid-the gas is passed through concentrated sulphuric acid in a wash bottle. The inlet delivery tube must be below the acid and the outlet delivery tube above the acid.

Concentrated sulphuric acid is not used to dry ammonia gas because it reacts with ammonia.

Anhydrous calcium chloride- the gas is passed through anhydrous calcium chloride in a U-tube.

Calcium oxide is used to dry ammonia gas. Dried gases must not be collected over water.

Other drying agents include anhydrous magnesium sulphate, anhydrous sodium sulphate, anhydrous calcium sulphate. Phosphorus pentoxide is another drying agent used to dry ether.

4. Collection

The method of collection depends on the density of the gas in comparison to the density of air and also the solubility of the gas in water.

(a) Over water

The gas jar is filled with water in the trough, inverted and the mouth of the jar placed over the beehive shelf. The gas displaces water from the gas jar when it is full. This method is suitable for gases that are insoluble in water eg nitrogen and oxygen.

(b) Upward delivery (Downward displacement of air)

Gases that are less dense than air are collected by upward delivery eg hydrogen and ammonia.

(c) Downward delivery (upward displacement of air)

Gases that are denser than air are collected by downward delivery eg carbon dioxide, sulphur dioxide, chlorine, hydrogen chloride.

(d) Freezing method

The gas is liquefied by cooling in ice-cold freezing mixture in a U-tube e.g. nitrogen dioxide.

Exercise

1. (a) What is chemistry?
(b) State three importance of studying chemistry
(c) Give any three laboratory rules.
(d) What first aid would you give to:
(i) A person who has swallowed poison
(ii) A person who has ingested acid
2. (a) What is a flame?
(b) Name the apparatus used for heating in the chemistry laboratory.
(c) Draw and label the apparatus used for heating in the chemistry laboratory.
(d) Which parts of the apparatus named in (c) above carry out the following functions:
(i) controls the amount of air entering
(ii) provides the necessary pressure for pushing the gas.
(i) Increases or decreases the size of the air hole.
3. (a) Name the two types of Bunsen flames.
(b) How are the flames named above produced.
(c) State four differences between the flames named in (b) above.
4. (a) What is burning back?

- (b) What causes burning back?
- (c) State what can be done when burning back occurs.
5. Draw a labelled diagram of
- (i) luminous flame
- (ii) non luminous flame
- (b) Indicate with letter R, the hottest part of a non luminous flame.
- (c) What would happen to the flame of a non luminous flame if the gas supply was slowly turned down when the air holes are fully open?
6. (a) Describe an experiment to find out the hottest part of a non luminous flame.
- (b) Outline the steps for lighting a Bunsen burner.
- (c) Why is it not advisable to use a luminous flame for heating chemicals in the laboratory?
7. (a) Why is it important to have laboratory rules and regulations?
- (b) State any three laboratory rules and regulations that must be observed;
- (i) before performing an experiment
- (ii) During an experiment
- (iii) After an experiment
8. (a) Complete the following table

	Diagram	Name	Use
(i)			Removing small samples of solid from container bottles.
(ii)			Spread heat when heating in a beaker.
(iii)		pipette	
(iv)		Test tube rack	
(v)			Holding the test tube
(vi)		Burette	

MATTER

Matter is anything that occupies space and has weight (mass). Every substance is made up of matter, which occurs in three major states.

States of matter

Matter exists in three states. These are solids, liquids and gases or vapour.

Characteristics of states of matter

Solid:

- Solid particles have definite size (volume) and shape. Force is required to change this shape
- They cannot be compressed and do not flow. Examples of solids include; cement, sand, wood, wool.

Liquids

- Liquids have a definite volume (size) but no definite shape. It takes up the shape of containing vessel. Liquids flow along a slope from a higher level to a low level. They cannot be compressed. Examples of liquids include; water, alcohol, nitric acid, mercury.

Gas (Vapour)

- A gas has no definite volume and shape, it fills any volume in which it is placed.
- Gases are easily compressed into a small space by pressure e.g. air into a bicycle tyre.
- A gas can flow in any direction. Examples of gases include; oxygen, nitrogen, carbon dioxide.

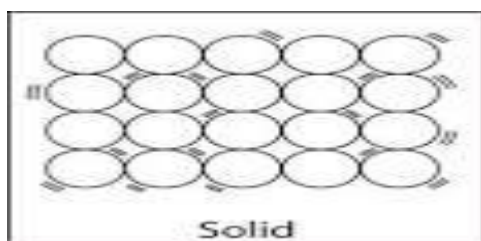
KINETIC THEORY OF MATTER

Kinetic theory states that all particles of matter have energy. This energy is kinetic energy i.e. Energy of movement.

Solids.

In a solid the particles are closely packed and held together by strong force of attraction (cohesion). Cohesion is force of attraction between particles of the same substance. The particles are fixed but can rotate and vibrate about their fixed positions.

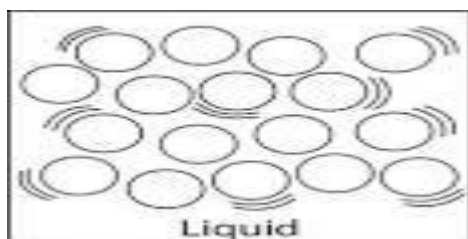
Heat increases the vibrations until melting point when the solid becomes a liquid.



The solid can change directly into gas when heated without passing through the liquid state by a process called sublimation.

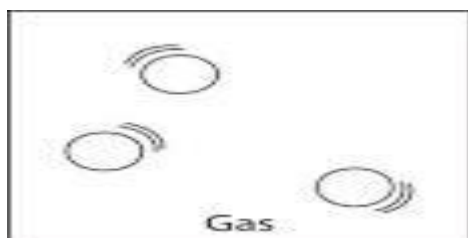
Liquids.

The particles in a liquid can move about within the liquid but are still held together by intermolecular forces. Heat makes the particles to move faster until boiling point when particles break free and the liquid turns into vapour (gas).



Gas.

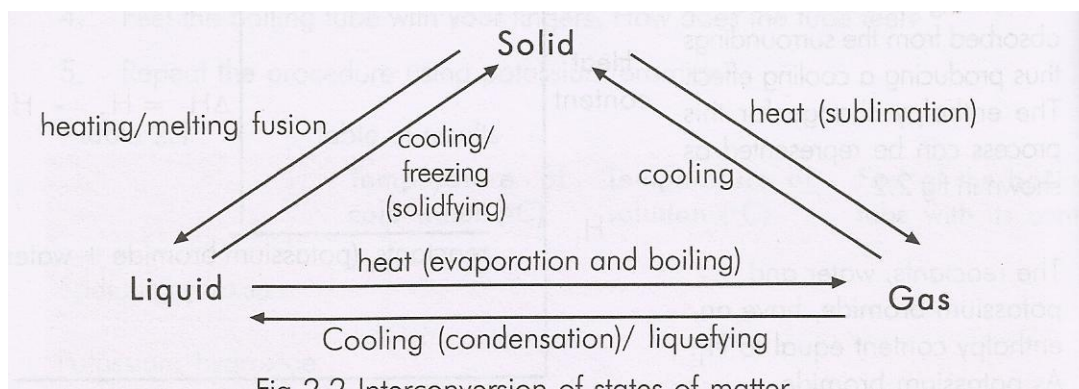
In gas the particles are widely spaced apart. The forces of attraction between them are very weak. They are free to move in all directions.



The gas can change to solid by sublimation. When the gas is cooled, it condenses into a liquid.

CHANGE OF STATE

The state of a substance depends on the temperature. A change in temperature can change the state of a substance.



Condensation is the change of state from gas to liquid due to cooling.

When the temperature of a liquid is reduced, the kinetic energy of the molecules decreases, then the liquid changes to solid.

When a solid is heated, the molecules acquire more kinetic energy to overcome the attractive forces hence the solid changes to liquid in a process called melting.

Sublimation is the change of state from solid to gas without passing through the liquid state.

Examples of substances that sublime

1. Ammonium chloride
2. Iodine
3. Iron (III) Chloride
4. Aluminium Chloride.

Evaporation(boiling)/ vaporization.

Vapour is the gaseous state of the substance which can be compressed to a liquid.

Vaporisation is the change of state from liquid to gas due to heating.

EFFECT OF HEAT ON SUBSTANCES

1. Magnesium:

The bright shiny metal burns with a bright flame forming a white powder.

2. Sulphur

Put sulphur powder in a test tube. Heat the test tube gently.

Observation

Sulphur melts and liquid sulphur changes back to sulphur again on cooling.

3. Iodine

Place solid iodine in a dry test tube and heat gently.

Observation:

The iodine changes directly to a violet vapour which condenses on cooling to form black shining crystals of iodine. This change is sublimation.

4. Wax: The wax, which is a mixture, melts gradually when it is warmed slowly. When very hot wax forms white fumes.

5. Sand: This substance does not change on heating, but any water in it may form water vapour.

6. Copper sulphate crystals: The blue crystals contain water, which comes off when the crystals are heated and white copper sulphate (without water) remains. The white solid reacts with cold water to form a blue substance again.
7. Potassium manganate: The purple solid decomposes and forms oxygen, which relights a glowing splint.
8. Zinc oxide: The cold oxide is white. It becomes yellow when hot, and becomes white again when cool. No change occurs and the substance is still zinc oxide.
9. Ammonium chloride: The white solid changes directly into white fumes which on cooling change back to white solid. No liquid forms.
10. Copper: Copper foil or wire is very thin and quickly changes to black copper oxide and makes the flame bluish-green.

PHYSICAL AND CHEMICAL CHANGES

When sand, zinc oxide, iodine, wax ammonium chloride are heated, no new substances are formed.

A physical change is a change in which no new substance is formed. The products have the same chemical properties as the reactants.

Chemical changes

When copper reacts with air to form copper oxide, magnesium burns in air to form magnesium oxide, copper(II) sulphate crystals change to white copper sulphate, potassium manganate giving off oxygen when heated. New substances are formed in all the above changes, which are chemical (permanent).

A chemical change is a change in which a new substance is formed.

The products have chemical properties different from those of the reactants.

Examples of Physical change

1. Melting of wax
2. Heating of sulphur
3. Boiling of liquids
4. Making of solutions

Examples of chemical Change

1. Rusting of iron
2. Combustion of substances in air
3. Precipitation of a solid from solution.

Differences between physical and chemical change.

Physical change	Chemical change
❖ No new substances are formed	❖ New substances are formed
❖ Usually the change is easily reversible	❖ Usually the change is irreversible
❖ No energy (heat and light) is given out or absorbed	❖ Usually energy (heat) is given out or absorbed.
❖ The mass of the substance does not alter.	❖ The mass of the substance does alter

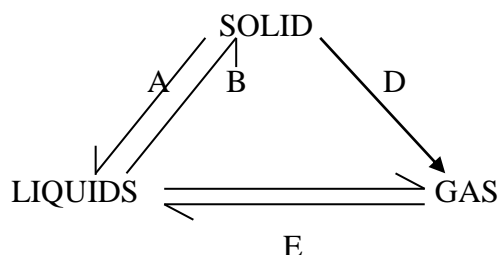
MISE - Mass, Irreversible, Substance, Energy.

Exercise

1. (a) Define the following terms:

- (i) Melting
- (ii) Condensation
- (iii) Vaporisation

(b) The figure below shows interconversions of states of matter.



- (i) Name the processes labelled A,B,E and D
 - (ii) Name three substances that undergo process D
 - (iii) State the condition for process E to occur
2. (a) What is matter?
- (b) Name the three states in which matter exists.
- (c) State two differences between
- (i) Solids and liquids
 - (ii) Solids and gases
 - (iii) Liquids and gases.
3. (a) Giving two examples in each case, explain the following terms:
- (i) A physical change
 - (ii) A chemical change
- (b) State four differences between a physical change and a chemical change.
- (c) State whether the following changes are physical or chemical.
- (i) Heating sand
 - (ii) Heating zinc oxide
 - (iii) Rusting of iron
 - (iv) Precipitation of solids from a solution.
4. (a) Name the drying agents for the following substances:
- (i) hydrogen
 - (ii) ether
 - (iii) Carbon dioxide.

(iv) ammonia

(b) Give two examples in each case of gases collected by the following methods:

(i) Upward displacement of air

(ii) Downward displacement of air

5. Classify the following changes into physical or chemical and give a reason for your choice.

(i) Burning of fire wood

(ii) melting sulphur

(iii) Dissolving sugar in water

(iv) Respiration

(v) Magnetization of iron

(vi) Chewing the PK

6. Explain the following observations:

(a) When white ammonium chloride is heated, it changes to white fumes which on cooling turn to white solid.

(b) When lead (II) oxide is heated, the colour changes from red-brown to yellow on cooling.

(c) When wax is heated, it melts but on cooling it changes back to wax.

7.(a) State the Kinetic theory of matter.

(b) In which state of matter;

(i) Are particles moving in all directions?

(ii) Do particles vibrate in fixed position?

(iii) Do particles fill any container in which they are placed?

(iv) Do particles have no definite shape and volume?

(c) Complete the table by showing how the three states of matter differ according to the properties given on the table.

Property	Solid	Liquid	Gas
Compressibility			
Arrangement of atoms			
Forces of attraction			

(d) How does pressure affect the boiling point of a substance?

(f) Using dots in each case, draw diagrams to show the arrangement of particles in each of the three states of matter.

State of matter	Diagram

8. During an experiment, a beam of light was passed through a dark room and dust particles were seen moving in all directions. State why dust particles moved in all directions.

(b) The temperature of the dark room was increased.

(i) State what was observed

(ii) Give a reason for your answer.

ELEMENTS, COMPOUNDS AND MIXTURES

Elements:

An element is a substance which cannot be split up into two or more simpler substances by chemical means.

All elements are made up of tiny particles called atoms.

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

Chemical symbols

The symbol of an element is one or two letters which mean one atom of the element.

A symbol is usually the first letter of the English or Latin name of the element.

Elements whose symbols are derived from Latin names include:

Examples

Element	Latin Name	Symbol
1. Sodium	Natrium	Na
2. Potassium	Kalium	Ka
3. Lead	Plumbum	Pb
4. Copper	Cuprum	Cu
5. Mercury	Hydragyrum	Hg
6. Silver	Argentum	Ag
7. Iron	Ferrum	Fe
8. Gold	Aurum	Au

Symbols for first 20 elements

Element	Symbol
Hydrogen	H
Helium	He
Lithium	Li
Beryllium	Be
Boron	B
Carbon	C
Nitrogen	N
Oxygen	O
Fluorine	F
Neon	Ne

Sodium	Na
Magnesium	Mg
Aluminium	Al
Silicon	Si
Phosphorous	P
Sulphur	S
Chlorine	Cl
Argon	Ar
Potassium	K
Calcium	Ca

Other common elements and their symbols

<i>Element</i>	<i>Symbol</i>
Copper	Cu
Iron	Fe
Silver	Ag
Lead	Pb
Mercury	Hg
Zinc	Zn
Barium	Ba
Gold	Au

METALS AND NON METALS

Most elements fall into two groups i.e. metals or non metals. A metal is an element which forms positive ions by losing electrons except hydrogen. A non metal is an element which forms negative ions by gaining electrons.

Examples of common metals include: Aluminium, Zinc, Copper, Iron, Lead and Magnesium.

Examples of common non metals include: sulphur, carbon, Phosphorus, Oxygen, Hydrogen and Nitrogen.

Physical properties of metals and non metals

Metals	Non-metals
<ul style="list-style-type: none"> • Lustrous, can be polished • Strong and tough, high tensile strength • Malleable i.e. can be drawn into wire • Ductile • High density • Good conductors of heat and electricity • Sonorous i.e. make sound when hit 	<ul style="list-style-type: none"> • Not lustrous, cannot be polished • Low tensile strength • Not malleable • Not ductile • Low density • Poor conductors • Not sonorous.

COMPOUNDS

A compound is a substance which consists of two or more elements chemically combined together. A compound is formed when one element combines chemically with another element of different type.

Examples of compounds

<i>Compound</i>	<i>Elements in the compound.</i>
1. Water	Hydrogen, and oxygen.
2. Common Salt (Sodium chloride)	Sodium and chlorine.
3. Baking powder (NaHCO_3)	Sodium, hydrogen, carbon and oxygen.
4. Carbon dioxide	carbon, and oxygen
5. Calcium carbonate	Calcium,carbon and oxygen
6. Calcium oxide(lime)	Calcium and oxygen

Chemical names of compounds

Many compounds have chemical names and common names. Chemical names are most commonly used because they are descriptive and tell the elements present in a compound. If the name of a compound ends with **-ide**, it contains usually only two elements eg chlorides such as calcium chloride containing elements calcium and chlorine, and sulphides such as zinc sulphide containing elements zinc and sulphur except hydroxides which end with **-ide** but contains three elements such as sodium hydroxide containing the elements sodium, oxygen and hydrogen.

If the chemical name of the compound ends with **-ate**, the compound contains at least three elements such as calcium carbonate containing elements calcium, carbon and oxygen.

If the chemical name of the compound ends with **-ite**, the compound contains at least three elements such as sodium sulphite containing the elements sodium, sulphur and oxygen, potassium nitrite containing the elements potassium, nitrogen and oxygen.

MIXTURES

A mixture is a substance which consists of two or more elements or compounds not chemically combined together.

Types of mixtures

1. Solid-solid mixtures e.g brass, bronze, sand and salt.
2. Liquid-liquid mixtures e.g water and ethanol.
3. Gas- gas mixture e.g air

4 .Solid-liquid mixtures e.g water and sand

5. Liquid-gas mixtures e.g water containing dissolved gases.

Examples of mixtures;

- Ink, milk, paint, Sea water, air
- Alloys e.g Brass, Steel, Bronze etc.

ALLOYS

An alloy is a uniform mixture of one metal with one or more other substances usually metals or carbon.

Common alloys, their composition and uses

Alloy	Components	uses
Solder	Lead and tin	Making electrical contacts and joining pipes
Brass	Copper and Zinc	making screws, light bulb caps and condenser tubes
Duralumin	Aluminium, magnesium and copper	making parts of aircraft
Steel	Iron and carbon	light weight machinery, utensils
		making car bodies, railway lines, building industry
Stainless steel	Iron, Chromium and Nickel	making cutlery, kitchen utensils, razor blades.

Differences between a mixture and a compound.

Mixture	Compound
<ul style="list-style-type: none">• The substances in it can be separated by physical means• Its properties are the average of those of the substances in it• Energy is not usually given out or absorbed when its made• It's composition is variable	<ul style="list-style-type: none">• Elements in it can't be separated by physical means• Its properties are quite different from those of the elements in it.• Energy is usually given out or absorbed when its made.• It's composition is not variable

SPEC- Separated, Properties, Energy, Composition.

METHODS OF SEPERATION OF MIXTURES

1. FILTRATION:

This method is used to separate a soluble component from insoluble component.

Example: *Seperation of a mixture of sand and common salt.*

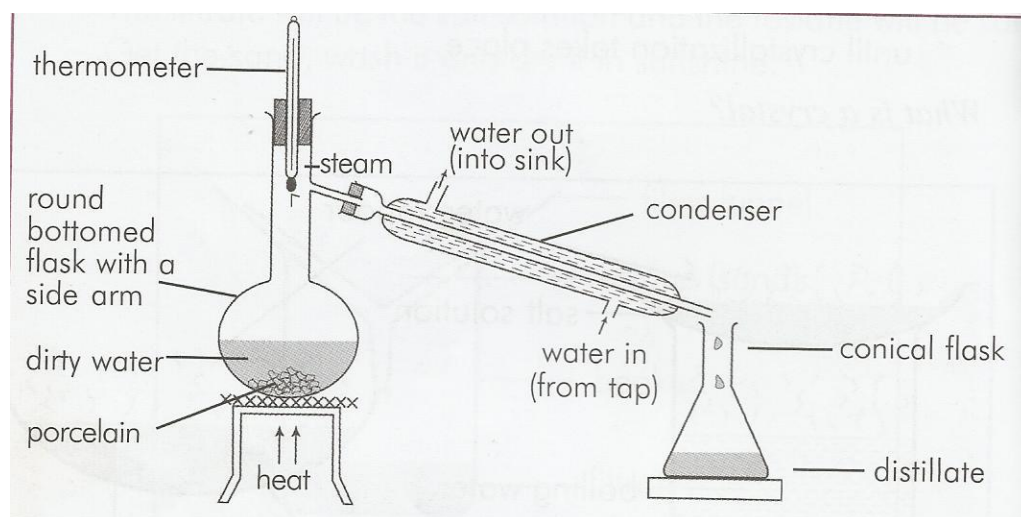
- Place a mixture of sand and salt in a beaker and add enough water to dissolve.
- Stir the mixture with stirring rod until all the salt dissolves.
- Filter the mixture and collect the filtrate(salt solution) in the evaporating dish
- Evaporate the salt solution to dryness.
- The sand remains on the filter paper as the residue.

2. DISTILLATION:

Distillation is a process of boiling a liquid to form vapour and then cooling the vapour to obtain the liquid.

Example: *To obtain pure water from impure water.*

- Pour water into a distilling flask.
- Add sand and copper (II) sulphate to colour the water.
- Pass the water up the condenser until it flows slowly into the sink.
- Heat the flask and boil the water, the steam condenses in the condenser and collects as a distillate in a receiving flask.



3. FRACTIONAL DISTILLATION.

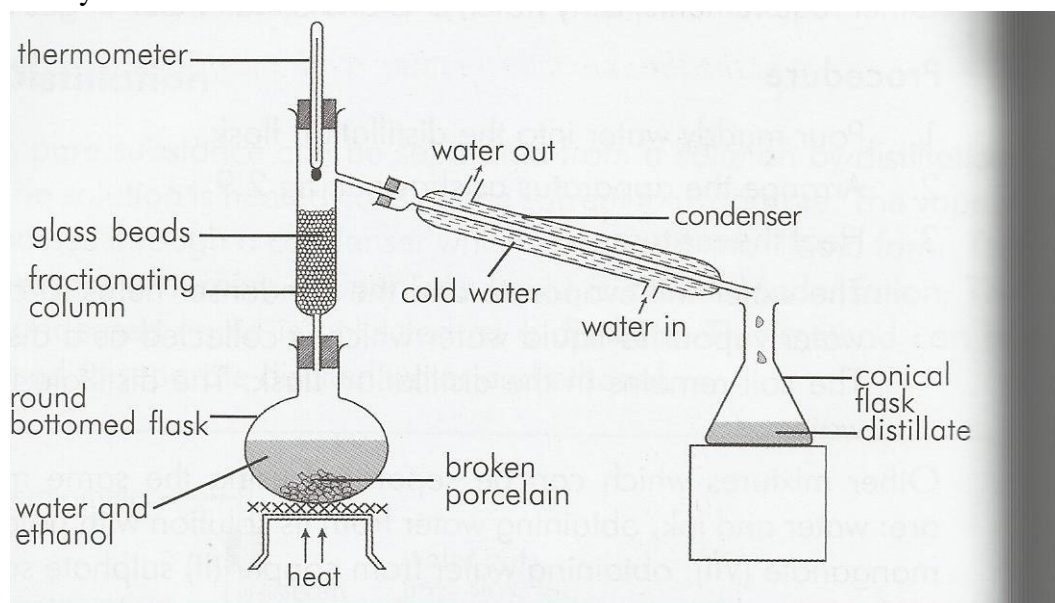
Fractional distillation is the process of separating substances from a mixture by making use of the different boiling points of the substances. The distillate is collected as fractions which boil at different temperatures.

Example: Separating a mixture of ethanol and water

Mix some ethanol with water and add the mixture to a distilling flask and fit the flask with a fractionating column.

Heat the mixture and collect the distillate formed until the thermometer records 82°C .

Continue heating and again collect the distillate boiling between 95°C and 100°C , which is mainly water.



A fractionating column provides a large cooling surface. The glass beads ensure that vapour and liquid mix well.

Application of fractional distillation.

- (i) Separation of liquid air into oxygen and nitrogen.
- (ii) Separation of crude oil into fractions e.g Petrol, diesel, kerosene etc.

The principle behind separation by fractional distillation is that the liquids must have different boiling points.

4. SUBLIMATION.

Sublimation is the change of state from solid to gas without passing through the liquid state.

Example.

Separating a mixture of sodium chloride and iodine.

- A mixture of iodine and sodium chloride is put in a glass beaker.
- The mixture is heated so that the iodine crystal changes to purple vapour.
- Iodine vapour condenses on cold surface forming iodine sublimate

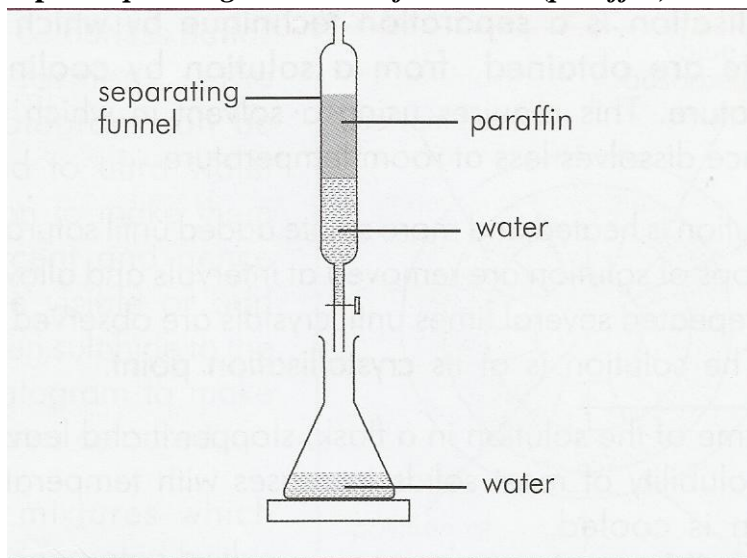
Other mixtures which can be separated by this method are:-

1. Ammonium chloride and potassium chloride
2. iron (III) chloride and sand.
3. Iodine and sand.

5. USE OF A SEPERATING FUNNEL

This method is used to separate immiscible liquids. Immiscible liquids are liquids which do not mix and form layers. The less dense liquid settles on top of the denser one.

Example: Separating a mixture of kerosene (paraffin) and water



- Pour some kerosene in a separating funnel followed by water and shake vigorously.
- Open the tap, the lower layer (water) is completely run off.
- Pour the remaining liquid (paraffin) into another container.

Other mixtures which can be separated by this method are:-

1. Petrol and water
2. Water and oil

6. USE OF A MAGNET (MAGNETISM)

This is a method of separating a magnetic substance from a non magnetic substance using a magnet.

Example: *Separation of a mixture of iron filings and sulphur*

- Put a mixture of iron filings and sulphur in a Petri-dish.
- Pass a magnet over the mixture in a Petri dish, iron filings are attracted to the magnet leaving behind sulphur.

Chemically the mixture of iron filings and sulphur can be separated by use of carbon disulphide.

Sulphur dissolves in carbon disulphide leaving iron filings to settle at the bottom. Filter the mixture to obtain iron. Warm the solution of sulphur in carbon disulphide to evaporate it and collect sulphur crystals.

7. CHROMATOGRAPHY

Chromatography means ‘colour writing’. Paper chromatography is a process for separating coloured substances by using porous paper e.g. filter paper. The substances are moved over the paper at different rates by a solvent. The components are separated into bands called a chromatogram.

Example: *Separating the solutes in ink*

- Lay a filter paper flat on the rim of an evaporating basin or small dish.
- Add one drop of ink to the centre of the paper. Wait until the ink stops spreading and then add a second drop. Wait until the ink dries, leaving a ring of coloured solutes.
- Add one drop of water (solvent) to the middle of the coloured ring. The water spreads and moves the solutes. Wait until the water is almost dry and add a second drop. Continue until a disc of coloured substance almost reaches the edge of the paper.

The dried paper with separate discs of coloured solutes obtained is called a chromatogram.

Note. The solute which has the highest solubility in the solvent travels fastest and ends up nearest to the solvent front. The least soluble solute travel slowest and ends up nearest to the original spot.

8. CRYSTALLISATION.

Crystallization is a separation technique by which crystals of a solute are obtained from a solution by cooling to room temperature e.g separation of copper (II) sulphate from water.

9. FRACTIONAL CRYSTALLISATION

This is a method of separating a mixture by repeated partial crystallization of the solution. The principle behind the separation is the different solubilities of the solutes in a given solvent.

Examples of mixtures separated by this method:

1. Sodium carbonate and sodium hydrogen carbonate
2. Potassium chloride and potassium chlorate

10 EVAPORATION

This is a method used to separate a solid from a liquid in which it has dissolved e.g. separating a mixture of sodium chloride and water.

11. DECANTING.

This method is used to separate a mixture of solid and liquid where by the solid settles on standing e.g. chalk powder and water, sand and water

Exercise

- 2(a). What is chromatography?
- (b). State the principle behind separation of mixtures by chromatography.
3. State the method by which each of the following mixtures can be separated:
 - (a) Water and ethanol
 - (b) Water and oil
 - (c) Ink
 - (d) Sand and iodine
 - (e) salt and sand

SOLUTIONS AND CRYSTALS

Miscible and immiscible liquids

Miscible liquids are liquids which mix completely e.g. water and ethanol, benzene and paraffin. Immiscible liquids are liquids which do not mix and form two separate layers e.g. paraffin and water, petrol and water, oil and water.

Solutions

When you shake sugar with water, the sugar mixes uniformly with the water and each part of the liquid is exactly the same. Sugar is soluble and dissolves in water. The mixture of water and sugar is called a solution.

1. Solution

A solution is a uniform mixture of two or more substances.

Saturated solutions

If we shake a lot of sugar with water, some of the solid dissolves but the rest does not. It settles to the bottom of the water on standing. The water can not dissolve any more solid and the solution is said to be saturated.

A saturated solution of a solute at a particular temperature is one which can dissolve no more solute.

2. Solute.

A solute is a dissolved substance. It can be solid, liquid or gas.

3. Solvent:

A solvent is a substance which dissolves a solute. A solution in water is an aqueous solution. Substances that are insoluble in water may be soluble in other solvents.

Rubber solution is used to mend holes in tubes of bicycles and cars; it is a solution of rubber in petrol.

Iodine solution is used on cuts and wounds on the body; it is a solution of iodine in ethanol.

4. A suspension

A suspension is a liquid containing small particles of solid spread throughout it and which settle on standing.

A suspension differs from a solution in the following ways:

- It contains solid particles which can be seen.
- It's solid particles settle on standing.
- Separation is by filtration which separates it into a filtrate and residue.

Crystals

Some solids are usually more soluble in hot solvents than in cold. When a hot saturated solution cools, some solid settles out. If the pieces of the solid have a regular shape, flat sides and sharp edges, they are called crystals.

Crystals form when a cold saturated solution stands in air for some time. Part of the solvent evaporates and the rest can not hold all the solid in solution. The excess solid settles out as crystals.

A crystal is a solid that has solidified in a definite regular shape.

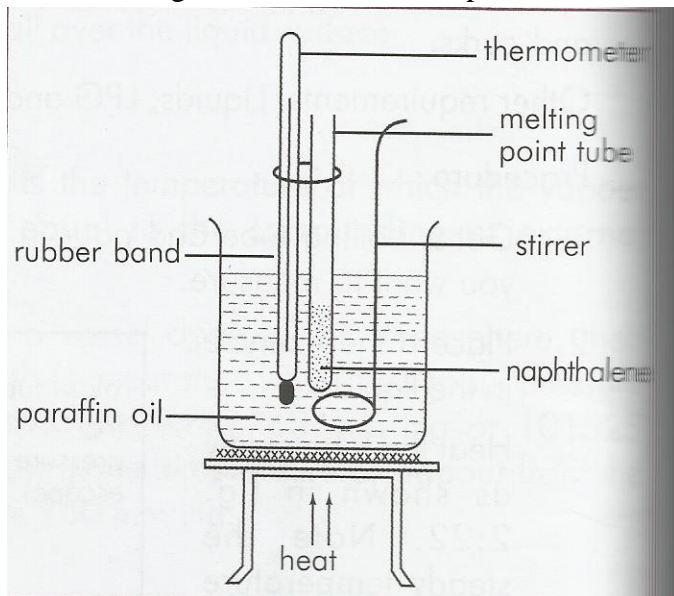
TEST OF PURITY

Melting point: This is the temperature at which a pure solid changes to liquid. Melting point is a characteristic of each substance and can be used to determine whether a substance is pure or not.

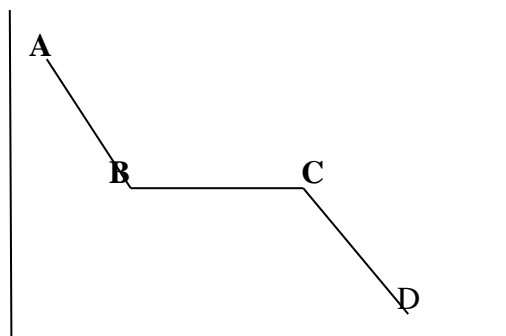
Measuring melting point of Napthalene

- Put some Napthalene powder into capillary tube.

- Tie the thermometer to the capillary tube using rubber bands.
- Place the capillary tin a beaker of cooking oil.
- Heat the oil gently while stirring continuously
- Note the temperature at which Napthalene melts
- The solid will begin to melt and the temperature is read off as the melting point of the solid.



Cooling curve for Naphthalene.



AB - Cooling of liquid naphthalene up to B where solidifying starts.

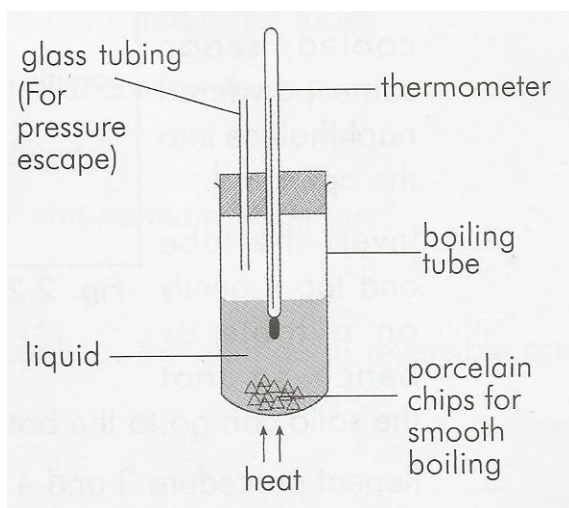
BC - Solidifying of Naphthalene at constant temperature.

CD - Cooling of solid Naphthalene.

Measuring boiling point of a liquid.

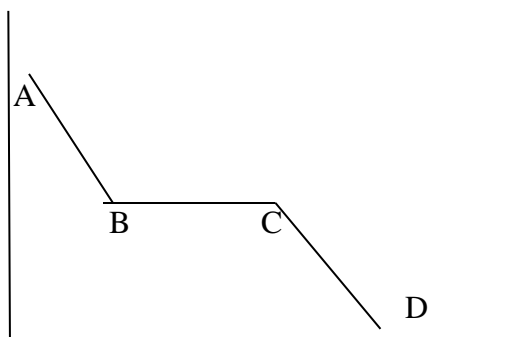
Boiling point: This is a temperature at which a liquid changes to vapour. Boiling occurs when the vapour pressure of a liquid is equal to the external pressure.

- Place a liquid whose boiling point is to be determined in a boiling tube.
- Place the thermometer in the liquid in the tube.
- Heat the liquid gently until a steady temperature when boiling occurs.



Exercise

1. The figure below shows a cooling curve of Naphthalene



(a) State what is happening at point marked

AB

BC

CD

(b) What is the melting point of Naphthalene

(c) The table below shows the melting points and boiling points of pure substances.

Substances	MP°C	BP°C
P	-90	-30
Q	-8	60
R	-5	230
S	40	250
T	330	1800

Which substance Q(s) is/are

(i). gases at room temperature?

(ii). Liquids at room temperature?

(iii). Solids at room temperature?

2. Explain the the following terms;

- (i) Solute
- (ii) Solvent
- (iii) Solution
- (iv) Suspension

(b) Give two examples in each case of;

- (i) Solvent
- (ii) Solute
- (iii) Solution

3. (a) Name one method by which the following mixtures can be separated:

- (i) Ethanol and water
- (ii) Oil and water
- (iii) Sand and water
- (iv) Iron filings and sulphur
- (v) Green pigments in a leaf
- (vi) Salt and water
- (vii) Iodine and sand
- (viii) Sodium carbonate and sodium hydrogen carbonate
- (ix) Copper (II) sulphate and water
- (x) Blood cells from blood sample.

(b) State the principle property used in the methods;

- (i) (a)(i)
- (ii) (a)(ii)
- (iii) (a)(iv)
- (iv) (a)(iv)
- (v) (a)(vii)
- (vi) (a)(viii)

4.(a) Giving two examples in each case, define the following terms:

- (i) Immiscible liquids
- (ii) Miscible liquids

(b) The table below shows results of a paper chromatography experiment.

•	•			•		
•						•
	•				•	
•			•			
•	•	•				
P	Q	A	B	C	D	E

P and Q are different mixtures of some of the pure substances A,B,C,D and E.

(a). Identify the substances in the ;

- (i). Mixture P

(ii).Mixture Q

(b). Which substances are present in both mixtures

(c). Which substances are present in mixture Q only?

(d). Which substances are present in mixture P only?

5.(a).Define the term compound?

(b) Classify each of the following substances into element,compound or mixture.

(i) Potassium chloride

(ii) milk

(iii)Paint

(iv)Chlorine

(c) Give any four differences between metals and non metals.

Metals	Non metals
(i)	
(ii)	
(iii)	
(iv)	

(d) Define the following terms:

(i) Element

(ii) Mixture

(iii) Compound

6.(a) .Define the term Alloy?

(b)Give the composition of the following alloys:

(i). Brass

(ii).Solder

(iii).Bronze

(iv).Stainless steel

(v).Duralumin

(c). Give the uses of;

(i). Brass

(ii).Duralumin

(iii).Stainless steel

7.(a). State four differences between a compound and a mixture.

Mixture	Compound
(i)	
(ii)	
(iii)	
(iv)	

(b).Give reasons why water is a compound and not a mixture.

8.(a).Write the chemical symbols for the following elements:

- (i). Sodium
- (ii).Iron
- (iii).Magnesium
- (iv).Aluminium
- (v).Copper
- (vi).Silicon
- (vii).Argon
- (viii).Helium
- (ix).Silver
- (x).Chlorine

(b).Write the names of the following elements whose symbols have been written:

- (i).Pb
- (ii).Be
- (iii).S
- (iv).P
- (v).Ca
- (vi).K
- (vii).Ne
- (viii).Zn
- (ix).Hg
- (x).H

FORMULAE AND EQUATIONS

Chemical formulae

The formula of an element or compound is the symbols and numbers which mean one molecule.

Molecules

A molecule is the smallest particle of an element or compound which can normally exist on its own. Substances such as water and gases consist of molecules.

Some substances contain atoms of the same elements while others consist of atoms of different elements chemically combined together.

Atomicity is the number of atoms in one molecule of the element.

Types of molecules

1. Diatomic molecules - this consist of two atoms of the same element.

Examples:

H₂ - Hydrogen molecule

O₂ - Oxygen molecule

Cl₂ – Chlorine molecule

N₂ - Nitrogen molecules

2. Tri-atomic molecules – This consist of three atoms of the same element e.g O₃

3. Tetra- atomic molecules – This consist of four atoms of the same element e.g P₄

4. Polyatomic molecules - This consist of many atoms of the same element e.g S₈

VALENCY

This is the power of an element or radical to combine with others. The valency of an element or radial is the number of hydrogen atoms that it will combine with or that it will displace from an acid. The size of charge on an ion is valency.

VALENCIES OF ELEMENTS

Metals

Potassium	1
Sodium	1
Lithium	1
Silver	1
Calcium	2
Magnesium	2
Copper	2
Iron	2 or 3
Lead	2
Zinc	2
Barium	2
Mercury	2
Aluminium	3

Non- metals

H	1
Cl	1
O	2
S	2
N	3
C	4

RADICALS

A radical is a group of atoms that exist in several compounds but does not exist on its own. They are a group of atoms that carry a charge. Radicals may be negatively or positively charged ions.

Name	Formulae	Valency
Nitrate	NO ₃	1
Ammonium	NH ₄	1
Hydroxide	OH	1
Chloride	Cl	1

Hydrogen carbonate	HCO_3	1
Hydrogen Sulphate	HSO_4	1
Nitrite	NO_2	1
Carbonate	CO_3	2
Sulphate	SO_4	2
Sulphite	SO_3	2
Oxide	O	2
Phosphate	PO_4	3
Nitride	N	3

WRITING CHEMICAL FORMULAE

Formulae are written using symbols and valencies of elements and radicals

Steps in writing chemical formula

1. Write the symbols of the elements and formula of the radicals involved
2. Write the valencies of the species in the top right-hand corner.
3. Reverse the valencies and write them on the bottom right-hand corner. Never write **1** in the chemical formula.

Examples.

(a) To write the formula of calcium chloride

Step1 Ca Cl

Step2 Ca^2 Cl^1

Formula CaCl_2

(b) To write the formula of magnesium carbonate.

Step Mg CO_3

Step Mg^2 CO_3^2

Formula MgCO_3

(c) To write the formula of Zinc nitrate

Step1 Zn NO_3

Step Zn^2 NO_3^1

Formula $\text{Zn}(\text{NO}_3)_2$

(d) To write the formula of carbondioxide

Step1 C O

Step2 C^4 O^2

Formula CO_2

(e) To write the formula of water.

Step1 H O

Step2 H^1 O^2

Formula H_2O

(f) To write the formula of iron (III) hydroxide

Step1 Fe OH

Step2 Fe^3 OH^1
Formula $\text{Fe}(\text{OH})_3$

Exercise

Write the chemical formulae of the following compounds:

1. Sodium Carbonate
2. Ammonium Sulphate
3. Copper (II) hydroxide
4. Sodium Oxide
5. Calcium hydrogen carbonate
6. Potassium hydrogen carbonate
7. Ammonium Phosphate
8. Calcium Phosphate
9. Iron (III) Sulphate
10. Aluminium Sulphate

CHEMICAL EQUATIONS

A chemical equation represents a chemical change or reaction by means of symbols and formulae.

A word equation is written using the chemical names of the substances in the reaction.

Example:

Magnesium burns in oxygen to form magnesium oxide

Magnesium + Oxygen \longrightarrow Magnesium Oxide.

The substances on the left hand side of the equation are called reactants and those on the right hand side are products.

The symbol + on the left of equation means **reacts with** and on the right of the equation, it means ‘**and**’

The arrow \longrightarrow means to form

State symbols

State symbol for solid is (s)

State symbol for liquid (l)

State symbol for gas (g)

State symbol for an aqueous solution (aq).

Rules for writing Balanced chemical equations.

1. On the left hand side, write formulae or symbols of reactants, on the right write formulae of products.
2. Write a word equation for the reaction
3. Balance the equation by writing small whole numbers in front of one or more formulae.

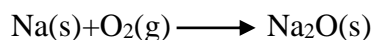
The process of making the number and kind of atoms equal on both sides of an equation is called Balancing equation

Equations

Examples on writing balanced equations.

1. Equations where only one product is formed

(a) Sodium + oxygen \longrightarrow sodium oxide



L.H.S

R.H.S

Na=1

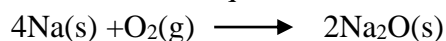
Na=2

O=2

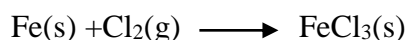
O=1

On the left-hand side of the equation, the sodium atom is one, on the R.H.S the sodium atoms are two. To balance the sodium atoms insert a 2 before sodium atom on the left. The oxygen atoms on the right are balanced by inserting a 2 before sodium oxide. But this produces 4 sodium atoms On the right while the left remains with 2, therefore to balance the equation insert a 4 on sodium atom on the left.

Thus chemical equation is:



(b).Iron + chlorine \longrightarrow iron (III) chloride



L.H.S

R.H.S

Fe=1

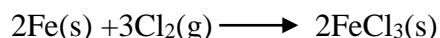
Fe=1

Cl=2

Cl=3

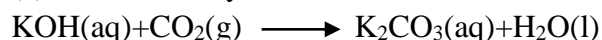
On both sides of the equation there is 1 iron atom, while the chlorine atoms on the L.H.S are 2, the chlorine atoms on the R.H.S are 3. To balance the chlorine, we insert a 3 before chlorine on the L.H.S and insert a 2 before iron (III) chloride making the number of chlorine on both sides 6, then insert a 2 before iron on the L.H.S.

Thus the chemical equation is:



2. Equations in which two products are formed

(a) Potassium hydroxide + carbon dioxide \longrightarrow potassium carbonate+ water



L.H.S

R.H.S

K=1

K=2

O=3

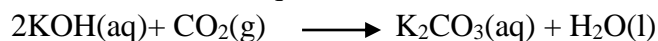
O=4

H=1

H=2

When we insert a 2 before KOH, the number of atoms of K=2, H=2 and O=2 hence making total number of O on L.H.S=4.

Thus the chemical equation is:



(b) Iron + hydrochloric acid \longrightarrow iron (II) chloride + hydrogen



L.H.S

R.H.S

Fe=1

Fe=1

H=1

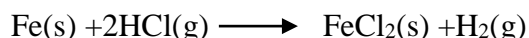
H=2

Cl=1

Cl=2

The number of hydrogen atoms on L.H.S =1 and on the R.H.S=2, to make the number of hydrogen balance, insert a 2 in front of hydrogen chloride.

Thus the chemical equation is :



3. Equations in which three products are formed

(a) Decomposition of lead (II) nitrate crystal \longrightarrow lead (II) oxide +nitrogen dioxide oxygen



L.H.S

R.H.S

Pb=1

Pb=1

N=2

N=1

O=6

O=5

The numbers of oxygen atoms on the L.H.S are 6 and on the R.H.S they are 5. The atoms are not even in numbers because of one oxygen atom on lead oxide. To make the oxygen atoms equal, insert a 2 in front of PbO, a 2 in front of Pb(NO₃)₂. To make nitrogen atom equal, insert a 4 in front of NO₂.

Thus the equation is:



Exercise

1. Write balanced chemical equations from the following word equations:

(a) Sulphur + oxygen \longrightarrow sulphur dioxide

(b) Aluminium +oxygen \longrightarrow Aluminium oxide

(c) Iron + oxygen \longrightarrow Iron(III) oxide

(d) Calcium + chlorine \longrightarrow Calcium chloride

(e) Magnesium + nitrogen \longrightarrow Magnesium nitride

2. Balance the following equations:

(a) $\text{KClO}_3\text{(s)} \longrightarrow \text{KCl(s)} + \text{O}_2\text{(g)}$

(b) $\text{Fe(s)} + \text{O}_2\text{(g)} \longrightarrow \text{Fe}_3\text{O}_4\text{(s)}$

(c) $\text{Fe(s)} + \text{Cl}_2\text{(g)} \longrightarrow \text{FeCl}_3\text{(s)}$

(d) $\text{Mg(s)} + \text{CO}_2\text{(g)} \longrightarrow \text{MgO(s)} + \text{C(s)}$

(e) $\text{Mg(s)} + \text{N}_2\text{(g)} \longrightarrow \text{Mg}_3\text{N}_2\text{(s)}$

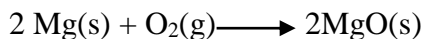
Principles of forming equations

To form equations, one needs to consider the different types of reactions, which include:

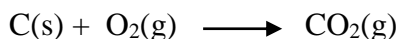
1. Combination reaction

This reaction occurs when two or more elements combine to form a single substance e.g. combustion of elements in air, in chlorine and reaction of metals with sulphur to form metal sulphides, and reaction of nitrogen with metals to nitrides.

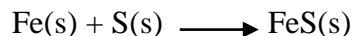
(i) Metal + oxygen \longrightarrow metal oxide



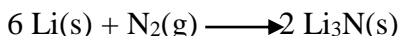
(ii) Non metal +oxygen \longrightarrow non metal oxide



(iii) Metal + sulphur \longrightarrow metal sulphide



(iv) Metal + nitrogen \longrightarrow metal nitride



2. Decomposition reaction

This reaction occurs when a compound splits up into simpler substances.

(i) Metal carbonates decompose to metal oxide and carbon dioxide.



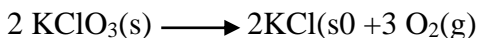
(ii) Metal hydrogen carbonates decompose to form metal carbonate + water + carbon dioxide



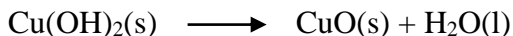
(iii) Metal nitrates (Cu, Pb, Ca, Zn, and Mg) decompose to form metal oxide + nitrogen dioxide + oxygen



(iv) Potassium chlorate decomposes \longrightarrow potassium chloride + oxygen.

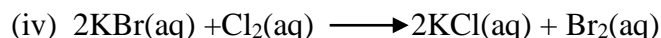
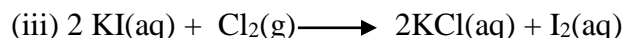
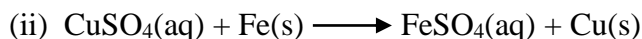
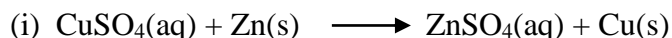


(v) Some metal hydroxides decompose \longrightarrow metal oxide + water



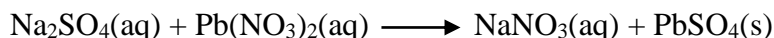
3. Displacement reactions

This reaction occurs when one element (more reactive) takes the place of another element (less reactive) from its solution.



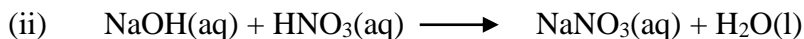
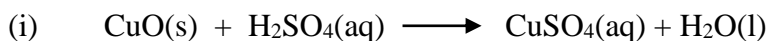
4. Precipitation reaction (double decomposition)

In this type of reaction, both compounds involved decompose and two new substances are formed by exchange of radicals.

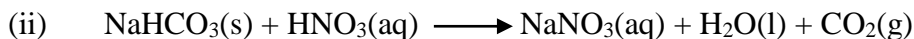
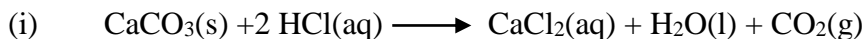


5. Neutralization reaction

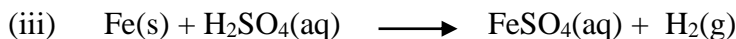
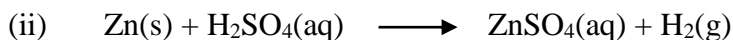
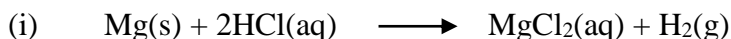
In this reaction an acid reacts with a base (metal oxides and hydroxides) to form a salt and water only.



6. **Acids** react with carbonates and hydrogen carbonates to form salt, water and carbon dioxide.



7. **Acids** react with metals (Mg, Zn, and Fe) to liberate hydrogen gas



Exercise

1. Complete and balance the following chemical equations:

1. $\text{Pb}(\text{NO}_3)_2(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \longrightarrow$
2. $\text{Mg}(\text{s}) + \text{Cl}_2(\text{g}) \longrightarrow$
3. $\text{Na}(\text{s}) + \text{O}_2(\text{s}) \longrightarrow$
4. $\text{Pb}(\text{NO}_3)_2(\text{aq}) + \text{Zn}(\text{s}) \longrightarrow$
5. $\text{Fe}(\text{s}) + \text{H}_2\text{SO}_2(\text{aq}) \longrightarrow$
6. $\text{FeSO}_4(\text{aq}) + \text{Zn}(\text{s}) \longrightarrow$
7. $\text{K}_2\text{SO}_2(\text{aq}) + \text{Pb}(\text{NO}_3)_2(\text{aq}) \longrightarrow$
8. $\text{ZnSO}_4(\text{aq}) + \text{NaOH}(\text{aq}) \longrightarrow$
9. $\text{AgNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \longrightarrow$
10. $\text{H}_2\text{O}_2(\text{l}) \longrightarrow$
11. $\text{KNO}_3(\text{s}) \longrightarrow$

2. (a). Define the following terms:

- | | |
|----------------|----------------|
| (i). Valency | (ii). Atom |
| (iii). Radical | (iv). Molecule |
| (v). Element | |

(b). Using symbols, valency of radical and elements, write the chemical formulae for the following compounds:

- | | |
|-------------------------------|------------------------|
| 1. Iron (II) sulphate | 2. iron (III) sulphate |
| 3. Calcium nitrate | 4. Aluminium sulphate |
| 5. Ammonium phosphate | 6. Aluminium oxide |
| 7. Lead nitrate | 8. Potassium carbonate |
| 9. Calcium hydrogen carbonate | 10. Magnesium nitride |
| 11. Sodium oxide | 12. Calcium phosphate |
| 13. Sodium hydrogen sulphate | 14. Aluminium chloride |
| 15. Calcium hydroxide | 16. Sodium sulphite |
| 17. Carbon dioxide | 18. Sodium sulphate |

(c). Give the chemical names for the following compounds:

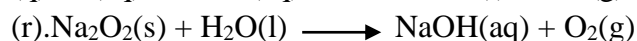
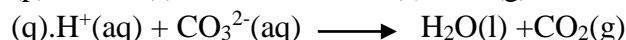
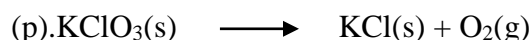
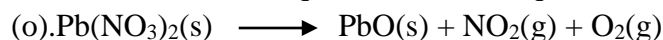
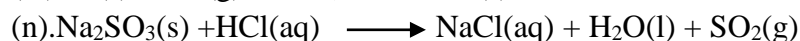
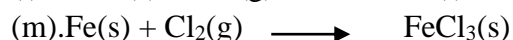
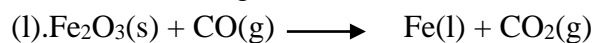
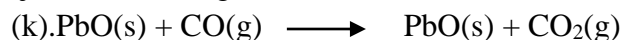
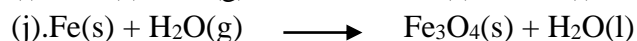
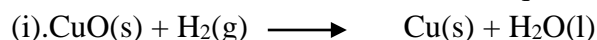
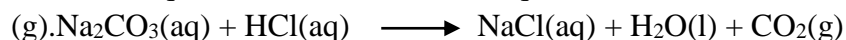
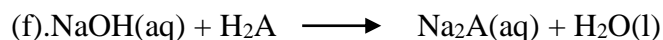
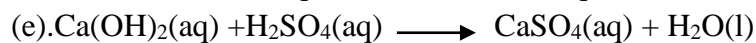
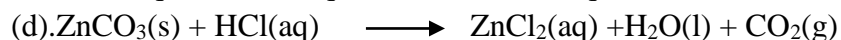
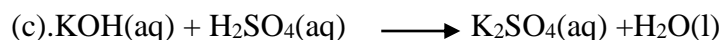
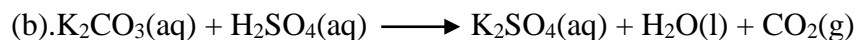
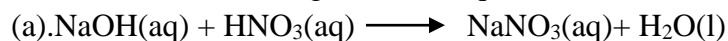
- | | |
|------------------------------------|-----------------------------------|
| (i). $\text{Cu}(\text{NO}_3)_2$ | (ii). AlCl_3 |
| (iii). $\text{Fe}(\text{OH})_3$ | (iv). $\text{Mg}(\text{HCO}_3)_2$ |
| (v). NaOH | (vi). CO |
| (vii). SO_2 | (viii). H_2SO_4 |
| (ix). $(\text{NH}_4)_2\text{CO}_3$ | (x). AgCl |

3. Complete and balance the following equations:

- (a). $\text{Ca}(\text{NO}_3)_2(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \longrightarrow$
- (b). $\text{K}_2\text{SO}_4(\text{aq}) + \text{Pb}(\text{NO}_3)_2(\text{aq}) \longrightarrow$
- (c). $\text{Pb}(\text{NO}_3)_2(\text{aq}) + \text{KI}(\text{aq}) \longrightarrow$
- (d). $\text{CuSO}_4(\text{aq}) + \text{NaOH}(\text{aq}) \longrightarrow$



4. Balance the following chemical equations:



AIR AND COMBUSTION

Air is a mixture of gases. Its composition varies at different times and in different places.

Air is a mixture and not a compound because of the following reasons:

- (i) Air can be made by mixing nitrogen and oxygen. No heat or light is given out.
- (ii) The nitrogen and oxygen (and other gases) can be separated from each other by physical means.
- (iii) The properties of air are just the average of the properties of nitrogen and oxygen.
- (iv) The proportions of nitrogen and oxygen (and other gases) in air are not definite, they vary from place to place and from time to time.

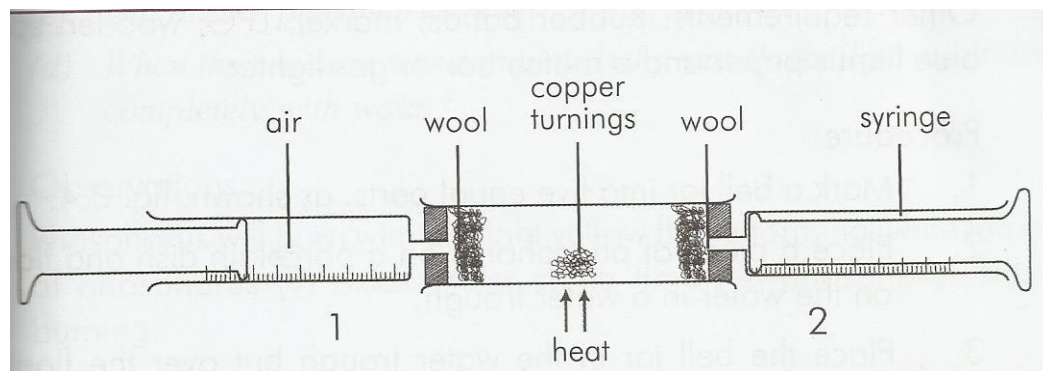
Composition of Air

The approximate composition of air by volume is as follows:

- | | |
|-----------------------------|---------------|
| 1. Nitrogen | 78% |
| 2. Oxygen | 21% |
| 3. Carbondioxide | 0.03% |
| 4. Inert / rare/noble gases | 1.0% |
| 5. Water vapour – variable | (0.5% - 4.0%) |

Oxygen is the reactive gas on which combustion, rusting and respiration depend. Nitrogen dilutes the oxygen and slows down the three processes. Noble gases such as helium, argon and neon are very inert.

EXPERIMENT TO DETERMINE THE PERCENTAGE BY VOLUME OF OXYGEN IN AIR USING COPPER



- Put some copper turnings into a combustion tube fitted to two syringes.
- Suck air into one of the gas syringes up to 100cm^3 and plug its mouth with cotton wool.
- Heat the combustion tube containing copper turning strongly and pass the air over copper turnings, by pushing the plunger of the syringe containing air to an empty syringe.
- Allow the combustion tube to cool and record the volume of air in the syringes. The decrease in volume of air is approximately 21cm^3 which is 21% of oxygen in air.

Note: The cotton wool prevents copper turnings from being blown into the syringes.

At the end of the experiment, the brown copper turnings turned black.

To determine the percentage of oxygen in air using a burning candle

Diag

- Fix a candle on a gas jar and stand
- Place the stand and candle in a trough and add water until the water level is about 2cm^3 up the candle.
- Lower a dry gas jar quickly over the burning candle.
- Lift the gas and push a burning splint in the gas.

Observation

- The candle is extinguished after a short time and the water level rises in the trough.

Conclusion

- Some part of air was used up by the burning candle. The water rose to occupy the space of the used air. The part used up is about 20% by volume. This part of air is oxygen.

COMBUSTION

Combustion (burning) is a chemical change in which substances combine with oxygen to form oxides and liberate energy.

To find out what substance is formed when a candle burns.

When a candle is lit, the suction pump is switched on. This pulls the gaseous products of the burning candle and keeps it burning. The products are passed through a cooled U-tube and lime water.

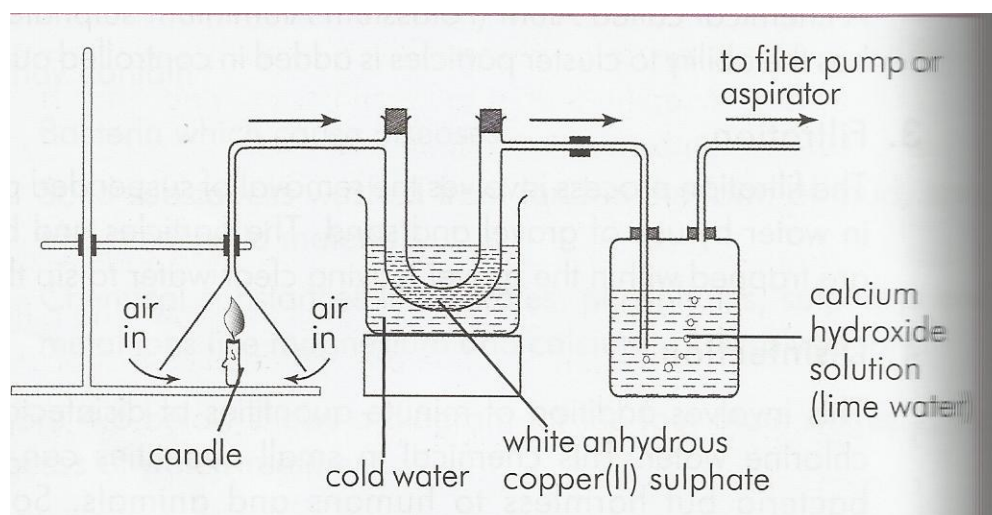
Observation.

A colourless liquid condenses in the U-tube and turns white anhydrous copper (II) sulphate to blue. The lime water turns milky.

Conclusion.

This shows that water and carbondioxide are formed when a candle burns in air.

Candle + Air \longrightarrow Water + Carbondioxide.

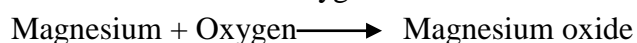


To find out if there is a change in mass when a metal burns in Air.

- Clean the magnesium ribbon with sand paper.
- Weigh 0.5g of magnesium ribbon; put it in a crucible and close the lid.
- Weigh the mass of the crucible, lid and magnesium.
- Heat the crucible gently.
- Raise the lid from time to time to allow in air to enter and combine with magnesium.
- Allow to cool and weigh the crucible, lid and magnesium oxide.

Observation.

There is increase in mass when a substance burns in air. The increase in mass is because the metal combines with oxygen of the air to form an oxide.



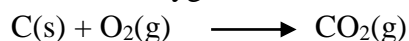
COMBUSTION OF ELEMENTS IN AIR

When substances burn in air, they combine with oxygen of the air to form oxides.

The elements can be heated on asbestos paper, porous pot, ceramic wool, crucible or metal bottle top.

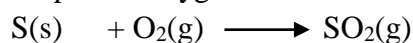
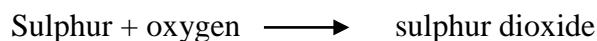
1. Carbon:

Carbon burns in air, becomes red hot and does not melt. It gradually burns away, forming a colourless gas called carbon dioxide.



2. Sulphur:

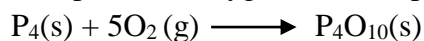
Sulphur melts and then burns with a bright blue flame, forming cloudy white fumes with a pungent; choking smell. The gas formed is a mixture of sulphur dioxide and sulphur trioxide.



3. White phosphorus

Phosphorus melts and burns readily with a bright yellow flame, forming white fumes of phosphorus (V) oxide. Red phosphorus burns less readily with a yellow flame.

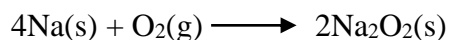
Phosphorus + oxygen \longrightarrow phosphorus (V) oxide.



4. Sodium

Sodium melts and then burns with a bright yellow flame, forming a yellow solid, sodium peroxide. In limited air, it burns with a yellow flame to form a white solid, sodium oxide.

Sodium + oxygen \longrightarrow Sodium peroxide



RUSTING OF IRON

When iron or steel wool is left in damp air for sometime, the iron becomes covered with a red – brown substance termed as rust. The rust forms only on the out side of the metal. Rust is soft and readily drops off and then the metal below starts to rust. Rust is hydrated Iron (III) Oxide, $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$.

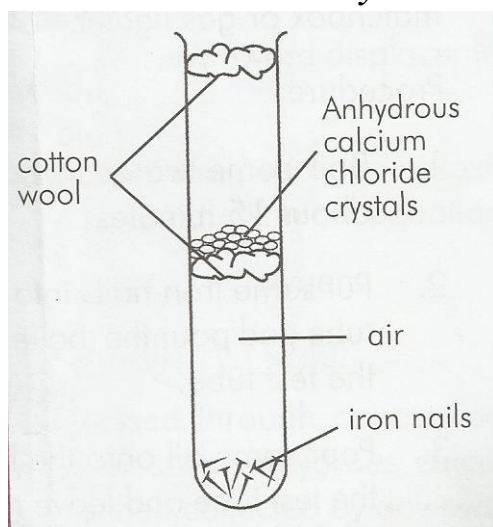
Rusting is a chemical change in which iron combines with water in the presence of air to form a red-brown substance.

Iron + Oxygen + Water \longrightarrow Iron rust

Rusting corrodes and wears off iron and reduces its strength. Rusting is a slow process.

To find conditions needed for Iron to rust.

(a). *To show that iron does not rust in dry Air.*

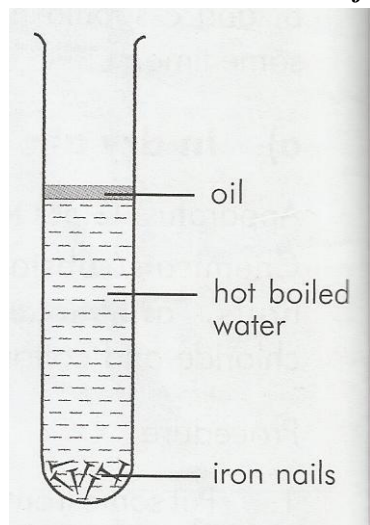


- Place some clean iron nails at the bottom of a glass test tube.
- Push in some cotton wool half way down the test tube.
- Add anhydrous calcium chloride above the cotton wool to keep the air in the tube dry.
- Close the mouth of the test tube with cotton wool and leave for several days.

Observation.

The iron nails do not rust.

(b). *To show that iron does not rust in air –free water*

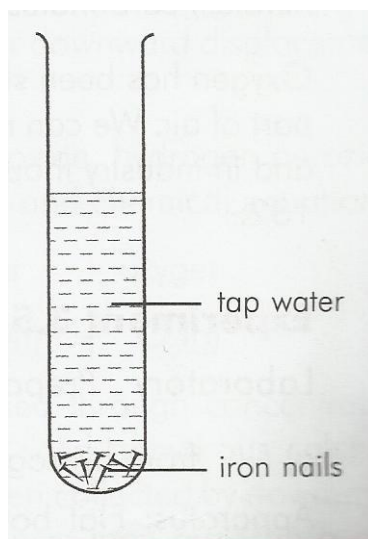


- Boil water in a test tube, to remove any dissolved air from water.
- Put some clean nails in the water in a glass test tube.
- Add oil or Vaseline to the water to form a layer on the surface. This keeps air out of the water.
- Leave the set up to stand for some days.

Observation.

The iron nails do not rust.

(c) *To show that iron rusts in both Air and Water.*



- Put some iron nails in a glass test tube
- Pour some water in the test tube and leave it to stand for some days.

Observation.

The iron nails rusted after a few days.

Conclusion.

Oxygen and water are necessary for rusting.

METHODS OF PREVENTING RUSTING

Rusting is prevented by keeping away water and air from iron. This can be achieved by:

1. Oiling or greasing the iron.

A layer of oil or grease is put on a metal. This cuts off air and stops metal tools and parts of machines from rusting. The oil or grease does not stay on the iron for long, and a fresh layer must be put on after some time.

2. Painting or tarring:

Paint is usually put on the iron parts of windows, doors, bicycles, cars etc. Tar is used instead of paint in parts that we do not see, e.g. bottom parts of bridges, ships. Tar is cheaper and lasts longer than paint.

3. Coating with other metals:

A layer of another metal that does not rust is put on the iron.

Iron coated with tin is called tin plate. Tin plate is used for making the 'tins' that contain food, paint, and petrol. Tin plate rusts quickly if the coat of tin is broken at any point.

Iron coated with zinc is called galvanized iron. It is used for making roofs, baths, buckets. It is better than tin-plate because it does not rust easily even if the zinc coat is broken.

Iron coated with chromium is chromium-plated. It is used on motor cars, bicycles and other things where a nice appearance is needed because it shines like silver.

4. Enamelling.

Enamel is very hard and does not let air or water pass through into the iron. It is used in cookers and refrigerators. It has no action on food and therefore quite safe to use on cooking vessels. Iron rusts easily if enamel is broken.

5. Using stain less steel:

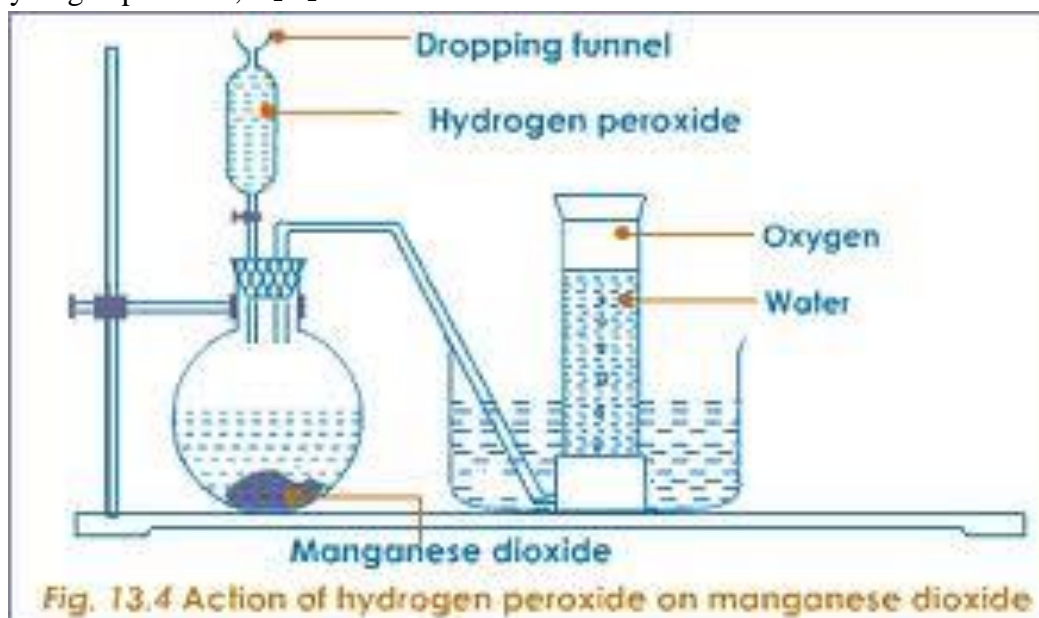
It is a very hard form of iron and contains carbon, chromium, and nickel. Knives, scissors and other tools are made from stainless steel.

OXYGEN

Oxygen is a gas that occupies 21% by volume of air.

Laboratory Preparation of oxygen

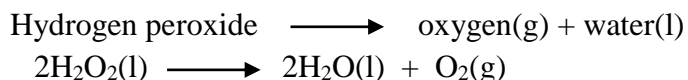
(a). From hydrogen peroxide, H_2O_2



- ❖ Put manganese (IV) Oxide powder at the bottom of the flask.
- ❖ Add hydrogen peroxide solution drop by drop using a dropping funnel on manganese (IV) oxide. Rapid effervescence occurs at room temperature. Manganese (IV) oxide acts as a catalyst.

A catalyst is a substance which changes the speed of a chemical reaction but remains chemically unchanged at the end of the reaction.

- ❖ The gas is collected by upward delivery as it's less dense than air.

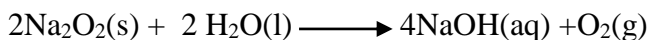


Note: If the gas is required dry, it's passed through concentrated sulphuric acid in a wash bottle or Solid anhydrous calcium chloride in a u-tube to dry it.

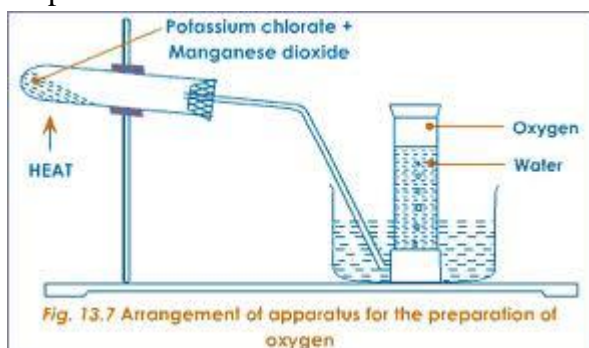
(b). From sodium peroxide

Water is added to solid sodium peroxide in a flask. Effervescence occurs and oxygen is evolved. The gas is then collected by upward delivery as it's less dense than air.

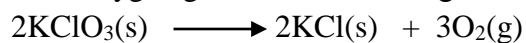
Equation.



(c) From potassium chlorate



A mixture of potassium chlorate and manganese (IV) oxide was heated in a hard glass tube. Oxygen gas is evolved. The gas is collected by upward delivery as it's less dense than air.



INDUSTRIAL MANUFACTURE OF OXYGEN

Oxygen is manufactured on an industrial scale by fractional distillation of liquid air. Nitrogen distills off at -196°C . Oxygen has Bp -183°C

Test for oxygen.

This is done by inserting a glowing wooden splint into a test tube containing oxygen.

Observation.

The glowing splint is relighted.

PROPERTIES OF OXYGEN

Physical Properties

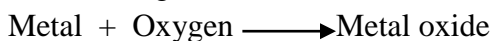
1. It's a colourless, odourless and tasteless gas.
2. It's neutral to litmus
3. It's slightly soluble in cold water
4. It's slightly denser than air.

Chemical properties of oxygen

Oxygen supports the burning of:-

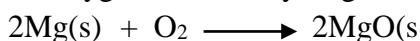
(a). Metals:

Metals such as magnesium, sodium, calcium, Iron etc burn in oxygen to form basic oxides.

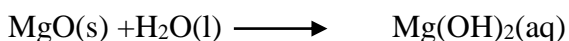


1. Magnesium

It burns in oxygen with a very bright white flame forming a white ash of magnesium oxide.

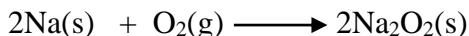


The oxide dissolves slightly in water to form magnesium hydroxide solution that turns litmus blue.

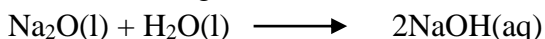
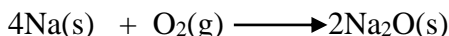


2. Sodium

Sodium burns with a brilliant yellow flame to form a yellow solid of sodium peroxide.

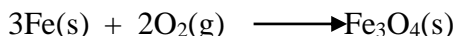


In limited oxygen supply, sodium burns to form a white solid of sodium oxide



3. Iron

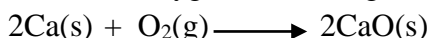
Iron burns in oxygen with bright sparks to form blue-black beads of tri-iron tetra oxide.



The oxide does not dissolve in water and has no effect on litmus.

4. Calcium

Calcium burns in oxygen with a bright red flame to form a white calcium oxide.



The oxide dissolves in water to form calcium hydroxide, which turns litmus blue.

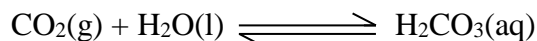
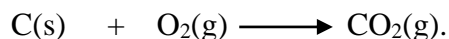
When metals(e.g magnesium,iron,sodium,calcium)burn in oxygen they form basic oxides,which dissolve in water to form alkalis.

(b). Non-metals:

Non-metals such as carbon, sulphur, phosphorus burn in oxygen to form acidic oxides.

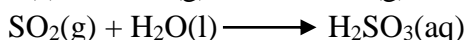
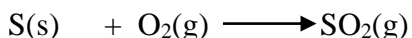
1. Carbon

Carbon burns with an orange flame and makes bright sparks.It forms carbon dioxide which turns calcium hydroxide milky.Carbon dioxide dissolves in water to form carbonic acid which turns blue litmus pale red.



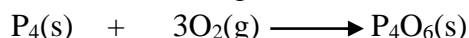
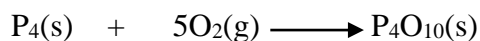
2. Sulphur

Sulphur burns with a bright blue flame to form cloudy fumes with choking smell.The gas is sulphur dioxide which dissolves in water to form sulphurous acid which turns blue litmus red.



3. Phosphorus

Phosphorus burns with a very bright yellow flame forming dense white fumes.The oxide dissolves in water to form phosphoric acid which turns blue litmus red.



USES OF OXYGEN

1. In oxy-acetylene flame for welding
2. in the manufacture of steel. Its used to remove impurities in iron
3. Liquid oxygen is used in space rockets.A mixture of liquid oxygen and hydrogen is a powerful rocket fuel.
4. In aiding respiration, life support in sea divers, hospital and mountain climbers.

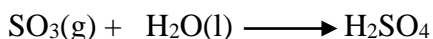
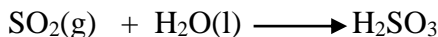
CLASSIFICATION OF OXIDES

1. Acidic oxides:

These are oxides of non metals which dissolve in water to form acids. They are called acid – anhydride.

Examples

- Sulphur dioxide, (SO₂)
- Sulphur trioxide, (SO₃)
- Nitrogen dioxide, (NO₂)
- Carbondioxide, (CO₂)



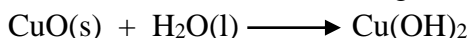
2. Basic Oxides

These are oxides of metals which react with acid to form a salt and water only.

Examples:

- Magnesium oxide, (MgO)
- Copper(II) oxide, (CuO)
- Calcium oxide, (CaO)

Basic oxides dissolve in water forming alkaline solutions.



3. Amphoteric oxides.

These are oxides which have both acidic and basic properties when they react with acids and alkalis

Examples

- Aluminium oxide, (Al₂O₃)
- Zinc oxide, (ZnO)
- Lead (II) oxide (PbO).

4. Neutral oxides

These are oxides which have neither acidic nor basic properties.

Examples

- Water, (H₂O)
- Carbon monoxide, (CO)
- Nitrogen monoxide, (NO)
- Dinitrogen oxide, (N₂O)

Neutral oxides are usually lower oxides of non metals.

5. Mixed oxides:

These are oxides which react like a mixture of two simpler oxides.

Examples

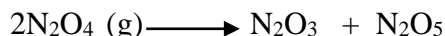
- Tri lead tetra oxide reacts like a mixture of lead(II) oxide and lead (IV) oxide.



- Tri iron tetra oxide reacts like a mixture of iron(II) oxide and iron (III) oxide



- Dinitrogen tetra oxide reacts like a mixture of dinitrogen trioxide and dinitrogen pentoxide.



6. Peroxides

Peroxides are oxides which contain twice as much oxygen as expected from the usual valency of the other elements in the oxides. Peroxides contain the O^{2-} ion and react with dilute acids to form hydrogen peroxide.

Examples

- Sodium peroxide, Na_2O_2
- Hydrogen Peroxide, H_2O_2
- Barium peroxide, BaO_2
- Potassium peroxide, K_2O_2

Exercise.

- (a) What is rust?
 - (b) Give the chemical name and formula of rust.
 - (c) State the conditions necessary for iron to rust.
 - (d) State the methods of preventing rusting.
- (a) Oxygen can be prepared in the laboratory from hydrogen peroxide and substance R.
 - (b) Name substance R and give its formula.
 - (c) What role does substance R play?
 - (d) Write equation for the reaction that takes place.
 - (e) State the uses of oxygen
- (a) Oxides are grouped into classes. Name the six classes of oxides.
 - (b) Categorise the oxides given below into their respective classes.
Nitrogen dioxide, nitrogen monoxide, copper oxide, zinc oxide, sodium peroxide, tri iron tetraoxide.
 - (c) Calcium was burnt in air.
 - (i) State what was observed
 - (ii) Write equation for the reaction that took place.
 - (d) The oxide in (c) was dissolved in water and the resultant solution tested with red litmus.
 - (i) State what was observed.
 - (ii) Write the equation for the reaction of the oxide with water.
- (a) What is combustion?
 - (b) Name the products of burning candle.
 - (c) State the tests for products of burning candle and the observations made.
- (a) Give reasons why air is a mixture and not a compound.
 - (b) Describe an experiment to determine the percentage of oxygen in air using copper of burning candle.

6. (a). State two major components of air.
- (b). (i) Name the process by which the components of air named in (a) can be separated.
- (ii). Explain why the process you have named in b(i) can be used to separate the components of air.
- (c). Give two industrial uses of one of the components of air that has the highest percentage by volume.
7. (a). What is atmosphere?
- (b). Name the components of air which is removed from air when:
- (i). Iron rusts
- (ii). Substances burn in air
- (iii). Air is passed over heated copper turning
- (iv). Air is passed through calcium hydroxide solution (Lime water).
- (c). Calculate the percentage of oxygen in a sample of air from the following experimental results:
- Initial volume of air = 40.0cm^3
- Volume of air remaining after reacting with metal = 31.60cm^3
8. (a). 79cm^3 of air was passed backwards and forwards over heated copper turnings in a combustion tube fitted with gas syringes plugged with cotton wool. The volume of air slowly reduced to 63cm^3 .
- (i). State what was observed in the combustion tube.
- (ii). Why were the syringes plugged with cotton wool?
- (iii). What caused the decrease in the volume of air?
- (iv). Name the gas used in this experiment.
- (b). Calculate the percentage of the gas named in the given sample of air above.
9. A sample of iron wool was placed in a boiling tube containing some water and the tube inverted in a trough of water. After 3 days the volume of air in the tube changed from 24.0cm^3 to 20.0cm^3 and a reddish-brown layer formed on iron wool.
- (a). Name the process taking place on cotton wool.
- (b). Give the:
- (i). common name of the reddish-brown layer on iron wool.
- (ii). Chemical name of the reddish-brown layer.
- (c). Write the chemical formula of the reddish-brown layer.
- (d). Calculate the percentage change in volume of air in the boiling tube.
- (e). A few drops of common salt were added to water in the boiling tube and iron wool placed in it. State whether the reddish brown layer formed in less, 3 or more than 3 days.
10. (a) Define the following terms, giving one example in each case;
- (i) Acidic oxide (ii) Basic oxide (iii) Neutral oxide (iv) Amphoteric oxide.
- (b) Some oxides are categorized as peroxides. What are peroxides? Give two examples of peroxides and write their formulae.

WATER AND HYDROGEN

WATER

Water is of fundamental importance to all kinds of plants and animals. Water supports vital processes necessary to life and growth of plants and animals such as digestion of food and photosynthesis.

Occurrence

Pure water does not exist in a natural state, but supplies of water are obtainable all over the world.

Water constitutes about 67% of the earth's surface and occurs in rocks, the atmosphere and in all living organisms.

Sources of water:

The natural sources of water are rain, spring water, river water and sea water.

THE WATER CYCLE

A cycle is a number of changes which go back to the starting point and which can be repeated. Water vapour forms in the air when the sun shines and when wind blows across water. Water vapour is less dense than air and therefore rises. As it rises it cools. Finally, it is so cold that the vapour condenses forming tiny drops of liquid water. These drops float in the sky and form a cloud.

The tiny drops of water in a cloud join together and fall to the ground as rain. The rain water passes into springs, rivers and seas. Once again it evaporates. Water on earth's surface is always changing into vapour, going into the air and then falling as rain.

CHEMICAL TESTS FOR WATER

Water can be tested chemically in two ways:

1. By using white anhydrous copper (II) sulphate which turns blue.
2. By using blue cobalt (II) chloride paper that turns from blue to pink.

PROPERTIES OF WATER

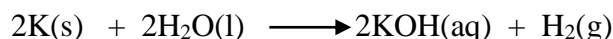
- Pure water is clear, colourless and odourless liquid.
- Water is neutral to litmus
- Water freezes at 0°C .
- Water boils at 100°C at sea level. It boils away completely and does not leave any residue.
- Water has a density of 1gcm^{-3} at 4°C .
- Water is an excellent solvent. It dissolves practically everything.

ACTION OF WATER AND STEAM ON METALS

Most reactive metals react with water forming metal hydroxide and hydrogen gas is liberated. The reaction varies from metal to metal depending on the position of the metal in the reactivity series.

Potassium:

Potassium reacts vigorously with cold water; it melts into a silvery ball and darts on the surface of water with a hissing sound and gives off hydrogen gas to form sodium hydroxide solution which turns red litmus blue.



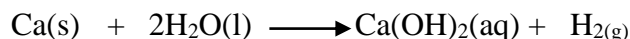
Sodium

Sodium melts into a silvery ball and darts on the surface of water with hissing sound. The hissing sound is due to evolution of hydrogen gas.



Calcium

Calcium sinks and reacts quietly to form bubbles of hydrogen gas. The lime-water becomes milky or cloudy because calcium hydroxide is not very soluble and forms a suspension.



Effects of steam on metals.

Magnesium: This metal reacts very slowly with cold water but reacts with steam.

Conditions for the reaction

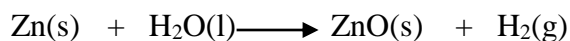
- A clean piece of magnesium ribbon
- Heating water to form steam.

Hot magnesium catches fire in steam and burns brightly to form a white ash of magnesium oxide.



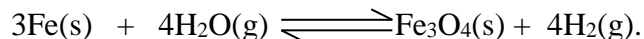
Zinc

Zinc reacts only when heated strongly in steam forming yellow solid when hot and white when cold. This solid is zinc oxide



Iron

Steam reacts slowly with heated iron to form a blue-black solid. This reaction is reversible.



Lead and copper do not react with cold water and steam.

WATER POLLUTION

Pollution is the presence in or introduction into the environment of a substance which has harmful effects to plant and animal life. These substances which bring about pollution are called pollutants.

Common water pollutants

- Industrial wastes e.g. oils, and gases such as sulphur dioxide dissolve and are washed by rain into water systems.
- Agricultural wastes e.g. herbicides, insecticides and fertilizers are washed down the streams.
- Sewage, mainly from homes and factories which find their way into water systems.
- Detergents, in homes and laundries, waste soap may escape into water bodies.

SEWAGE

This is water containing waste matter from toilets, bathrooms and factories. The sewage is normally carried away by underground pipes into prepared areas for treatment.

Sewage treatment

Sewage treatment involves the following steps:

1. Primary settlement tank
Solid wastes sink to the bottom and are sent to anaerobic digester.
2. Anaerobic digester.
Anaerobic bacteria digest the sludge producing methane gas. The digested sludge is sent to aeration tank.
3. Aeration tank.
Liquid from top of the primary settlement tank flows in the aeration tank. Air is bubbled into it where aerobic micro organisms break down any remaining organic matter.
4. Secondary settlement tank.
The activated sludge is sent back to aeration tank where aerobic micro-organisms digest the organic material.
5. Effluent
The liquid at the top of the secondary settlement tank which is clear has no smell and no disease causing organisms is called effluent. It's released into rivers or lakes.

Uses of sludge

- Road surfacing as it can set into a hard solid.
- As a fertilizer, it contains phosphorus and nitrogen.
- As a raw material to produce biogas
- To fill quarries for land reclamation.

WATER TREATMENT

There are four main stages in the process of water treatment.

1. Screening
This involves the use of screens or wire mesh to remove large floating objects from water.
2. Sedimentation.
This involves the clustering together of suspended particles into big lumps which sink and settle at the bottom. Alum (potassium Aluminium sulphate) which has the ability to cluster particles is added in controlled quantities.
3. Filtration.
This involves the removal of suspended particles in water by use of gravel and sand.
4. Chlorination
This involves addition of minute quantities of chlorine water. This destroy bacteria but harmless to humans. Soda ash (Na_2CO_3) may be added to normalize the PH of the chlorinated water.

USES OF WATER

- As a coolant and reactant in industry.
- As a solvent since it dissolves many substances
- In production of electricity
- For domestic purposes –(drinking, cooking, washing).

WATER IN THE ATMOSPHERE

Many substances absorb water from the air. Such substances are said to be hygroscopic.

Examples: anhydrous copper (II) sulphate, calcium oxide, sulphuric acid.

A hygroscopic substance is one which absorbs water from atmosphere.

Some substances absorb so much water from the atmosphere that they dissolve in it. Such substances are said to be deliquescent.

Examples: Calcium chloride, sodium hydroxide, potassium hydroxide, iron (III) chloride.

All deliquescent substances are hygroscopic.

Deliquescence is the absorbing of water from the atmosphere by a solid to form a solution.

Water of crystallization

Some substances form crystals that contain water. The water is chemically combined with the substance. The water comes off when the crystals are heated, and the crystals lose their shape and sometimes colour. A crystalline compound containing water is called *hydrated*, and the one without water is called *anhydrous*.

Hydrated salts include, copper sulphate crystals, Iron (II) sulphate crystals, sodium carbonate crystals, and sodium sulphate crystals.

Water of crystallization is the definite amount of water with which some substances chemically combine when they form crystals from their solutions in water.

Efflorescence

This is the giving up of water of crystallization by a crystal to the atmosphere.

HARD WATER

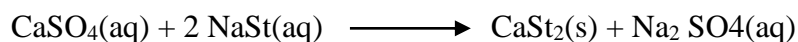
Hard water is water which does not readily form lather with soap.

Soft water readily forms lather with soap.

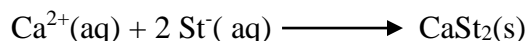
When soap is used in hard water, a white precipitate called scum forms and the water turns milky. Scum is an insoluble salt formed when soap reacts with Mg^{2+} or Ca^{2+} in hard water.

The chemical name for soap is sodium stearate (NaSt) and reacts with hard water according to the following equations.

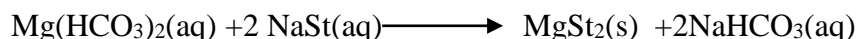
Permanent hardness



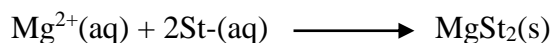
Ionic equation:



Temporary hardness

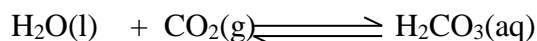


Ionic equation:

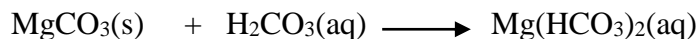
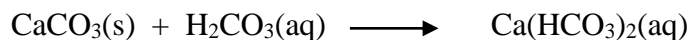


HOW WATER BECOMES HARD

(a). Rain dissolves carbondioxide as it falls forming dilute carbonic acid.



If this water passes through rocks containing calcium carbonate or magnesium carbonate, some dissolves to form calcium hydrogen carbonate or magnesium hydrogen carbonate.



(b). If the rain water passes through the rocks containing calcium sulphate or magnesium sulphate, some dissolves. The water is hard because calcium and magnesium ions in solution react with soap and stop it from forming lather readily.

TYPES OF WATER HARDNESS

Water is said to be hard if it contains calcium or magnesium ions. There are two types of water hardness.

1. Temporary hardness

This type of hardness is caused by the presence of dissolved magnesium hydrogen carbonate or calcium hydrogen carbonate in water and can be removed by boiling.

2. Permanent hardness.

This type of hardness is caused by the presence of magnesium sulphate or calcium sulphate in water which can not be decomposed by boiling hence 'permanent'.

REMOVAL OF HARDNESS

Temporary hardness of water can be removed by the following methods.

1. Boiling:

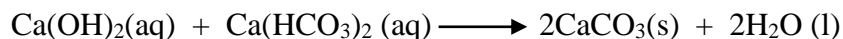
Boiling decomposes calcium hydrogen carbonate to form insoluble calcium carbonate and magnesium carbonate respectively. The insoluble carbonates are filtered off leaving soft water.



2. Addition of calcium hydroxide (lime water).

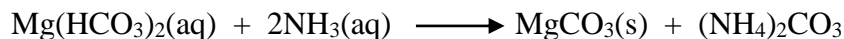
Calculated quantity of calcium hydroxide is used to remove temporary hardness of water.

It precipitates insoluble calcium carbonate hence removing the calcium ions that cause hardness.



3. Addition of aqueous Ammonia.

This precipitates out calcium or magnesium carbonate.



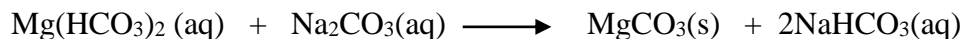
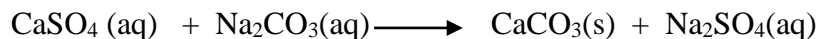
Methods of removing both temporary and permanent hardness

1. Distillation.

Distillation of water removes all hardness; water is boiled to form steam which is condensed to form pure and soft water. Solid impurities are left in the distillation flask.

2. Addition of sodium carbonate (washing soda)

This precipitates insoluble calcium carbonate or magnesium carbonate which can be filtered off.



Disadvantages of using sodium carbonate;

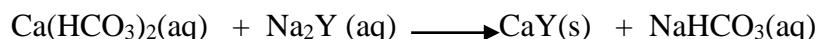
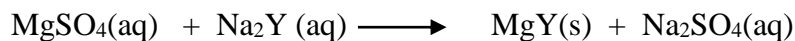
- ❖ The reaction forms insoluble carbonates which stain the clothes.

❖ Sodium carbonate hydrolyses in water to form alkali which attacks the skin.

3. By passing water through permutit (Sodium aluminium silicate, Na_2Y)

When water is passed through permutit, the Ca^{2+} or Mg^{2+} in hard water are removed and replaced with sodium ions of permutit.

This method works on the principle of ion exchange.



Experiment: To compare hardness of water

- Fill a burette with soap solution
- Pipette 25.0cm^3 of tap water into a conical flask.
- Titrate with soap solution adding 0.5cm^3 at a time. Cork the conical flask and shake the mixture vigorously. Continue adding the soap solution and shaking until a permanent lather forms on the water.
- Record the volume of soap solution used.
- Repeat the procedure with rain water
- The water which uses little soap solution to form permanent lather is soft and one which uses a lot of soap solution to form lather is hard.

Advantages of Hard water

- Rich in calcium which builds strong bones and teeth.
- It has good taste
- It prevents lead poisoning
- It builds strong egg shells

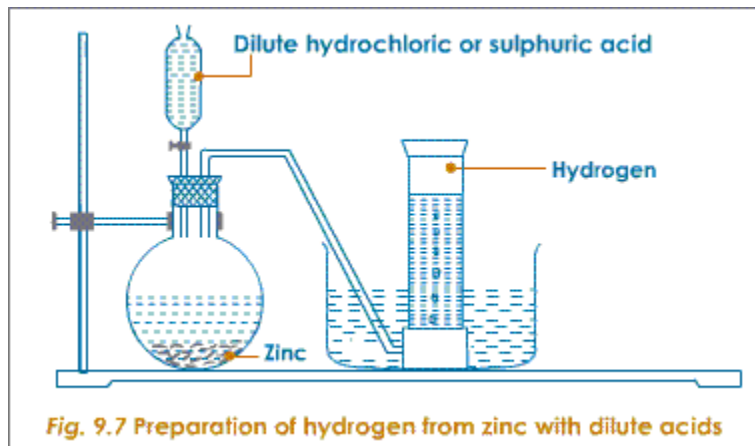
Disadvantages of Hard water

- Hard water wastes soap due to formation of scum.
- Causes dirty marks on clothes due to scum and also damage silk and nylon.
- Causes fur in kettles and pans leading to wastage of fuel. Fur is a bad conductor of heat.
- Causes wrong colours in dyeing process and poor finish during leather tanning.
- Causes formation of boiler scale which wastes fuel because it's a bad conductor of heat.

HYDROGEN

Hydrogen is the lightest gas.

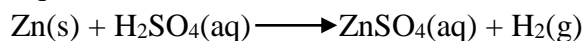
Laboratory preparation of hydrogen



Dilute sulphuric acid is added to zinc granules through the thistle funnel. Effervescence occurs and hydrogen gas is given off.

Copper (II) Sulphate solution is added as a catalyst to speed up the reaction. The gas may be dried by passing it through concentrated sulphuric acid and collected by upward delivery as it is less dense than air.

Equation:



INDUSTRIAL MANUFACTURE OF HYDROGEN

Hydrogen can be prepared on a large scale by passing steam over heated carbon.

Physical properties of Hydrogen

- ❖ It's a colourless, odourless and tasteless gas.
- ❖ It's neutral to litmus paper.
- ❖ It's less dense than air.
- ❖ It is only slightly soluble in water.
- ❖ It is the lightest substance known (air is 14.5 times denser).

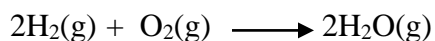
Test for hydrogen

A mixture of hydrogen and air explodes with a 'pop' when lighted.

Chemical reactions of Hydrogen

1. Combustion

Hydrogen burns quietly with a pale blue flame in air forming water.



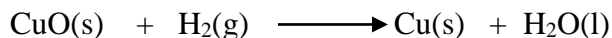
2. Reducing Action

Hydrogen removes oxygen from the metal oxides are reduced to the metals and hydrogen oxidized to water.

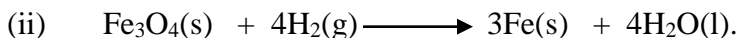
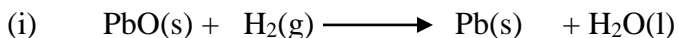
Dry hydrogen is passed over heated copper (II) oxide in a combustion tube.

Observation: The black oxide turns to brown and a colourless liquid forms on the cooler parts of the tube.

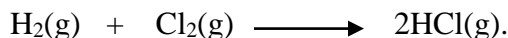
Metal oxide(s) + hydrogen (g) \longrightarrow Metal(s) + water(l)



Hydrogen also reduces copper (II) Oxide and tri iron tetra oxide.



3. Hydrogen combines readily with chlorine to form misty fumes of hydrogen chloride.



USES OF HYDROGEN

- ❖ Hardening of oils in fats to make margarine. This process is called hydrogenation.
- ❖ Manufacture of Ammonia by the Haber process.
- ❖ In oxy-hydrogen flame for cutting and welding steel.
- ❖ To fill weather balloons because of its lightness.
- ❖ Converting coal into petrol
- ❖ As a rocket fuel.

OXIDATION AND REDUCTION

Oxidation is addition of oxygen to or the removal of hydrogen from a substance. Reduction is the removal of oxygen from or addition of hydrogen to a substance.

Oxidising agent adds oxygen to another substance or removes hydrogen from another substance.

Reducing agent is a substance which adds hydrogen to another substance.

Electrons in Oxidation and Reduction

Oxidation is the removal of electrons from a substance.

Reduction is the addition of electrons to a substance.

An oxidizing agent is a substance that accepts electrons.

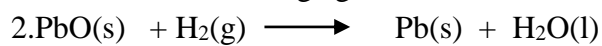
A reducing agent is a substance that supplies(donates) electrons to another substance.

Examples



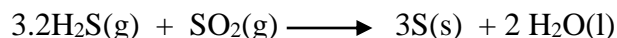
PbO is oxidizing agent

NH₃ is reducing agent



PbO is oxidizing agent

H₂ reducing agent



H₂S is reducing agent

SO₂ is oxidizing agent

REDOX REACTIONS

Redox reactions are reactions in which reduction and oxidation occur at the same time.

Non metals are categorized as electronegative elements because they have high tendency to gain electrons. Such elements act as oxidizing agents e.g chlorine, fluorine, oxygen, bromine, sulphur and nitrogen.

Metals are categorized as electropositive elements because they have the tendency to lose electrons. Such elements act as reducing agents e.g iron, zinc, aluminium, and calcium.

When a reaction occurs between a metal and a non metal to form a compound, metals act as reducing agents and non metals act as oxidizing agents.

REACTIVITY SERIES

Reactivity series refers to an arrangement of elements beginning with the most reactive and ending with the least reactive element.

The series is based on the reactions with air, water and dilute acids.

Potassium
Sodium
Calcium
Magnesium
Aluminium
Zinc
Iron
Lead
Hydrogen
Copper
Mercury
Silver
Gold

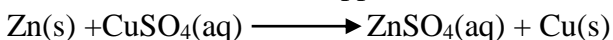
Displacement Redox reactions

Metals higher in the reactivity series displace metals low in the series from their salts in solution. Displacement occurs most readily if there is a big gap between the two metals in the series.

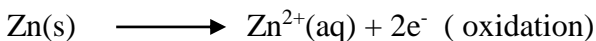
Examples.

1. Addition of zinc powder to copper (II) sulphate solution.

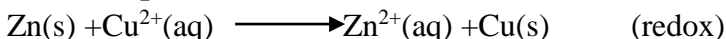
Observation: The blue copper (II) sulphate solution turns to colourless and a brown solid is formed. Zinc is higher than copper in the reactivity series and displaces copper from its solution to form brown copper and a colourless zinc sulphate solution.



Ionic equation:



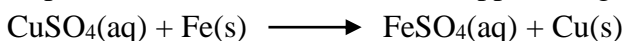
Overall equation:



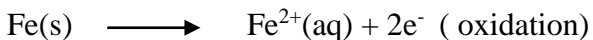
2. Addition of iron filings to copper (II) sulphate solution.

Observation: The blue copper (II) sulphate solution turns to green and a brown solid is formed.

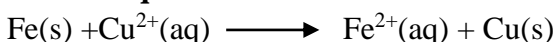
Iron is higher than copper in the reactivity series and displaces copper from blue copper (II) sulphate solution to form a brown copper and green iron (II) sulphate solution.



Ionic equation:



Overall equation



Exercise

1. (a).What do you understand by reactivity series?

(b).The following elements form part of the reactivity series: sodium, lead, magnesium, iron and calcium.

(i). Arrange the elements from the least reactive to the most reactive.

(ii).Arrange the elements from the most reactive to the least reactive.

(c).Briefly describe how the elements above react with water.Write equation where necessary.

(d).Element **P** removes oxygen from element **Q** but not from **R**.Element **S** removes oxygen from **R**.Arrange the elements from the least reactive to the most reactive.

2. (a).Define oxidation and reduction in terms of:

(i).Oxygen (ii). Hydrogen (iii). Electrons.

(b).What is a redox reaction?

(c). Zinc dust was added to copper(II) sulphate solution.

(i). State what was observed.

(ii).Write half equation for oxidation reaction

(iii).Write half equation for reduction reaction

(iv).Write overall redox equation for the reaction between zinc and copper sulphate.

3. Iron reacts with excess steam to form solid **K** and gas **L**.

(a).Name:

(i) Solid **K**

(ii) Gas **L**

(b). State what was observed during the reaction

(c). State how gas **L** could be tested

(d).A gaseous product was dried and passed over heated lead oxide in a combustion tube.

(i). State what was observed.

(ii). Write an equation for the reaction.

4. Hydrogen gas can be prepared in the laboratory from acid **X** and metal **Y**.

(a).Name the ;

(i). Acid **X** (ii) Metal **Y**

(b).Name the catalyst normally used in the preparation of hydrogen gas in the laboratory.

(c).Write the equation to show how hydrogen gas is prepared from **X** and **Y**.

(d).Dry hydrogen gas was passed over heated copper (II) oxide in a combustion tube.

(i).State what was observed in the combustion tube?

(ii).Write the equation for the reaction that took place.

(e).Give four uses of hydrogen gas.

5. (a) What is hard water?

(b).Name two types of hard water.

(c).Describe the methods of removing both temporary and permanent hardness in water.

(d).State the advantages and disadvantages of using hard water.

(e). Describe the process of water treatment.

(g).Give three uses of water to the society.

THE STRUCTURE OF THE ATOM

Atoms are the building blocks of elements. Each atom is made up of two regions.

The nucleus and the energy levels (shells).

The nucleus

An atom consists of electrons moving around a small nucleus. The nucleus of hydrogen atom consists of a single proton. The nuclei of other atoms consist of both protons and neutrons.

The number of protons vary from 1 to 92. The number of protons in the nucleus of an atom is called the atomic number.

Composition of an atom

The atom is made up of three fundamental particles, Protons, Neutrons and Electrons.

Protons:

- They are present in the nucleus of an atom .
- A proton has one positive charge, the smallest unit positive charge that can exist.

Neutrons:

- They are present in the nucleus of all atoms except hydrogen.
- They have no charge.
- They have a relative mass of 1.

The particles protons and neutrons within the nucleus of an atom are collectively called Nucleons.

Electrons:

- An electron has one negative charge, the smallest unit of charge that can exist.
- An electron has a very small mass, $\frac{1}{1840}$.

Electrons move in orbits around the nucleus, like planets move around the sun.

An atom is neutral, therefore the number of electrons (- ve) equals the number of protons (+ve) in the nucleus.

Summary of the properties of particles.

	Charge	Mass
--	--------	------

Proton	+1	1
Electron	-1	1/ 1840
Neutron	Nil	1

ATOMIC CHARACTERISTICS

1. Atomic Number

Atomic number is the number of protons in the nucleus of an atom.

2. Mass Number (Atomic Mass)

Mass number is the sum of protons and neutrons in the nucleus of an atom.

Symbols

The mass number and atomic number of an element can be written together with the symbol to completely describe the atom of an element.

The mass number is written on the top left hand corner and atomic number is written on the bottom left hand corner of the symbol.

Example



37 represent mass number

17 represent atomic number

$37 - 17 =$ number of neutrons.

ISOTOPES

Isotopes are different atoms of an element having the same number of protons but different number of neutrons therefore different mass numbers.

Examples

Each hydrogen atom has 1 proton and 1 electron.

1. Hydrogen, ${}^1_1\text{H}$, ${}^2_1\text{H}$, ${}^3_1\text{H}$

Each chlorine atom has 17 protons and 17 electrons.

2. Chlorine ${}^{35}_{17}\text{Cl}$, ${}^{37}_{17}\text{Cl}$

Each carbon atom has 6 protons and 6 electrons.

3. Carbon, ${}^{12}_6\text{C}$, ${}^{13}_6\text{C}$, ${}^{14}_6\text{C}$

Each oxygen atom has 8 protons and 8 electrons.

4. Oxygen, ${}^{16}_8\text{O}$, ${}^{17}_8\text{O}$, ${}^{18}_8\text{O}$

ISOTOPE

Isotopy is the existence of atoms of an element having the same number of protons but different mass numbers due to the difference in the number of neutrons.

Electronic structures of elements

The electrons move in definite orbits around the nucleus. The orbits which electrons occupy have definite energies.

The orbit nearer the nucleus has less energy than subsequent orbits. The further away the orbit is from the nucleus the greater the energy. Therefore, when electrons gain energy they move to higher orbits. When they lose energy they move into lower orbits.

Number of electrons in a shell

The number of electrons that can fill an energy level or shell numbered **n** is **2n²**. However, a given energy shell can have less than the maximum number of electrons but not more.

1. The first shell (orbit) can have a maximum of 2 electrons.
2. The second and third orbits, which are farther away from the nucleus, can each have a maximum of 8 electrons.
3. In the outer most shell of any atom, the maximum number of electrons is 8.

ELECTRONIC CONFIGURATION

The electronic configuration of an atom is an arrangement of electrons according to the sequence of energy levels. The electronic configuration is written by showing the number of electrons in each energy level starting with electrons in the first energy level.

Element	Atomic number	Electronic configuration
Hydrogen	1	1
Helium	2	2
Lithium	3	2:1
Beryllium	4	2:2
Boron	5	2:3
Carbon	6	2:4
Nitrogen	7	2:5
Oxygen	8	2:6
Fluorine	9	2:7
Neon	10	2:8
Sodium	11	2:8:1
Magnesium	12	2: 8:2
Aluminium	13	2:8:3
Silicon	14	2:8:4
Phosphorus	15	2:8:5
Sulphur	16	2:8:6
Chlorine	17	2:8:7
Argon	18	2:8:8

Potassium	19	2:8:8:1
Calcium	20	2:8:8:2

RELATIVE ATOMIC MASS

The relative atomic mass of an element is the mass of one atom of an element divided by the mass of one-twelfth.

$$\text{Relative atomic mass} = \frac{\text{mass of 1 atom of an element}}{\text{mass of 1 atom of carbon}} \times 12$$

Calculation of Relative atomic mass from relative abundance of isotopes of an element.

Example.

1. Natural chlorine consists of 75% chlorine – 35 and 25% chlorine – 37. Calculate the relative atomic mass of chlorine.

$$\text{Chlorine} - 35 = \frac{75 \times 35}{100} = 26.25$$

$$\text{Chlorine} - 37 = \frac{25 \times 37}{100} = 9.25$$

$$\begin{aligned} \text{Relative atomic mass of chlorine} &= 26.25 + 9.25 \\ &= 35.5. \end{aligned}$$

2. A sample of chlorine is a mixture of two isotopes, chlorine – 35 and chlorine-37 in the ratio 3:1. Calculate the Relative atomic mass of chlorine atom.

$$\text{Total ratio } 3 + 1 = 4$$

$$\text{Chlorine} - 35 = \frac{3}{4} \times 35 = 26.25$$

$$\text{Chlorine} - 37 = \frac{1}{4} \times 37 = 9.25$$

$$\begin{aligned} &= 26.25 + 9.25 \\ &35.5. \end{aligned}$$

Exercise

3. Naturally occurring Gallium consists of isotopes, Gallium – 69 and Gallium –71 in the ratio 3:2. Calculate the relative atomic mass of gallium.
4. The percentage abundances of the isotopes of lead are as follows:
Lead - 207 = 25%, Lead - 206 = 25% and lead - 205 = 50%. Calculate the relative atomic mass of lead.

5. The relative percentage abundances of the isotopes of lithium are as follows: calculate the relative atomic mass of lithium.

VALENCY AND ELECTRONIC STRUCTURE

Valency of an element can be derived from electronic structure of its atom.

1. Elements with atoms having **1-4** electrons in their outer most shells have valencies equivalent to the number of electrons in the outer shell.

Example

Element	electronic structure	valency
Na	2:8:1	1
Ca	2:8:8:2	2
Al	2:8:3	3
Si	2:8:4	4

2. Elements with atoms having **5 – 7** electrons in the outer most shell have valencies equivalent to 8 – electrons in the outer shell.

Element	electronic structure	valency
Phosphorus	2:8:5	3
Sulphur	2:8:6	2
Chlorine	2:8:7	1

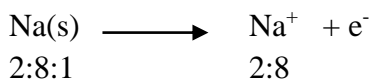
ION FORMATION

An ion is an electrically charged particle and is formed when an atom loses or gains one or more electrons.

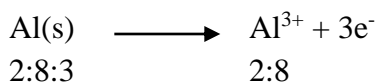
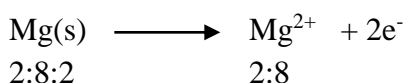
Positive ion formation

1. Metal atoms having **1-3** electrons in the outer most shells lose these electrons to form positive ions.

These ions are called cations. Cations have more protons than electrons.

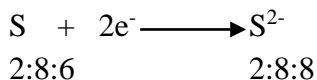
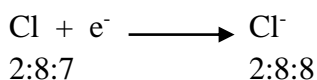
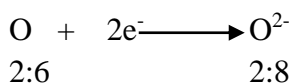


2:1 2



2. Non metals having **5-7** electrons in their outer most shell gain electrons to form negative ions. These ions are called anions. Anions have more electrons than protons.

Example



Some atoms, their electronic configurations, the formulae of the ions and electronic configurations of the ions.

Atom	Atomic number	Electronic configuration	Formula of ion	Electronic configuration of ion
Li	3	2:1	Li^+	2
B	5	2:3	B^{3+}	2
Na	11	2:8:1	Na^+	2:8
K	19	2:8:8:1	K^+	2:8:8
Mg	12	2:8:2	Mg^{2+}	2:8
N	7	2:5	N^{3-}	2:8
P	15	2:8:5	P^{3-}	2:8:8
S	16	2:8:6	S^{2-}	2:8:8
Cl	17	2:8:7	Cl^-	2:8:8
Ca	20	2:8:8:2	Ca^{2+}	2:8:8

Exercise

1. The number of protons and neutrons in atoms A, B, C and D are shown below.

Atom	No. of Protons	No. of neutrons
A	6	6
B	12	12
C	6	8
D	17	20

(a) Write electronic configuration of:-

- (i). D
- (ii). D^-
- (iii). B
- (iv). B^{2+}

(b). which of the atoms are isotopes? Give a reason for your answer

(c) Determine the mass number of B

2(a). An atom of element is represented as ${}_{15}^{31}P$

- (i). State the number of protons in atom P.
- (ii). Calculate the number of neutrons in P
- (iii). Write electronic configuration of P

(b). another atom of the same element is represented as ${}_{15}^{32}P$

State the difference between the two atoms.

3. The table below shows the number of electrons, protons and neutrons in atoms P,Q,R S and S.

Atom	Electrons	Protons	Neutrons
P	8	8	8
Q	16	16	16
R	13	13	14
S	Y	3	4

(a) Calculate

- (i) The value of Y
- (ii) the relative atomic mass

(b) Write the electronic configurations of the following atoms and ions.

- (i) P
- (ii) P^{2-}
- (iii) R
- (iv) R^{3+}
- (v) Q

(c) Write the formula of the ion of Q.

4. The table below shows the number of protons, electrons, and neutrons in particles A, B, C and D.

Particles	Protons	Electrons	Neutrons
A	6	6	6
B	9	10	10
C	12	10	12
D	11	11	12

- (a) Which one of the particle is
- (i) A positive ion
 - (ii) A negative ion
- (b) Write the electronic structure of D
- (c) State the valency of A. Give a reason for your answer.

5. The number of protons, neutrons and electrons in particles A-F are given below.

Particle	Protons	Neutrons	Electrons
A	3	4	2
B	9	10	10
C	12	12	12
D	17	18	17
E	17	20	17
F	18	22	18

- (a) Choose from the table the letters that represent:
- (i) A neutral atom of a metal
 - (ii) A neutral atom of a non metal
 - (iii) An atom of a noble gas
 - (iv) A pair of isotopes
 - (v) A cation(positive ion)
 - (vi) An anion(negative ion)
6. The number of protons, neutrons and electrons in atoms A, B, C and D are given in the table below.

Atom	Protons	Neutrons	Electrons
A	11	12	11
B	17	18	17
C	13	14	13
D	17	20	17

- (a) Which atoms are isotopes? Give a reason for your answer.
- (b) Select from the table, an atom which forms a cation with
 - (i) One positive charge.
 - (ii) Three positive charges
- (c) State the mass number of
 - (i) A
 - (ii) B

STRUCTURE AND BONDING

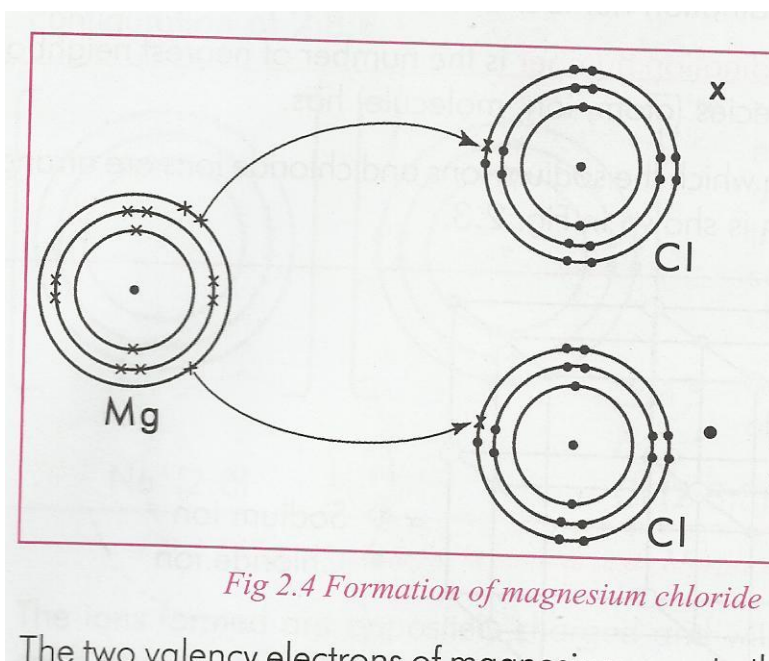
A chemical bond is a force that holds two or more atoms together in a chemical union. Electron arrangements in noble gas are very stable. They have stable octet of electrons or duplet as in Helium in the outer most energy level.

Atoms of other elements take part in chemical reactions in order to acquire the electronic structure of noble gases; they do so by transfer of electrons to form ions or by sharing of electrons.

1. THE IONIC BOND (ELECTROVALENT BOND)

Ionic bond is formed between a metal and a non metal by transfer of electrons from a metal atom to a non metal atom. The metal atom loses electrons and non metal atom gains electrons lost by the metal to acquire a stable structure.

(a). Formation of magnesium chloride ($\text{Mg} = 12$, $\text{Cl} = 17$)



Exercise

- (b). Formation of sodium chloride (Na = 12, Cl = 17)
- (c) Formation of calcium oxide (Ca=20, O = 8)
- (d) Formation of magnesium oxide
- (e). Formation of magnesium lithium oxide.

Characteristic properties of ionic compounds

- ❖ They are crystalline ionic solids at room temperature and do not vaporize easily.
 - ❖ They are soluble in water but insoluble in organic solvents
 - ❖ They have high melting and boiling points due to strong forces of attraction between the ions.
 - ❖ They conduct electricity when in solution or molten form because the ions are mobile.
- They do not conduct electricity in solid state as the ions are not free to move about.
- ❖ They consist of ions and have giant ionic structures.

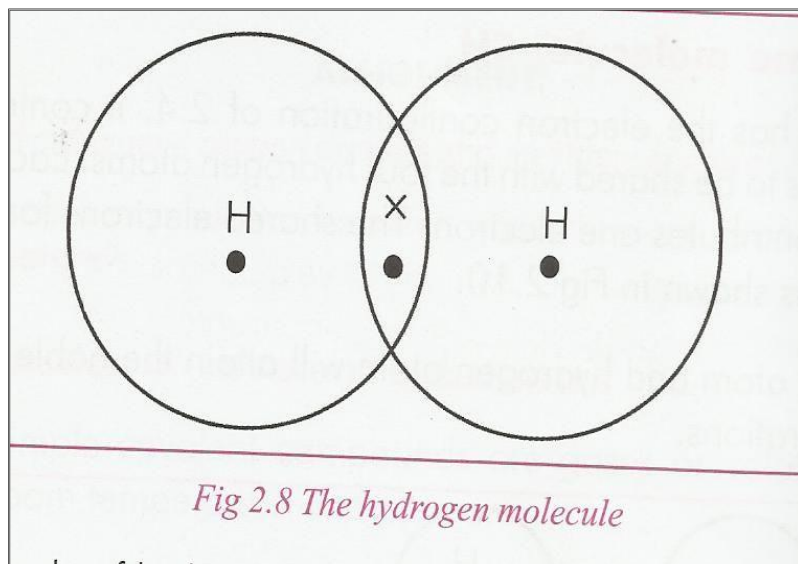
2. THE COVALENT BOND

Covalent bond occurs between non metal atoms of the same element or different elements by sharing of electrons. The electrons for sharing are contributed by each of the atoms sharing them. When electrons are shared between atoms, each atom attains an octet in the outermost energy level. A molecule is normally formed of a covalent compound.

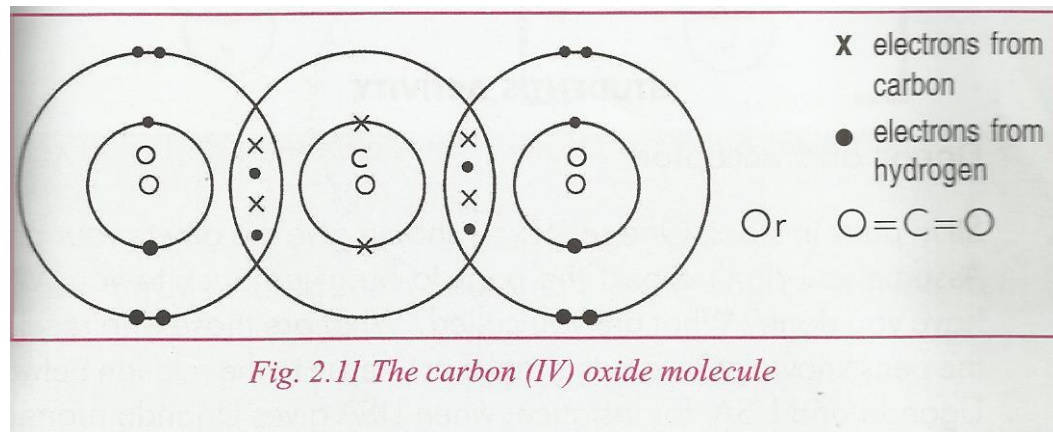
A covalent bond can be represented by a dash line (—) e.g. Cl-Cl which stands for the molecule Cl₂.

The **covalency** of an atom is the number of electrons it contributes to be shared. A covalent bond is a weak force of attraction that holds two atoms that shared their electrons. Covalent bonds can be single or multiple.

(a). Formation of Hydrogen molecule (H = 1).



(b) Formation of carbon dioxide



Exercise

(c) Formation of methane molecule, CH₄ (c = 6, H = 1)

(d). Formation of ammonia molecule ($N=7$, $H=1$)

(e). Formation of Water molecules ($H=1$, $O=8$)

(f). Formation of chlorine molecule ($Cl=17$).

Characteristic properties of covalent compounds.

- ❖ They do not conduct electricity as they contain no ions.
- ❖ They are insoluble in water but soluble in organic solvents such as benzene or carbon disulphide.
- ❖ Have low melting and boiling points.
- ❖ Occur as gases or volatile liquids at room temperature e.g. ammonia, carbon dioxide, ethanol.
This is because their molecules are electrically neutral and have little attractive force for each other.
- ❖ They consist of molecules.

3. THE METALLIC BOND

The metallic bonding occurs between metal atoms. The outermost shell electrons of each atom are easily removed with the formation of metal cations. The electrons are free to move through the lattice structure of the metal. These electrons are said to be delocalized and no longer confined to a particular atom. The metal and the 'sea' of delocalized electrons will bond the atoms together. This type of bonding is very strong in some metals such as iron which are difficult to shatter, but much weaker in potassium and sodium, which can be cut with a knife because they are soft.

Properties of metals.

- ❖ Generally solids and have high melting and boiling points except mercury which is a liquid.
- ❖ Good conductor of electricity and heat because the mobile electrons can carry an electric current.
- ❖ They are ductile
- ❖ They are malleable and have shiny surfaces.

THE CO-ORDINATE OR DATIVE BOND

A coordinate bond is a covalent bond in which the shared pair of electrons is provided by only one of the bonded atoms. The atom that provides both electrons is called the *donor* and the other atom is called the *acceptor*.

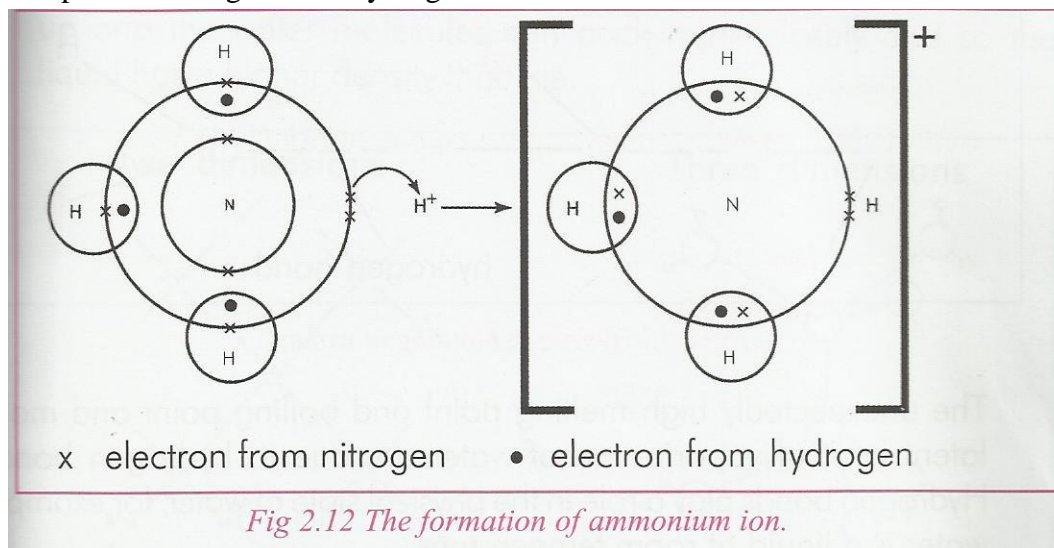
The donor must have an unshared pair of electrons (lone pair of electrons) in its outermost shell and the acceptor should have an empty energy level. A lone pair of electrons is a pair that is not directly concerned in its existing valency bonds.

Formation of Ammonium ion

The Nitrogen atom in ammonia has lone pair of electrons. It can donate the lone pair of electrons to an acceptor species.

The bonding supplies an electron duplet to hydrogen nucleus while maintaining the electron octet of Nitrogen atom.

The positive charge on the hydrogen atom is distributed over the whole ammonium ion.



STRUCTURES OF SOLIDS

Solid substances have 3-dimensional shapes called structures. Structures are made of many small building units (atoms, ions or molecules) joined together.

Solid substances have structures that are either crystalline with definite shapes or amorphous. The crystalline solids have two broad structures.

1. Simple molecular structures

Molecular structures are formed mainly by covalent compounds. They consist of small molecules containing a few atoms. The molecules are held together by weak intermolecular forces of attraction called Vander Waals forces.

Examples of simple molecules include.

❖ Iodine, Solid carbondioxide (dry ice), Water, Ammonia, Phosphorus, ethanol and sulphur

Properties of simple molecules

- ❖ They have low melting and boiling points
- ❖ They do not conduct electricity in either molten or in solution.
- ❖ They are insoluble in water but soluble in organic solvents.

2. **Giant structures**

(i). Giant ionic structures

These solids consist of cations and anions held in a regular repeating pattern by strong ionic bonds due to electrostatic forces of attraction between oppositely charged ions.

Examples

- ❖ Sodium chloride, Magnesium chloride, Aluminium oxide.

Properties of ionic crystals.

- ❖ Hard and brittle.
- ❖ High melting and boiling points.
- ❖ Soluble in water and insoluble in organic solvents.
- ❖ Good conductors of electricity in molten or aqueous solution.

(ii). Giant atomic (Giant molecular) structure

In this structure, the atoms of a non metal are joined by a continuous net work of covalent bonds.

Examples

Graphite, Diamond, Silicon dioxide.

Properties of Giant atomic structures.

- ❖ Hard solids with very high melting points.
- ❖ Insoluble in common solvents.

(iii). Giant metallic structure

These are formed by metals. It's made up of millions of metal atoms. The delocalized (mobile) electrons hold the inner positive core of the metal atoms through electrostatic attraction. The atoms are closely packed and the degree of packing and the final shape are determined by the size of the metal atoms.

Properties of Giant metallic structures.

- ❖ Occur as hard solids at room temperature except mercury.
- ❖ Have high melting and boiling points.
- ❖ Malleable and ductile i.e. can be hammered without breaking and can be drawn into shapes
- ❖ Have a shiny appearance when freshly cut.
- ❖ Good conductors of electricity due to free mobile electrons.

Exercise

1. The table below shows the properties of substances A, B, C and D. Study it and answer the following questions.

			Electrical conductivity
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Substance	Melting point($^{\circ}\text{C}$)	Boiling point($^{\circ}\text{C}$)	As solid	As a liquid
A	1083	2567	Good	Good
B	-182	-164	Poor	Poor
C	1723	2230	Poor	Good
D	993	1695	poor	Poor

State with a reason which one of the substances

- (a) Is ionic
- (b) Is metallic
- (c) Has giant covalent structure.

2. Use the table below to answer the following questions.

Substance	Density	Melting point($^{\circ}\text{C}$)	Electrical conductivity
A	10.5	961	Good
B	2.1	119	Poor
C	2.2	800	Poor
D	2.2	3700	good

- (i) Which substance is likely to be a metal?
- (ii) Which substance is likely to have a giant covalent structure?
- (iii) Which substance is likely to be molecular?

3. (a).Name the structures adopted by the following substances;

- (i) Water (ii). Sodium chloride (iii). Diamond (iv). Magnesium

(b).Show how bonding takes place in the following:[C=6,O=8,H=1,Mg=12,Na=11,n=7]

- (i).Ammonia (ii).Carbon dioxide (iii).Methane (iv).Magnesium chloride
- (v). Sodium chloride (vi).Water. (vii). Ammonium ion.

(c).Define the term covalency

(d).State four physical properties of covalent compounds.

4. The table below provides information about some substances. Study the table and answer the questions that follow by writing the letters A-H.

Substance	Melting point($^{\circ}\text{C}$)	Boiling point($^{\circ}\text{C}$)	Electrical conductivity		
			of solid	of liquid	In water
A	1540	3000	Good	Good	Insoluble
B	-114	-85	Poor	Poor	Good
C	712	1418	Poor	Good	Good
D	-68	57	Poor	Poor	Good
E	-25	144	Poor	Poor	Insoluble

F	-39	357	Good	Good	Insoluble
G	1700	2776	Poor	Poor	Insoluble
H	2045	3000	poor	Good	Insoluble

Name:

- (a) The substance with the lowest melting point.
 - (b) The substance which is a gas at room temperature.
 - (c) One substance which is a liquid at room temperature.
 - (d) One substance which could be a metal.
 - (e) One substance which is likely to be an ionic solid at room temperature.
5. (a). Define the term bonding.
- (b). Give four types of bonding and briefly state how each type occurs.
- (c). State the type of bond that occurs between the atoms of
- (i). Chlorine
 - (ii). Aluminium
 - (iii). Calcium and oxygen
 - (iv). Carbon in graphite
 - (v). carbon in diamond
- (d). Which of the compounds in (c) have
- (i). Giant ionic structure
 - (ii). Simple molecular structure
 - (iii). Giant atomic structure
 - (iv). Giant metallic
- (e). State 3 properties of
- (i). Ionic structures
 - (ii). Metallic structures
 - (iii). Giant atomic structures

THE PERIODIC TABLE

This is a table of elements arranged in order of increasing atomic number. The elements which have similar chemical properties are in the same columns called groups. The group number is written in Roman numerals. The horizontal rows of the elements are called periods. The period number is written in Arabic numeral.

I							VII	VIII
	II		III	IV	V	VI		
		Transition						
		elements						

- ❖ There are 8 groups in the periodic table numbered I – VIII. Group 8 is also known as group **O**.
- ❖ Elements of the same group have similar chemical properties because they have the same number of outer most energy level electrons (valency electrons).
- ❖ Atomic size (radius) increases down the group as more energy levels are added.
- ❖ Elements in the same group have a common valency.
- ❖ The group number is equal to the number of electrons in the outer most energy level of the element.

Example.

Element	electron structure	Group
Li	2:1	1
Mg	2:8:2	2
Al	2:8:3	3
C	2:4	4
N	2:5	5
O	2:6	6
F	2:7	7
Ne	2:8	8

- ❖ Elements in group 4 have a valency 4 because their atoms require 4 electrons to attain a noble gas structure.
- ❖ Elements in groups 5-7 react by gaining electrons to form negative ions. They are non metals. The number of electrons gained is equal to the valency of the element.

Group	No. of electrons gained	Valency
5	3	3
6	2	2
7	1	1

Periods

There are 7 periods in the periodic table. Elements in the same period have the same number of energy levels.

Period **1** contains 2 elements, hydrogen and helium.

Period **2** and **3** are short periods and contain 8 elements.

Periods **4** and **5** are long periods and contain 18 elements each.

Characteristics of a period

- ❖ Elements in the same period have the same number of energy levels, occupied by electrons and give the period number.
- ❖ The atomic radius (size) decreases across a period from left to right.
- ❖ Elements in the same period have different chemical properties.
- ❖ Elements in the same period have different valencies.

CHEMICAL FAMILIES

This refers to elements with the same chemical properties i.e. elements in the same group of the periodic table with the same number of electrons in the outer most shell.

GROUP 1: THE ALKALI METALS

All group **1** elements have one electron in the outer most shell. They are monovalent. They are called alkali metals because they react with water to form an alkaline solution.

Element	Electronic structure.
Li	2:1
Na	2:8:1
K	2:8:8:1

Physical properties of Alkali metals

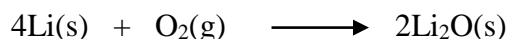
- ❖ Have low melting and boiling points because each metal contributes only one electron to the sea of electrons hence weak metallic bonding.
- ❖ They are soft and shows shiny surface when freshly cut.
- ❖ Conduct electricity due to mobile electrons.
- ❖ Have low densities. Their density is less than that of water so they float on water.

Group I elements are the most reactive metals.

REACTIONS OF ALKALI METALS WITH AIR, WATER AND CHLORINE

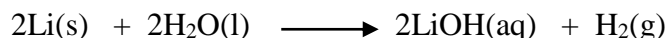
(a) Lithium

Lithium burns in oxygen with a crimson flame (intensive white) to form a white solid, lithium oxide.



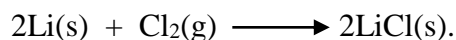
Reaction with water.

Lithium floats on water surface, gradually dissolving producing hydrogen gas. An alkaline solution is formed.



Reaction with chlorine

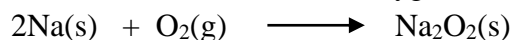
Lithium reacts instantly with chlorine to form lithium chloride.



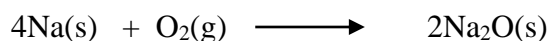
(b) Sodium

Reaction with air

Sodium melts and burns in oxygen with a yellow flame to form yellow sodium peroxide

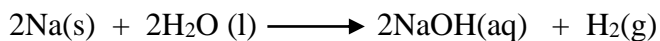


In limited oxygen it burns with a yellow flame to form a white solid, sodium oxide.



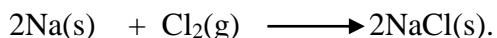
Reaction with water.

It darts around on water surface with hissing sound producing hydrogen gas. Floats and melts to a silver ball as it dissolves.



Reaction with chlorine

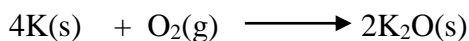
Sodium burns in chlorine with a bright yellow flame to form white clouds of sodium chloride.



(c) Potassium

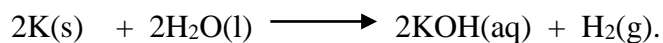
Reaction with air

Potassium burns in air with lilac (purple) flame to form potassium oxide



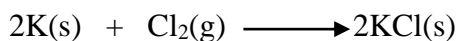
Reaction with water

Darts around the water surface. Floats and melts bursting into purple flame.



Reaction with chlorine

Potassium reacts with chlorine to form potassium chloride.



Storage of group I metals

Group 1 metals are very reactive and are stored under oil to prevent them from coming into contact with air and water.

Uses of sodium

- ❖ Manufacture of sodium hydroxide used in paper industry.
- ❖ In glass industry as sodium carbonate.
- ❖ Sodium vapour lamps are used for street lighting.
- ❖ Produce sodium chloride, a food additive.
- ❖ Molten sodium is used as a coolant in some nuclear reactors.

THE ALKALINE – EARTH METALS (Group II elements)

All group II elements have 2 electrons in the outermost shell. They are divalent.

Element	electronic structure
Be	2:2
Mg	2:8:2
Ca	2:8:8:2

As the group is descended, the size of the atom increases as more energy levels are added.

Physical properties of group 2 metals.

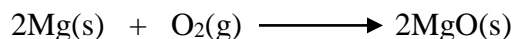
- ❖ High melting and boiling points because of strong metallic bond.
- ❖ High densities because of small size of atoms that allows close packing
- ❖ Good conductors of heat and electricity since they have mobile electrons.
- ❖ They are harder than the alkali metals.
- ❖ They are silver-grey in colour when pure and clean.

CHEMICAL REACTIONS OF ALKALINE EARTH METALS

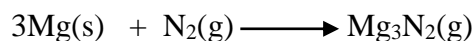
(a) Magnesium

Reaction with air

Magnesium burns in air with a bright white flame to form a white powder of magnesium oxide.

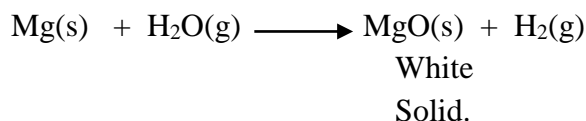


In limited oxygen, magnesium reacts with nitrogen of air to form a white powder of magnesium nitride.



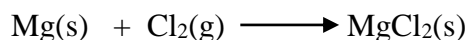
Reaction with water

Magnesium reacts with steam vigorously and slowly with cold water. Magnesium ribbon should be heated.



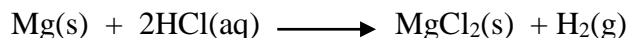
Reaction with Chlorine

Magnesium burns brightly in chlorine to form a white solid of magnesium chloride.

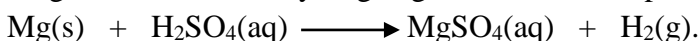


Reaction with dilute hydrochloric acid and Sulphuric acids.

(i). Magnesium reacts with dilute hydrochloric acid giving off hydrogen gas.



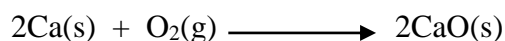
(ii). Magnesium liberates hydrogen gas with dilute sulphuric acid



(a) Calcium

Reaction with air

Calcium burns in air with a red flame when heated and forms a white solid of calcium oxide



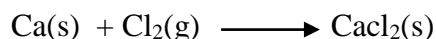
Reaction with water

Calcium sinks to the bottom as it's denser than water. Effervescence occurs and hydrogen gas is given off. A cloudy solution of calcium hydroxide is formed.



Reaction with chlorine

Calcium reacts with chlorine to form calcium chloride.



Uses of alkaline earth metals

- ❖ Beryllium is used in making parts of high speed air craft and in missiles.
- ❖ It's used in nuclear reactors and communication satellites.
- ❖ Calcium hydroxide is used as lime in building industry and road construction.
- ❖ Calcium is used in biological process for the formation of strong bones and teeth.
- ❖ Magnesium is used in photographic flash lights and fire works.
- ❖ It's also used in making magnesium hydroxide which is anti-acid medicine.

HALOGENS – GROUP VII ELEMENTS

The word halogen means ‘salt producer’, since all halogen elements react with most metals to form electrovalent, salt-like compounds.

All members of this group have 7 electrons in their outermost shells. They have one electron short of the noble gas structure.

Element	Electron structure	Formula	Colour
F	2:7	F ₂	Pale yellow gas
Cl	2:8:7	Cl ₂	Greenish yellow gas
Br	-	Br ₂	Dark brown solid
I	-	I ₂	Black crystalline solid

Physical properties of Halogens.

- ❖ Halogens exist as diatomic molecules in gas, liquid or solid state i.e F₂, Cl₂, Br₂ and I₂
- ❖ The size of the molecules increase down the group due to increasing size of the halogen atom.
- ❖ Melting point and boiling points of halogens increase down the group due to strong forces of attraction down the group.
- ❖ They are poor conductors of electricity and heat.

Halogens are the most reactive non metals. Their atomic radii increase down the group due to the effect of increasing number of energy levels increasing nuclear charge.

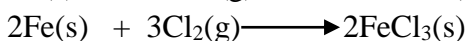
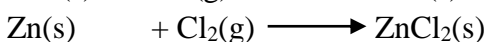
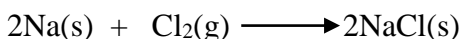
Chemical reactions of Halogens.

(a). Reaction with metals

Most heated metals combine with halogens to form metal halides.

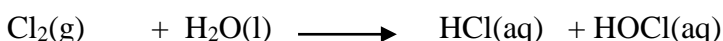


Chlorine is more reactive than bromine while iodine is the least reactive.



(b). Reaction with water

Chlorine dissolves in water to form hydrochloric acid and Hypochlorous acid.



(c) Reaction with sodium hydroxide solution

Cold dilute sodium hydroxide reacts with chlorine to form a pale yellow solution.



USES OF HALOGENS

Chlorine

- ❖ As a domestic bleach and as disinfectant in swimming pools.
- ❖ Manufacture of hydrochloric acid.
- ❖ In paper and textile industry.
- ❖ Manufacture of weed killer iodine.

Iodine

- ❖ Tincture of iodine used for medical purposes.
- ❖ Making of antibacterial agents.

Bromine

- ❖ In photographic industry as silver bromide.
- ❖ Used as petrol additives used in anti-knock petrol.

NOBLE GASES – GROUP VIII (O) ELEMENTS

They are called noble gases because they do not combine with themselves or with other elements. They are the least reactive non metals. This is because of completely filled outermost energy levels which make them stable.

He	2	2
Ne	10	2:8
Ar	18	2:8:8

USES OF NOBLE GASES

- ❖ Helium is a very light gas and is used in inflating airships, weather balloons and aeroplane tyres
- ❖ Neon is used in neon lights for advertisement.
- ❖ Argon is used in electric light bulbs. It stops oxidation of the hot filament that can darken the glass bulbs.

Transition metals

These are metals found between group II and group III e.g. Zinc, copper iron chromium, nickel.

Characteristics of transition metals

- ❖ They have high densities, high melting and boiling points.
- ❖ Their valencies are variable i.e they have more than one valency.
- ❖ They form coloured compounds.
- ❖ They behave as catalysts.

Exercise

- The electronic configuration of an element T is 2:8:6
 - Write the formula of the most common ion of T.
 - To which period of the periodic table does T belong?
 - Name an element that belongs to the same group as T.
- Elements W, X, Y, and Z belong to the same period of the periodic table. The formula of the oxide of W is W_2O and it is strongly basic forms a liquid, covalent chloride, $XC l_3$, the formula of the oxide of Y is Y_2O_3 and Z forms an ion Z^- .
 - Write equation for the reaction between the oxide, W_2O and water.
 - Write the formula of the sulphate of W.
 - Write the formula of the oxide of X.
 - Write the formula of the compound formed between W and Z.
- Part of the periodic table is shown below.

I							VIII
	II	III	IV	V	VI	VII	
			W		V		Z
	Y	T				Q	

- State:
 - The most reactive metal
 - The most reactive non metal
 - The atom that forms the largest anion
 - The most non reactive element
 - Write the formulae of the compounds formed between the following pairs of elements and in each case state the type of bonding.
 - W and Q
 - T and V
- Part of the periodic table is shown below. The letters used are not the correct symbols of the elements.

I								VIII
	I	II	III	IV	V	VI	VII	
				P		T	R	

ACIDS, BASES AND SALTS

ACIDS

An acid is a compound which when dissolved in water forms hydrogen ions as the only positively charged ions. It is the hydrogen ions that cause acidic properties and these are formed in the presence of water. Hydrogen ion is a proton because a hydrogen atom consists of a single proton and a single electron. The electron is lost when hydrogen ion is formed.

An acid turns blue litmus red and contains hydrogen ions which may be replaced, directly or indirectly, by a metal.

Examples of acids.

1. Mineral acids (common laboratory acids)

- Hydrochloric acid
- Sulphuric acid
- Nitric acid

The three acids are commonly used in the laboratory. They are sometimes called mineral acids because they were first prepared from certain minerals.

2. Organic acids

- Methanoic acid
- Ethanoic acid
- Citric acid

Common acids in every day life include: citric acid in lemons, tannic acid in tea, lactic acid in sour milk, tartaric acid in baking powder.

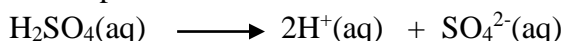
STRENGTH OF ACIDS

1. Strong acid

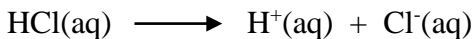
A strong acid is one that is completely ionized when in solution.

Examples

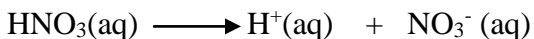
- Sulphuric acid



- Hydrochloric acid



- Nitric acid

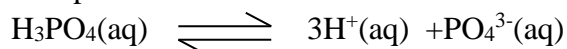


2. Weak acid

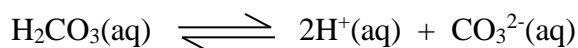
A weak acid is one which is only slightly ionized in dilute solutions and exist as molecules. Their ionization is reversible.

Examples.

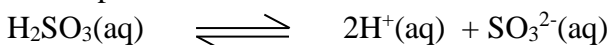
- Phosphoric acid



- Carbonic acid



- Sulphurous acid



BASICITY OF AN ACID

Basicity of an acid is the number of hydrogen ions that can be formed when one mole of acid ionizes completely. Sulphuric acid is dibasic and phosphoric acid is tribasic.

Acid	Basicity	
HCl	1	Monobasic
HNO ₃	1	Monobasic
H ₂ SO ₄	2	dibasic
H ₃ PO ₄	3	Tribasic.

TEST FOR ACIDS

Acids are tested using substances called indicators. An indicator is a substance that has different colours in acid and alkaline solutions.

Indicator	colour in acid	colour in alkaline
Litmus	red	blue
Methyl Orange	red	yellow
Phenolphthalein	Colourless	purple

NB: Weak acids turn litmus pink.

INDICATORS

Indicators are substances which change colour and indicate the degree of acidity or alkalinity of solutions. Some simple indicators can be made from coloured substances which occur naturally.

Extraction of simple indicator

- ❖ Take some coloured plant material e.g. rose petals or red cabbage leaves.
- ❖ Crush the material with a mortar and add it to a small beaker.
- ❖ Cover the material with a mixture of water and ethanol (5:1 by volume). Warm the material and solvent by placing in a larger beaker containing hot water.
- ❖ Stir the hot mixture well for 10 minutes. Filter and collect the coloured filtrate.
- ❖ Add the made indicator to various acidic solutions, water and alkaline solutions.

Indicators and pH scale

Universal indicator solutions are the most useful indicator. It shows a series of colour changes as the acidity or alkalinity of a solution changes.

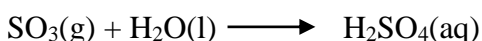
Add 2 or 3 drops of universal indicator to test the various solutions.

The pH scale is a scale of numbers from 0 to 14 used to express the degree of acidity or alkalinity of a substance. The pH number of water and other neutral substances is 7. Acidic solutions have a pH number less than 7, as the number decreases acidity increases. Alkaline solutions have a pH number of 8-14. The increase in number indicates increasing alkalinity.

pH	0 to 4	5	6	7	8	9	10 to 14
Colour	Red	Orange	Yellow	Green	Blue	Indigo	violet

PREPARATION OF ACIDS

1. By action of water on acidic oxides of non metals e.g sulphur dioxide, sulphur trioxide and carbondioxide.



2. Synthesis (Direct combination of elements).

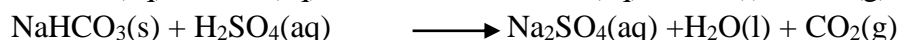
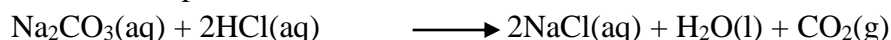
PROPERTIES OF ACIDS

1. Acids have sour sharp taste.
2. Acids turn blue litmus red. Sulphurous and carbonic acids are so weak that they turn litmus pale red or pink.
3. Acids are corrosive when concentrated and cause itching when dilute.
4. Acids react with carbonates and hydrogen carbonates liberating carbondioxide gas.

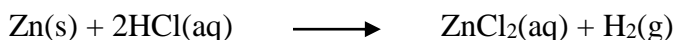
Carbonate + Acid \longrightarrow salt + water + carbondioxide

Hydrogen carbonate + Acid \longrightarrow salt + water + CO_2

Examples



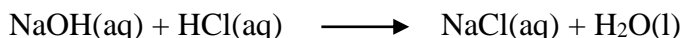
5. Acids react with metals which are more reactive than hydrogen to liberate hydrogen gas and form salt.



6. Acids react with bases (basic oxides and hydroxides) to form salt and water only.



Examples



7. Acids react with sulphites readily to form bubbles of colourless sulphur dioxide, which has a pungent, irritating smell.



USES OF ACIDS

1. Sulphuric acid is used in car batteries.
2. Nitric acid is used in the manufacture of dyes, explosives and fertilizers.
3. Hydrochloric acid is used in pickling metals.
4. Ethanoic acid is used in seasoning of food salads.

BASES AND ALKALIS

BASE:

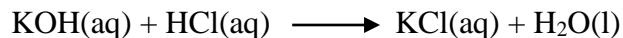
A base is a compound which contains oxide (O^{2-}) or hydroxide (OH^-) ions and reacts with an acid to form a salt and water only. Bases are oxides and hydroxides of metals. They neutralize acids forming a salt and water only. The reaction between bases and acids are called neutralizations.

Neutralization is a reaction between an acid and a base to form a salt and water only.

Examples of bases:

1. Calcium oxide (CaO)
2. Copper (II) oxide (CuO)
3. Magnesium hydroxide $\text{Mg}(\text{OH})_2$
4. Sodium hydroxide (NaOH)
5. Potassium hydroxide (KOH)

Acid + base \longrightarrow Salt + water



AN ALKALI

An alkali is a compound which when dissolved in water produces hydroxide ions (OH^-) as the only negatively charged ions. Soluble bases are called alkalis. Therefore, all alkalis are bases, but not all bases are alkalis.

An alkali is a base which is soluble in water. An alkaline solution is formed when a base dissolves in water.

Examples

1. Sodium hydroxide (NaOH)
2. Potassium hydroxide (KOH)
3. Calcium hydroxide [Ca(OH)_2]
4. Aqueous ammonia (NH_4OH)

STRENGTH OF ALKALIS

1. Strong alkalis

A strong alkali is one which is completely ionized in solution.

Examples.

Sodium hydroxide, potassium hydroxide and calcium hydroxide.

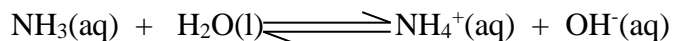
Strong alkalis are very good conductors of electricity.

2. Weak alkalis

A weak alkali is one which ionizes only partially in solution. It's characterized by reversible ionization. A weak alkali has low conductivity of electricity.

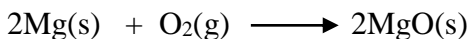
Example

Ammonia solution

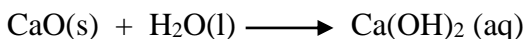


PREPARATION OF BASES

1. By burning metals in air



2. Addition of water to metal or metal oxide



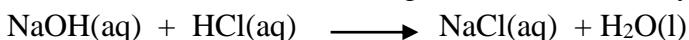
3. Heating the hydroxide, carbonate or nitrates of metals except sodium and potassium.



PROPERTIES OF BASES AND ALKALIS

1. Alkalis feel soapy and taste bitter.
2. Alkalis turn red litmus paper blue.
3. Alkalis react with ammonium salts when heated to form ammonia gas, which has choking smell and turn red litmus blue.

4. Alkalis react with acids forming a salt and water only.



5. Alkalis precipitate many insoluble hydroxides from a solution of their salts.



6. Sodium hydroxide and potassium hydroxide are deliquescent.

Sodium hydroxide solution absorbs carbon dioxide from air to form white powder of sodium carbonate.



USES OF ALKALIS AND BASES

1. Sodium hydroxide and potassium hydroxide are used in the manufacture of soap.
2. Calcium oxide is used:
 - (i). In building industry as slaked lime, in mortar and plaster.
 - (ii). In the manufacturing of bleaching powder.
 - (iii). As a drying agent to dry ammonia and ethanol.
3. Magnesium hydroxide is used as antacid to neutralize acidity in the stomach.

Exercise

1. Giving examples in each case write equations for the reaction of sulphuric acid with:

- (a). a metal
- (b). a metal oxide
- (c). a metal hydroxide
- (d). a metal carbonate
- (e). a metal hydrogen carbonate.

2.(a).What is an acid?

(b).State three physical properties of acids.

(c).Write the name and formula of the 3common mineral acids used in the laboratory.

(d).Write chemical equations to iilustrate the properties of acids using hydrochloric acid and the following compounds;

(i).Copper(II) oxide

(ii).Magnesium metal

(iii).Potassium carbonate solution

(iv).potassium hydrogen carbonate solution.

3.(a).Define the terms;

(i).Base

(ii).Alkali

(iii).Neutralization reaction

(b).Give two examples of;

(i).Bases

(ii).Alkalis

4.(a).What is an indicator?

(b).Describe an experiment to prepare an indicator in the laboratory from either coloured flower petals or red cabbage leaves.

(c).The table below shows the PH values for some aqueous solutions.

Solutions	R	S	T	U	V	W
PH	10.5	6.5	2.0	7.0	7.5	11.0

Which solution(s):

(i). is the most acidic?

(ii).is the most alkaline/

(iii).is likely to be lemon juice/

(iv).is pure water?

(d).Name the acid present in:

(i).sour milk

(ii).lemons

(iii).tea

(iv).baking powder

SALTS

A salt is the substance formed when either all or part of ionisable hydrogen ions of an acid is replaced by a metallic ion or ammonium ion. A salt is formed when an acid reacts with base.

TYPES OF SLATS

1. Normal salts

Normal salt is a salt formed when all the ionisable hydrogen ions of an acid are replaced by a metallic ion or ammonium ion.

Examples of normal salts

- ❖ Sodium sulphate
- ❖ Potassium nitrate
- ❖ Calcium chloride

In general normal salts include chlorides, sulphates, sulphites, nitrates, phosphates and carbonates of metals and ammonium radical.

2. Acid salts

Acid salt is a salt formed when ionisable hydrogen ions of an acid are partially replaced by a metallic ion or ammonium ion.

Examples

- ❖ Potassium hydrogen carbonate (KHCO_3)
- ❖ Sodium hydrogen sulphate (NaHSO_4)
- ❖ Sodium hydrogen sulphite (NaHSO_3)

In general acid salts include:-

Hydrogen carbonates, hydrogen phosphates of metals and ammonium radical.

Acid salts are formed when dibasic or tribasic acids are reacted with metal ions or ammonium ions.

3. Basic salts.

A basic salt is a compound of the normal salt and a base from which its derived e.g basic zinc chloride, $\text{Zn}(\text{OH})\text{Cl}$, basic magnesium chloride, $\text{Mg}(\text{OH})\text{Cl}$.

4. Double salts

A double salt is a combination of two normal salts. They consist of two different positive ions and one common negative ion. Most double salts contain water of crystallization loosely attached to the cations in the salt.

Examples of double salts.

- ❖ Potassium aluminium sulphate, $\text{K}_2\text{SO}_4.\text{Al}_2(\text{SO}_4)_3.24\text{H}_2\text{O}$
- ❖ Potassium magnesium chloride, $\text{KCl.MgCl}_2.6\text{H}_2\text{O}$.
- ❖ Ferrous ammonium sulphate, $(\text{NH}_4)_2\text{SO}_4.\text{FeSO}_4.6\text{H}_2\text{O}$.

HYDRATED SALTS

Hydrated salts are salts which contain water of crystallization. The water molecules are combined with the crystals. The number of water molecules varies depending on the salt. A crystalline compound containing water is hydrated and the water is sometimes called water of hydration.

Examples of hydrated salts.

Copper sulphate crystals, $\text{CuSO}_4.5\text{H}_2\text{O}$.

Iron(II) sulphate crystals, $\text{FeSO}_4.7\text{H}_2\text{O}$.

Magnesium sulphate crystals, $\text{MgSO}_4.7\text{H}_2\text{O}$.

Sodium sulphate crystals, $\text{Na}_2\text{SO}_4.10\text{H}_2\text{O}$

Sodium carbonate crystals, $\text{Na}_2\text{CO}_3.10\text{H}_2\text{O}$

Anhydrous salts are salts which do not contain water of crystallization. Anhydrous means 'without water'.

Examples of anhydrous salts

Sodium chloride

Potassium chloride

Potassium nitrate

Ammonium chloride

Potassium manganate.

PREPARATION OF SALTS

Salts are prepared basing on their solubility in water. A salt is either soluble when it dissolves in water or insoluble when it does not dissolve in water.

Solubility of salts.

Some salts can dissolve in water readily. Other salts are insoluble in water.

The solubility of a salt is the amount of the salt that is required to saturate 100g of water at a particular temperature.

Soluble

All nitrates

All chlorides except

All sulphates except

All carbonates of sodium, potassium and ammonium

insoluble

-

silver and Lead (II) Chloride

Lead (II) sulphate and barium sulphate.

all other carbonates.

Oxides of potassium, sodium

Calcium and ammonium

all other oxides.

Note: Oxides are not salts but also classified as soluble and insoluble.

Oxides are used in the preparation of salts. Soluble salts are prepared in solution and crystallized from saturated solution. Soluble salts are prepared by the action of an acid on:

- (i). a metal
- (ii). A hydroxide of a metal
- (iii). An oxide of a metal (basic oxides)
- (iv). A carbonate of a metal

The kind of salt to be prepared will depend on the kind of acid used.

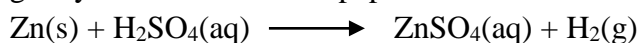
Name of acid	Salt formed	
	Normal salt	Acid salt
Hydrochloric acid	Chloride	None
Nitric acid	Nitrate	None
Carbonic acid	Carbonate	Hydrogen carbonate
Sulphurous acid	Sulphite	Hydrogen sulphite
Sulphuric acid	Sulphate	Hydrogen sulphate
Phosphoric acid	Phosphate	Dihydrogen phosphate

(a). **Action of acid on a metal.**

Preparation of Zinc sulphate crystals from zinc metal.

- ❖ Pour dilute sulphuric acid in a beaker
- ❖ Add Zinc granules until in excess and add copper (II) sulphate solution to speed up the reaction. Effervescence occurs.
- ❖ When the reaction stops, filter the excess Zinc. Evaporate the filtrate by heating in evaporating dish until crystallization point. Dip a glass rod in the solution and check if crystals appear in it or pour a little hot solution in a test tube and cool under tap shaking all the time.

- ❖ Wash the crystals 2-3 times with little cold distilled water and dry them by pressing gently between two filter papers.



(b). Action of acid on alkali

Preparation of sodium chloride from sodium hydroxide by neutralization.

- ❖ Pipette 25cm³ of sodium hydroxide into a conical flask and add 2-3 drops of phenolphthalein indicator.
- ❖ Fill the burette with hydrochloric acid solution .
- ❖ Run the acid from burette into sodium hydroxide in a conical flask. Shake constantly until the solution turns colourless.
- ❖ Transfer the solution into evaporating dish and evaporate until its concentrated.
- ❖ Pour the concentrated solution in a beaker and allow to cool and crystallize.
- ❖ Wash the crystals with little distilled water and dry between filter paper.

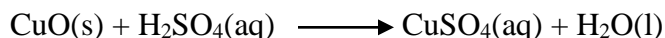
Sodium sulphate can be prepared in the same way.



(c) Action of acid on insoluble basic oxide

Preparation of copper (II) sulphate from copper (II) oxide.

- ❖ Put dilute sulphuric acid into a beaker and warm gently.
- ❖ Add copper (II) oxide to the warm acid little at a time .Continue to add the oxide until no more will react, showing all the acid has been neutralized.
- ❖ Filter off excess copper (II) oxide and collect a blue copper (II) sulphate filtrate.
- ❖ Evaporate the filtrate until crystallization point. Dip a glass rod into the hot solution to check if crystals form on it. If crystals form, stop heating and allow it to cool and crystallize.
- ❖ Wash the crystals with little distilled cold water and dry them between filter papers.



(d). Action of acid on insoluble carbonate.

Preparation of Lead (II) nitrate crystals from lead carbonate.

- ❖ Pour dilute nitric acid in a glass beaker
- ❖ Add Lead (II) carbonate little at a time while stirring, effervescence occurs.Continue to add until no more will react showing that all the acid has been neutralized.
- ❖ Filter off excess lead carbonate and collect a colourless filtrate of Lead (II) nitrate.
- ❖ Evaporate the filtrate until crystallization point. Dip a glass rod into the hot solution to check if crystals form on it. If crystals form, stop heating and leave to cool and crystallize.
- ❖ Wash the crystals with little cold distilled water and dry between filter papers.

Preparation of insoluble salts.

Insoluble salts are prepared by a method called double decomposition or precipitation.

Double decomposition means that the two compounds used in the reaction as reactants both decompose to form new compounds.

Precipitation is the act of throwing down a solid (precipitate) when two aqueous solutions are mixed.

Preparation of Lead (II) Sulphate

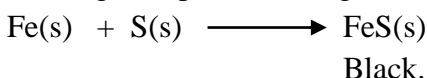
- ❖ Add dilute sulphuric acid to Lead (II) nitrate solution into glass beaker, stirring all the time.
- ❖ Filter off the precipitate and pour off the liquid.
- ❖ Wash the precipitate 2-3 times with hot distilled water to remove any acid.
- ❖ Dry the precipitate in a steam oven or by leaving it in the air.

Direct combination of elements (Direct synthesis)

(i) Preparation of Iron (II) Sulphide

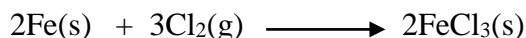
- ❖ Put three spatula endful of powder sulphur to iron filings in a test tube.
- ❖ Add three spatula endful of powdered sulphur to iron filings in a test tube.
- ❖ Heat the mixture strongly in Bunsen flame until no further change.

A red glow spreads through the whole mixture.



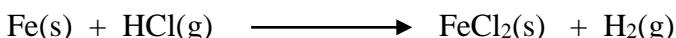
(ii). Preparation of iron (III) Chloride

- ❖ Heat the iron wire in the combustion tube strongly.
 - ❖ Pass a stream of chlorine over heated iron wire in the combustion tube
- The iron wires glows hot and react with chlorine.



(iii). Preparation of iron (II) chloride

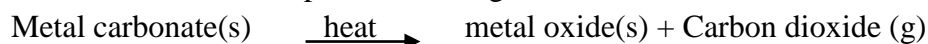
- ❖ Dry hydrogen chloride is passed over strongly heated iron wire.
- ❖ A white anhydrous iron (II) chloride is formed.



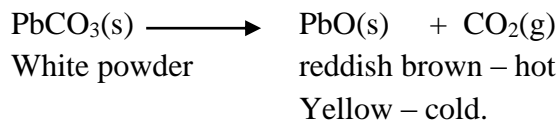
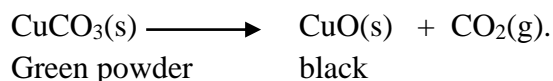
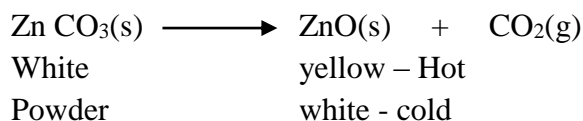
ACTION OF HEAT ON SALTS

(a). Carbonates

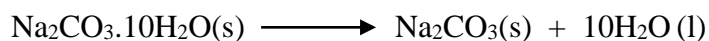
Potassium carbonate and sodium carbonate do not decompose when heated. The carbonates of other metals decompose on heating to form the oxide of the metal and carbon dioxide.



Zinc oxide is yellow when hot and white when cold, Lead (II) oxide is reddish-brown when hot and white when cold, Copper (II) oxide is black, and the other oxides are white.



Hydrated salts when heated gently lose their water of crystallization and then turn into anhydrous salts.



Ammonium carbonate when heated decompose to form ammonia, water and carbondioxide

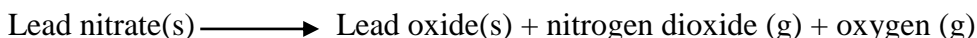


(b). Nitrates.

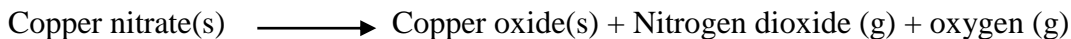
- ❖ Sodium and potassium nitrates melt to colourless liquids. They slowly decompose, forming oxygen and a nitrite, which is pale-yellow.



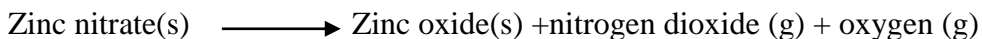
- ❖ Lead (II) nitrate crackles because gas forms inside its crystals and splits them. It then melts and forms a reddish-brown gas, which is a mixture of oxygen and nitrogen dioxide. The solid lead (II) oxide which remains is reddish-brown when hot and yellow when cold.



- ❖ Green copper nitrate splits up in the same way, leaving black copper oxide.



- ❖ Zinc nitrate decomposes to form zinc oxide, which is yellow when hot and white when cold.



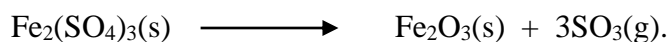
$\left. \begin{array}{l} \text{K} \\ \text{Na} \end{array} \right\}$ metal nitrite + oxygen

$\left. \begin{array}{l} \text{Ca} \\ \text{Mg} \\ \text{Al} \end{array} \right\}$

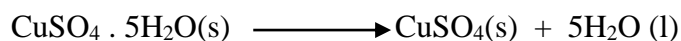
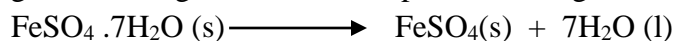
Zn metal oxide + Nitrogen dioxide + Oxygen
 Fe
 Pb
 Cu
 Hg } metal + Nitrogen dioxide + oxygen
 Ag }
 NH₄⁺ } water + dinitrogen oxide

(c) Sulphates.

Most sulphates are stable when heated gently but on strong heating they decompose forming SO₂ or SO₃ and an oxide of a metal.



Hydrated iron (II) sulphate and copper (II) sulphate when heated lose water of crystallization on gentle heating and then decompose on strong heating.

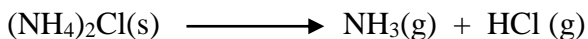


Ammonium sulphate melts and then decomposes to form ammonia and ammonium hydrogen sulphate.



(d). Chlorides

Most chlorides are stable to heat; however, ammonium chloride decomposes to form ammonia and hydrogen chloride

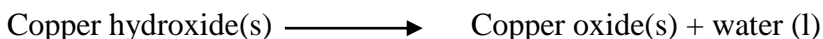


The gases can recombine on cooling.

(e). Hydroxides

Sodium hydroxide and potassium hydroxide melt to colourless liquids, but do not decompose.

All other hydroxides decompose on heating, forming water and metallic oxide.



The oxides of calcium, magnesium and aluminium are white.

SOLUBILITY AND SOLUBILITY CURVES

Solubility of a solute in a solvent at a particular temperature is the mass of a solute in grams required to saturate 100g of solvent at that temperature.

$$\text{Solubility} = \frac{\text{mass of solute that saturates a solution}}{\text{mass of solvent}} \times 100$$

Units of solubility = g/100g of H₂O at a given temperature.

Definitions.

- ❖ An unsaturated solution is one which can take up more solute at that temperature.
- ❖ A saturated solution is one which cannot dissolve any more solute at that temperature in the presence of undissolved solute.
- ❖ A super saturated solution is a solution which contains more solute than is required to form a saturated solution at a particular temperature.

Experiment: To determine the solubility of potassium nitrate

- ❖ Potassium nitrate is dissolved in water to form a saturated solution.
- ❖ The solution in the beaker is allowed to settle so as to allow undissolved solid to separate from the liquid.
- ❖ Weigh the evaporating dish.
- ❖ The solution is transferred to a weighed evaporating dish and the mass of the solution and dish recorded.
- ❖ Solution is evaporated to dryness on a water bath to avoid loss of solid.
- ❖ The solid is allowed to cool and the dish and the solid reweighed.

Treatment of results.

Mass of dish	=a g
Mass of dish+ solution	=b g
Mass of dish + salt	=c g
Mass of solution	= (b-a) g
Mass of salt	=(c-a) g
Mass of water	= (b-c)g
(b-c)g of water contain	= (c-a)
1g of water contain	$= \left(\frac{c-a}{b-c} \right) g$
100g of water contain	$= \left(\frac{c-a}{b-c} \times 100 \right) g$ of salt

Examples.

1(a). 75g of a saturated solution of sodium chloride contains 30g of sodium chloride at 20°C. Calculate the solubility of sodium chloride at 20°C.

Mass of water= mass of saturated solution-mass of solute.

$$\begin{aligned}\text{Mass of water} &= 75-30 \\ &= 45\text{g}\end{aligned}$$

45g of H₂O contain 30g of NaCl

1g of H₂O contains $\left(\frac{30}{45} \right) g$ of NaCl

100g of H₂O contain $\left(\frac{30}{45} \times 100 \right) g$ of NaCl

$$= 66.7\text{g} / 100\text{g of H}_2\text{O}$$

(b) Calculate the percentage of the salt in the saturated solution.

$$\frac{30}{75} \times 100$$

$$= 40\%$$

2. 12g of a saturated solution was evaporated and 4g of solid potassium nitrate was left.

Calculate the solubility of sodium nitrate.

Mass of water = 12 - 4

$$= 8\text{g}$$

8g of H₂O contain 4g of KNO₃

1g of H₂O contains $\left(\frac{4}{8}\right)$ g of KNO₃

100g of H₂O contain $\left(\frac{4}{8} \times 100\right)$ g of KNO₃

$$= 50\text{g} / 100\text{g of H}_2\text{O}$$

3. 23g of potassium chlorate at 21°C saturated 40cm³ of water. Calculate the solubility of potassium chlorate at 21°C.

40g of H₂O contain 23g of KClO₃

1g of H₂O contains $\left(\frac{23}{40}\right)$ g of KClO₃

100g of H₂O contain $\left(\frac{23}{40} \times 100\right)$ g of KClO₃

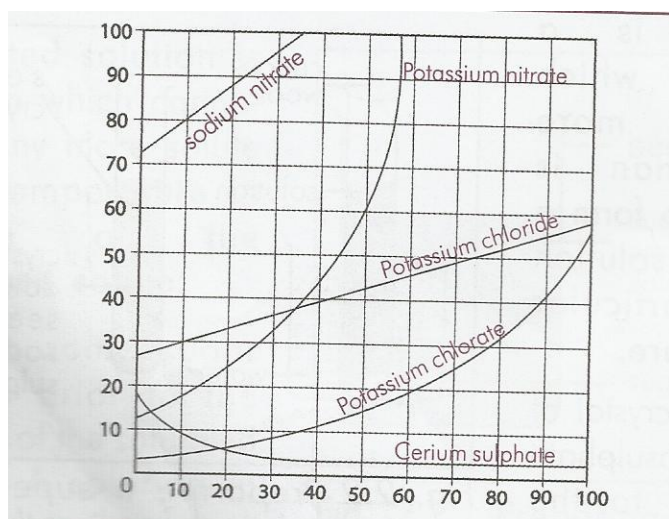
$$= 57\text{g} / 100\text{g of H}_2\text{O}$$

SOLUBILITY CURVES.

The solubility curve of a substance is a graph showing how its solubility varies with temperature. The solubility of some salts eg KNO₃, KClO₃ increase with temperature and their curves rise steeply.

The solubility of some salts like NaCl, KCl change little with rise in temperature and their curves rise gently.

Some compounds like CaSO₄ are less soluble in hot and cold water, their solubility curves fall gently.



Variation in solubility of salt with temperature is the basis of fractional crystallization of salts. Fractional crystallization is used in extraction of sodium carbonate and sodium chloride.

Uses of solubility curve

- To obtain solubility of the salt at various temperatures.
- To determine the temperature at which a certain mass of salt when dissolved in water, can form a saturated solution.
- To calculate mass of salt precipitated by cooling from higher temperature to a lower temperature.

Applications of solubility

- Production of fizzy drinks eg soda which contain carbon dioxide dissolved at high pressure and low temperatures. On opening the bottle, the gas bubbles escape out of the solution.
- Extraction of sodium chloride from sea water by fractional crystallization.
- Purification of salts by fractional crystallization.

Exercise

1. (a) Explain the following terms:

- Solubility of a salt
- Saturated solution

- The solubility of substance **B** is 25g per 100g of water at 25° C. Calculate the mass of **B** required to saturate 90g of water at 25° C.

- (c) 80.0g of a saturated solution of sodium nitrate contains 35.0g of sodium nitrate at 20° C. Calculate the solubility of sodium nitrate at this temperature.
2. (a) Describe an experiment to determine the solubility of potassium nitrate.
- (b) 75g of a saturated solution of potassium nitrate contains 30g of potassium nitrate at 20°C. Calculate the solubility of potassium nitrate at 20° C.
- (c) The table below shows the solubilities of potassium nitrate at different temperatures.

Temperature(° C)	10	30	50	70	90
Solubility in g/ 100g of water	4	10	18	30	46

- (i) Plot a graph of solubility of potassium nitrate against temperature.
- (ii) Use your graph to determine the solubility of potassium nitrate at 20° C and 60° C respectively.
3. The table shows the solubilities of potassium chloride and potassium nitrate at various temperatures.

Temperature (° C)	0	20	40	60
Solubility of KCl(g/ 100g of water)	28	34	39	45
Solubility of KNO ₃ (g/ 100 g of water)	13	32	61	108

- (a) On the same axes, plot graphs of solubilities of potassium chloride and potassium nitrate against temperatures.
- (b) Determine the temperature at which the solubilities of the two salts are equal.
- (c) Which of the two salts dissolves more rapidly with increase in temperature?
- (d) State what would happen if a saturated solution of potassium chloride at 40 ° C was cooled.

(e) A saturated solution of potassium chloride at 40° C was cooled to 35° C .Calculate the mass of potassium chloride crystals formed.

4. (a) Explain the following terms:

(i) Solute

(ii) Super saturated solution

(b) A saturated solution of sodium nitrate was made at 25° C .Use the following data to determine the solubility of sodium nitrate at 25°C.

Mass of evaporating dish = 58.05g

Mass of evaporating dish + salt solution= 123.53g

Mass of evaporating dish + dry salt = 61.28

5. (a) Explain giving example(s) what is meant by

(i) Basicity of an acid.

(ii) An acid salt.

(b) Outline how a pure dry sample of sodium hydrogen sulphate can be prepared in the laboratory.

(c) An acidified solution of barium chloride was added to aqueous sodium hydrogen sulphate.State what was observed and write equation for the reaction that took place.

6. Define the following terms

(a) Saturated solution

(b) Crystallization

(c) Neutralization

(d) Precipitation

(e) Describe how pure lead (II) sulphate can be prepared in the laboratory starting from lead nitrate.

7. (a) Define the terms

(i) A normal salt

(ii) An acid salt

(iii) Basic salt

(b) Give one example of

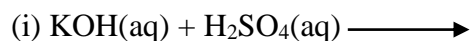
(i) Normal salt

(ii) An acid salt

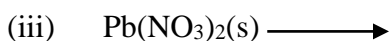
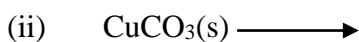
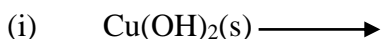
(iii) Basic salt

8. (a) What is basicity of an acid?

(b) Complete and balance the following equations



(c) Complete the following equations to show the effect of heat on the salts



9. (a) Sulphuric acid is a strong acid.

(i) Explain what strong acid means.

(ii) Write equation to show how sulphuric acid ionizes in water.

(b) Write equations for the reactions that occur when dilute sulphuric acid is added to

(i) Zinc carbonate

(ii) Sodium hydroxide solution.

(iii) Magnesium ribbon

10. (a) Give the name and formula of

(i) a dibasic acid

(ii) Amphoteric hydroxide

(iii) Amphoteric oxide

(b) Write equation for the reaction of the acid named with

(i) An amphoteric hydroxide (ii) An amphoteric oxide

(c) Salts can be prepared by the following methods:

(i) Neutralization of an acid by base (ii) Action of a metal on an acid

(ii) Double decomposition

Select the salt from the list below and a suitable method by which it can be prepared.

Salt	Method
Barium sulphate	
Calcium carbonate	
Copper(II) chloride	

Magnesium sulphate	
Sodium hydrogen carbonate	

10.(a).Name four general methods of preparing salts from dilute acids and name one salt prepared from each method.

(b) .Write an equation in each case to show how the acid is used to prepare the salt.

11.Sulphuric acid is a strong dibasic acid.

(a) Explain the terms strong acid, and basicity.

(b).Dilute sulphuric acid was added to a solution of lead(II) nitrate.

(i) State what was observed.

(ii)Write an equation for the reaction.

12.(a) What is a salt?

(b) What are the products of the reaction between acid and a base?

(c) Give three characteristics of a base and three characteristics of acid.

13.(a) Explain the following terms:

(i) Double decomposition

(ii) Precipitation

(c) Write ionic equations for the reaction between hydrochloric acid and

(i) Sodium carbonate solution

(ii) Sodium hydroxide solution

CARBON AND ITS COMPOUNDS.

Carbon is an element with mass number 12 and atomic number 6. It's a constituent of coal, Its symbol is C.

Occurrence

Pure carbon is found in the form of diamond and impure carbon as graphite. It is a constituent of numerous naturally occurring substances such as coal, mineral oils, carbonates, organic matter and occurs in air to a small but very important extent, 0.003% as carbon dioxide.

FORMS OF CARBON

Carbon exists in many forms most importantly as diamond and graphite. The other form in which carbon exists is called amorphous carbon e.g animal charcoal, wood charcoal, lamp black, sugar charcoal, coke and soot.

ALLOTROPY (POLYMORPHISM)

Allotropy is the existence of an element in different forms without change of physical state. The different forms of the element are known as allotropes.

Allotropes are different distinct forms in which an element exists without change of physical state.

Polymorphism is the existence of a substance in more than one crystalline form.

ALLOTROPE OF CARBON

Carbon exists in two allotropic modifications,

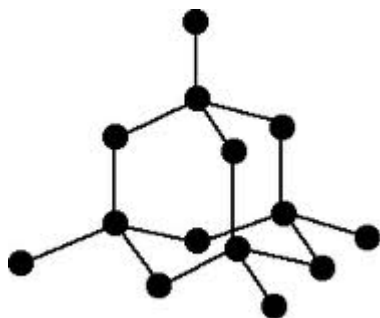
1. Diamond
2. Graphite

Proof for allotropy of carbon

When equal masses of graphite and diamond are separately burnt in excess pure oxygen, each produces the same quantity of carbon dioxide. The product turns lime water milky.

STRUCTURE OF DIAMOND.

In diamond, each carbon atom is bonded to four other carbon atoms by strong covalent bonds. The carbon atoms in diamond are tetrahedrally arranged and are equidistant from each other. The bonding extends in 3 dimensions to form a giant atomic structure.



PHYSICAL PROPERTIES OF DIAMOND

1. It's a non conductor of electricity, because all the electrons in the carbon atoms have been used for bonding. There are no delocalized electrons in diamond.
2. It has high density, due to close packing of carbon atoms. The density of diamond is 3.5gcm^{-3}
3. Its colourless, crystalline and transparent solid.
4. It's the hardest substance known.
5. It is resistant to melting. Melting point of diamond is 3930°C .

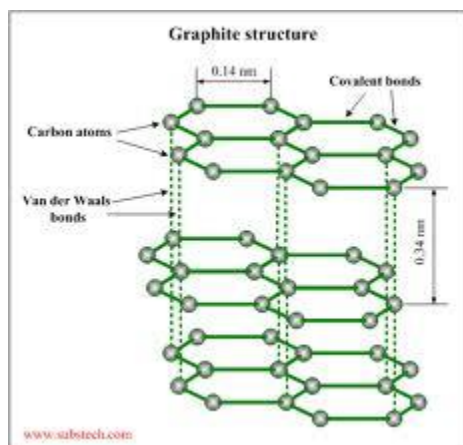
USES OF DIAMOND

1. It's used in making jewels. It can refract light producing sparkling brilliance.
2. It's used for making drills and glass cutters because of its hardness.

STRUCTURE OF GRAPHITE

In graphite, each carbon atom is covalently bonded to three other carbon atoms forming layers of hexagons in parallel planes. The layers are held by weak Vander Waals forces. These forces allow movement of the planes parallel to each other and this makes graphite very soft.

Note: Graphite conducts electricity because of delocalized electrons within the layers.



PHYSICAL PROPERTIES OF GRAPHITE.

1. Graphite is black and opaque.
2. It's soft, greasy or soapy.
3. It's a good conductor of electricity.
4. It has low density (2.3gcm^{-3})

USES OF GRAPHITE

- It's used as a lubricant because of its slippery nature.

- It's used as an electrode because it conducts electricity.
- It's used in the manufacturer of lead pencils because it's black.
- It is used in atomic piles in which uranium atoms split up to form energy.

DIFFERENCES BETWEEN GRAPHITE AND DIAMOND.

GRAPHITE	DIAMOND.
<ul style="list-style-type: none"> • Soft and slippery • Low density (2.3gcm^{-3}) • Good conductor of electricity • Black and Opaque 	<ul style="list-style-type: none"> • Hardest substance • High density (3.5gcm^{-3}) • Non conductor of electricity. • Colourless and transparent.

AMORPHOUS CARBON

These are forms of carbon without clearly defined shape or form.

Amorphous carbon exist in many forms such as wood charcoal, animal charcoal, sugar charcoal, lampblack and coke.

1. Wood Charcoal

This is made by burning wood in limited supply of air.

Uses

- ❖ It's used in gas masks as it is a good absorbent for gases such as ammonia, chlorine and bromine.
- ❖ It's as fuel in homes for cooking and ironing clothes.

2. Animal Charcoal.

It's made by heating bones in limited supply of air. It contains a lot calcium phosphate and can be used as a fertilizer.

Uses

- ❖ Its used to purify substances and to remove brown colour from sugar making it white.
- ❖ Its used for treating snake bites as magic stone. Bone charcoal has adsorbing property. It is placed at the snake bite wounds and pressed hard for some time. The animal charcoal adsorbs the snake poison.

3. Sugar charcoal.

This is pure carbon made by dehydrating cane sugar or glucose with concentrated sulphuric acid or by heating sugar in the absence of air to drive off water.



4. Lamp black

Its made when petroleum, kerosene, turpentine or candles burn in limited supply of air.

Uses.

- ❖ Used as a filler for rubber used to make tyres.
- ❖ Used in printers' ink, shoe polish, and carbon paper.

5. Coke.

This is the solid left when coal is heated in limited supply of air.

Uses

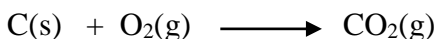
Coke is used as fuel in furnaces, boilers and oven.

CHEMICAL PROPERTIES OF CARBON

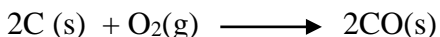
Carbon is not a very reactive element.

1. Combustion:

Carbon burns in plentiful supply of oxygen to form carbondioxide.



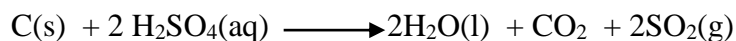
In limited air, carbon burns to form carbon monoxide.



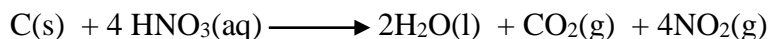
2. Reaction with Acids.

(i). Reaction with concentrated sulphuric acid

Concentrated sulphuric acid oxidizes carbon to carbondioxide which turns lime water milky and itself reduced to water and sulphur dioxide which turns acidified potassium dichromate from orange to green.

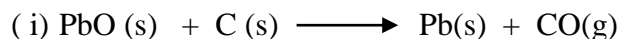


(ii). Reaction with concentrated nitric acid. Concentrated Nitric acid oxidizes carbon to carbondioxide and itself reduced to nitrogen dioxide which appears as reddish brown fumes.



3. Reducing Action of Carbon.

Carbon reduces oxide of zinc, lead, iron and copper to form a metal and carbon monoxide.

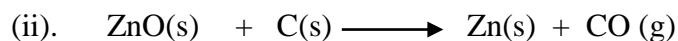


Yellow – Hot

Grey

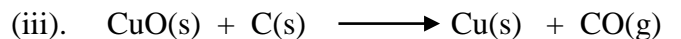
Redish brown – Hot

(Silvery beads)



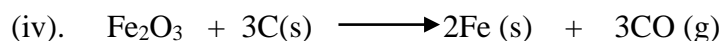
Yellow-Hot

White – cold



Black

reddish-brown



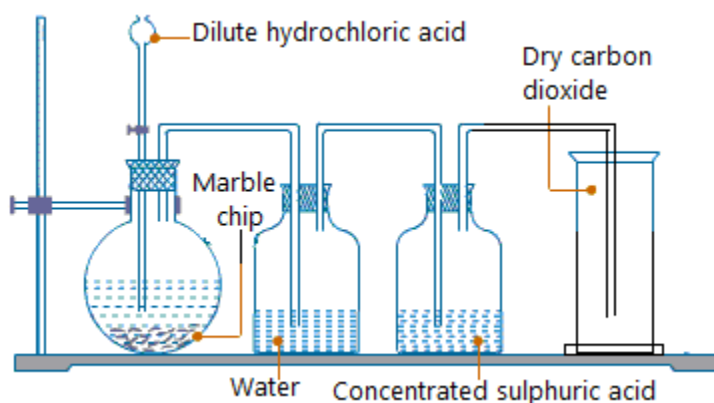
Note. Oxidation – reduction takes place in these reactions. This technique is used in the extraction of metals such as lead, iron and zinc.

Exercise

- 1.(a) Describe the structure of;
 - (i) Diamond
 - (ii) Graphite
- (b) Give three properties of diamond and graphite.
- (c) Give two uses of diamond and graphite basing on the properties in(b).
- 2.(a) What is allotropy?
 - (b) Name the two allotropes of carbon
 - (c) Give four differences between the allotropes of carbon named in (b).
 - (d) State 5 forms of amorphous carbon and give the uses of each form.
3. (a) Describe an experiment to show that diamond and graphite are allotropes of carbon.
 - (b) State what is observed and write equation for the reaction when carbon reacts with the following:
 - (i) Concentrated sulphuric acid.
 - (ii) Concentrated nitric acid.
 - (iii) Copper(II) Oxide
 - (iv) Lead (II) oxide
 - (v) Zinc oxide

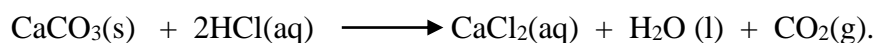
CARBONDIOXIDE.

Laboratory preparations of Dry Carbondioxide.



Dilute hydrochloric acid is poured down the thistle funnel on to calcium carbonate in the flask. Effervescence occurs and a colourless gas is evolved. The gas is passed through potassium hydrogen carbonate (or distilled water) in a wash bottle to remove hydrochloric acid spray. Its then passed through a wash bottle containing concentrated sulphuric acid to dry it. The gas is then collected by down ward delivery as it's denser than air.

Equation.



Ionic equation.



PHYSICAL PROPERTIES OF CARBON DIOXIDE

1. It's a colourless gas, with very faint smell and a sharp pleasant taste.
2. It does not support combustion.
3. it's slightly soluble in water, forming carbonic acid.

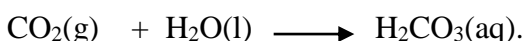
TEST FOR CARBON DIOXIDE.

Carbon dioxide can be tested for by passing the gas into a solution of calcium hydroxide (lime-water) which turns milky due to the formation of calcium carbonate.

REACTIONS OF CARBONDIOXIDE.

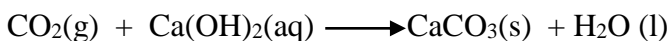
1. Reaction with water

Carbondioxide dissolves slightly in water forming carbonic acid which turns litmus solution pale-red.

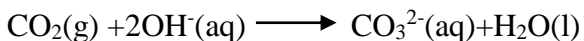


2. Reaction with lime water (calcium hydroxide solution)

When carbondioxide is passed through a solution of calcium hydroxide, a white precipitate is formed due to the formation of insoluble calcium carbonate.

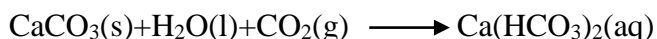


Ionic equation



When excess carbondioxide is passed, the calcium hydroxide solution will clear again, due to the formation of soluble calcium hydrogen carbonate.

Equation

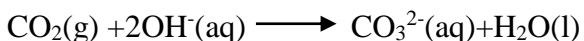


3. Reaction with other alkalis

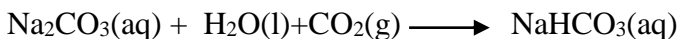
When carbondioxide is bubbled through cold concentrated sodium hydroxide or potassium hydroxide, it's quickly absorbed.



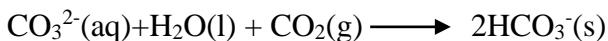
Ionic equation



When excess carbondioxide is bubbled through sodium hydroxide solution for some time, a white solid(sodium hydrogen carbonate) appears on the bottom of the boiling tube.

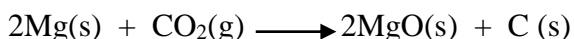


Ionic equation



4. Reaction with Magnesium.

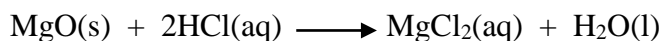
Magnesium continues to burn for a short time with a sputtering flame in a gas jar of carbondioxide forming black particles and a white solid.



Explanation.

Burning magnesium produces a lot of heat which decomposes carbondioxide into black carbon and oxygen. The oxygen supports the burning of magnesium for a short time forming white magnesium oxide.

When dilute hydrochloric acid is added to the products in a gas jar and shaken, the white powder dissolves in the acid leaving behind carbon. This indicates that carbon dioxide contains carbon and is an oxide of carbon.



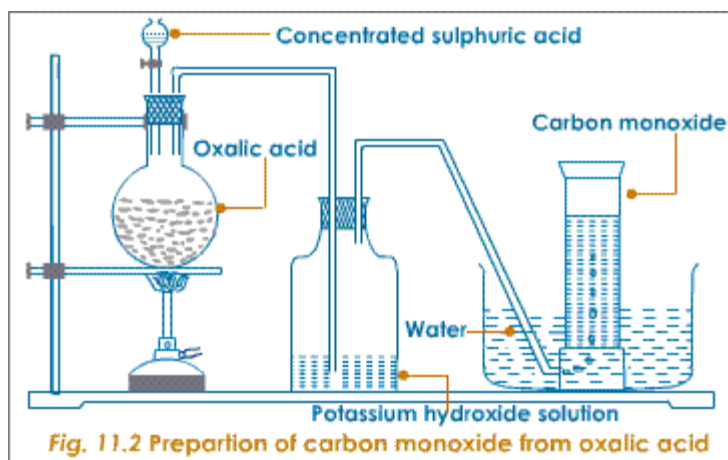
USES OF CARBON DIOXIDE.

1. In fire extinguishers, because it does not burn and its denser than air.
2. Its used in the manufacture of aerated drinks e.g. Soda
3. It's used as a refrigerant for perishable goods and ice cream.
4. It's used in Ammonia-soda or Solvay process for making sodium hydrogen carbonate and sodium carbonate.
5. Making rain-pieces of solid carbondioxide can be dropped into the clouds to cool them enough to form rain.
6. Carbondioxide is used in bakeries. Baking powder when heated form carbondioxide which forms bubbles inside bread and cakes making them light.

CARBONMONOXIDE, CO

Preparation of Carbon monoxide.

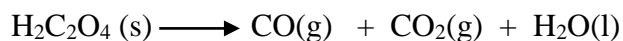
1. From Ethanedioic acid (Oxalic acid)



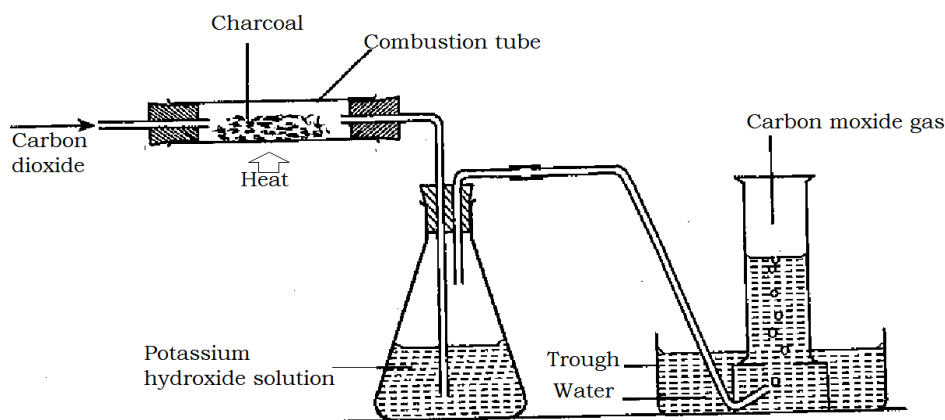
Concentrated sulphuric acid is poured down the thistle funnel on to oxalic acid crystals in the flask and the mixture gently warmed.

Oxalic acid crystals dissolve, Effervescence of colourless gases evolved. The mixture of gases is passed through wash bottle containing concentrated potassium hydroxide solution to absorb carbondioxide.

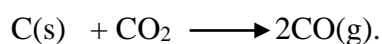
Carbon monoxide passes and is collected over water by upward delivery because it's less dense than air.



2. From carbon dioxide



Carbondioxide is passed over red-hot charcoal in combustion tube. Excess carbondioxide is passed through a wash bottle containing concentrated Potassium hydroxide solution to absorb it.



PHYSICAL PROPERTIES OF CARBON-MONOXIDE.

1. It's colourless, tasteless and odourless.
2. It's neutral to litmus.
3. It's slightly less dense than air.
4. It's only slightly soluble in water.

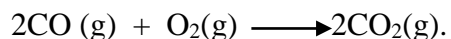
TEST FOR CARBON MONOXIDE

Carbon monoxide burns in air with a blue flame to form carbon dioxide.

REACTIONS OF CARBON MONOXIDE.

1. Combustion:

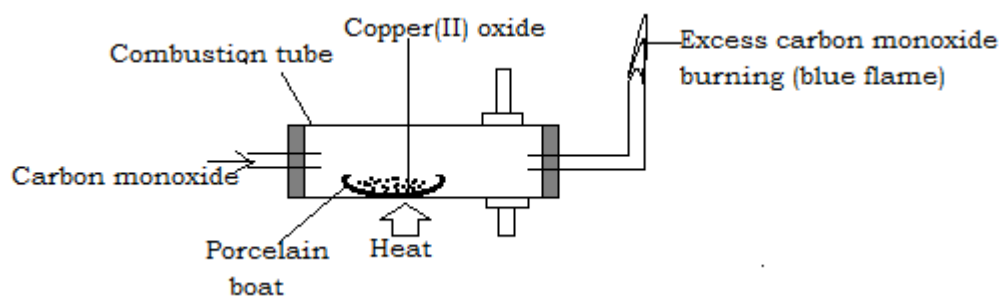
Carbon monoxide burns in air with a blue flame forming carbon dioxide.



2. Reducing Action.

Carbon monoxide is quite a strong reducing agent and reduces oxides of copper, lead, iron and Zinc to the metal and itself oxidized to carbon dioxide.

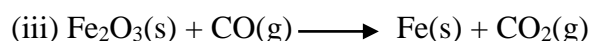
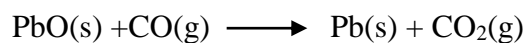
(i) *Effect of carbon monoxide on copper (II) oxide*



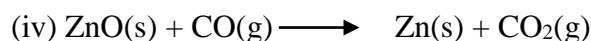
Dry carbon monoxide reduces black copper (II) oxide to form a brown solid.



(ii) When carbon monoxide gas is passed over heated lead (II) oxide, a grey metallic lead is formed.

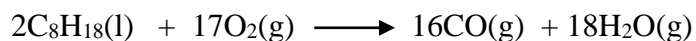


This reaction occurs at high temperature in the blast furnace.



CARBON MONOXIDE AS A POLLUTANT.

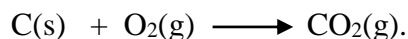
Car exhaust fumes contain carbon monoxide formed by incomplete combustion of fuel.



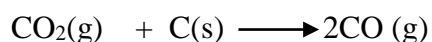
(Petrol)

In charcoal stove, carbon monoxide is produced when carbon burns in the absence of air. The reaction can occur in two stages:

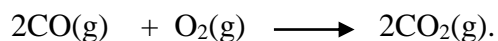
Carbon burns in oxygen to form carbon dioxide



The carbon dioxide reacts with more carbon to form carbon monoxide



The carbon monoxide burns in more oxygen with a blue flame to form carbon dioxide.



POISONOUS NATURE OF CARBON MONOXIDE

Blood contains red haemoglobin, which combines with oxygen to form oxy-haemoglobin. This moves to all parts of the body and supplies them with oxygen.

Carbon monoxide combines with haemoglobin to form bright red substances, carboxyhaemoglobin which stops blood from carrying oxygen.

The person becomes ill or dies of suffocation.

Differences between carbon dioxide and carbon monoxide.

Carbondioxide	Carbon monoxide
❖ Not poisonous	❖ Poisonous
❖ Soluble in water and alkalis	❖ Insoluble in water and alkalis
❖ Denser than air	❖ Less dense than air
❖ Does not burn	❖ Burns with a blue flame

Both gases are colourless and odourless and do not support combustion.

PRODUCER GAS.

Producer gas is a mixture of carbon monoxide and Nitrogen.

Uses

- ❖ It's used to heat the retorts during the manufacture of Zinc to heat lime kilns, steel furnaces and glass furnaces.
- ❖ It's also the source of nitrogen in the Haber process for the manufacture of ammonia.

WATER GAS

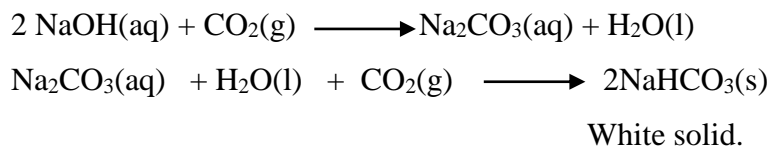
Water gas is a mixture of hydrogen and carbon monoxide.

Use: It's used in the manufacture of Hydrogen.

CARBONATES AND HYDROGEN CARBONATES.

Preparation of sodium carbonate:

Pass carbon dioxide into a moderately concentrated sodium hydroxide solution for some time until a white solid is formed.



Filter of the solid and wash it 2-3 times with a little cold distilled water.

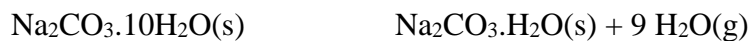
Transfer the solid to a hard glass tube and heat, sodium carbonate is left as a fine white powder.



Dissolve the powder in the minimum amount of hot water and crystallize to obtain

$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ (washing soda).

Washing soda, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$, large translucent crystal, on exposure to air lose mass and becomes coated with a fine powder which renders them opaque. Each molecule of washing soda gives up 9 molecules of water of crystallization. This process of giving up water of crystallization to the atmosphere is termed efflorescence.

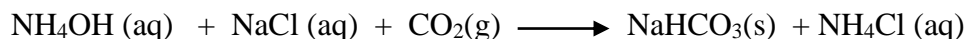


SOLVAY PROCESS.

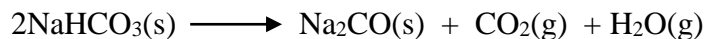
On large scale, sodium carbonate is manufactured by the Solvay process.

Sodium carbonate is manufactured by dropping brine slowly down an absorber tower up which Ammonia passes. The brine becomes saturated with ammonia.

The Ammonia – brine flows slowly down a Solvay tower up which carbondioxide passes to form sodium hydrogen carbonate and ammonium chloride.



The sodium hydrogen carbonate is filtered off and heated to form anhydrous sodium carbonate



(Soda ash)

USES OF SODIUM CARBONATE.

- ❖ In the making of glass.
- ❖ In the softening of water.
- ❖ In soap industries.
- ❖ In paper and textile industries (NaOH).

Test for carbonates

Put some of the suspected carbonate in a test tube and add dilute nitric acid. If the carbonate is present there will be effervescence and the gas which turns calcium hydroxide solution milky evolved.

Preparation of sodium hydrogen carbonate

Bubble carbondioxide through concentrated sodium hydroxide solution, a white powder is precipitated.



The sodium hydrogen carbonate is deposited as a white powder which is filtered off, washed 2-3 times with a little cold water and allowed to dry.

USES OF SODIUM HYDROGEN CARBONATE.

1. It's used in baking powder for making cakes and bread.
2. It's used in fire extinguishers.

EFFECTS OF HEAT ON CARBONATES AND HYDROGEN CARBONATES.

Carbonates.

All carbonates are decomposed by heat except sodium and potassium carbonates. They decompose to form a metal oxide and carbondioxide.

Metal carbonate \longrightarrow metal oxide + carbon dioxide + water

Examples.

1. **Zinc carbonate**; Zinc carbonate is a white powder which when heated decomposes to form a yellow solid residue when hot and turns white on cooling.



2. **Lead carbonate**: The white powder decomposes on heating to form a reddish-brown residue when hot and turns yellow on cooling.



3. **Copper(II) carbonate**: The bright green powder decomposes on heating to liberate carbon dioxide and a black solid residue is left.



4. **Ammonium carbonate**: Ammonium carbonate is non metallic and decomposes on heating to liberate ammonia gas which turns moist red litmus paper blue.



Hydrogen carbonates:

All hydrogen carbonates are decomposed by heat to form carbonates of the metal, water and carbondioxide.

Metal hydrogen carbonate \longrightarrow Carbonate + Water + carbon dioxide

1. Sodium hydrogen carbonate.

When a small amount of sodium hydrogen carbonate is heated in a dry test tube, a gas is given off which turns the calcium hydroxide solution milky and water is seen to collect on the cooler parts of the test tube.



2. Calcium hydrogen carbonate decomposes when heated to form calcium carbonate, water and carbon dioxide.



3. Magnesium hydrogen carbonate decomposes when heated to form magnesium carbonate, water and carbon dioxide.



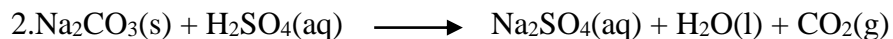
EFFECTS OF ACIDS ON CARBONATES AND HYDROGEN CARBONATES.

Carbonates.

Carbonates react with mineral acids to form acid salt, water and carbon dioxide



Examples



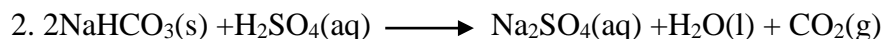
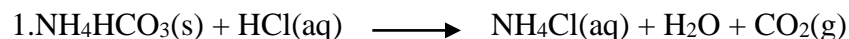
Ionic equation for reaction of carbonates with acid:



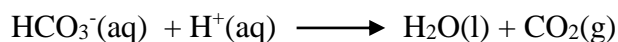
Hydrogen carbonates

Hydrogen carbonates react in a similar way to carbonates forming salt, water and carbon dioxide.

Examples



The ionic equation for reaction of hydrogen carbonates with acids:



TEST FOR CO_3^{2-} AND HCO_3^-

Reagent: Magnesium sulphate solution.

Observation

With CO_3^{2-} - a white precipitate is formed.

With HCO_3^- - no observable change / no precipitate.

Stalactites and Stalagmites.

Stalactites are solid structures which grow hanging from the roof of a cave.

Stalagmites are solid structures which grow from the floor of a cave.

They are deposits of calcium carbonate or magnesium carbonate formed when rain water containing calcium hydrogen carbonate or magnesium hydrogen carbonate evaporate.

CARBON CYCLE.

Carbon dioxide is added to the atmosphere by:-

1. Combustion: The burning in air of fuels containing carbon e.g wood.
2. Respiration.
3. Making lime: lime is made from lime stone.



4. Making of beer and wines by fermentation of sugar or starch.



Carbondioxide is removed from the atmosphere by:-

- Solution in water: rain dissolves carbon dioxide and rivers, lakes, seas and oceans contain much of it.
- Photosynthesis – green plants absorb carbon dioxide to make sugar.
- Hardening of Mortar – Mortar and white wash contain calcium hydroxide which slowly reacts with CO_2 to form calcium carbonate.

Exercise

- 1 .(a) Describe briefly, a simple experiment that can be carried out in the laboratory to prove that graphite and diamond consists of carbon atoms only.
(b) Carbon dioxide is normally prepared in the laboratory by action of dilute hydrochloric acid on calcium carbonate.
(i) Write equation for the reaction.
(ii) Explain why sulphuric acid is not used to prepare carbon dioxide from calcium carbonate.

(c) Name a reagent that can be used to distinguish between the following pairs of ions and in each case, state what was observed when each ion is separately reacted with the reagent.

(i) $\text{CO}_3^{2-}(\text{aq})$ and $\text{HCO}_3^-(\text{aq})$.

(ii) $\text{CO}_3^{2-}(\text{aq})$ and $\text{SO}_4^{2-}(\text{aq})$.

2. (a) What is meant by hard water?

(b)(i) Name two ions responsible for the hardness of water.

(ii) Write an ionic equation for the reaction that takes place when soap solution is added to hard water.

(iii) State one advantage of hard water.

3. Carbon dioxide was bubbled into a fairly concentrated solution of sodium hydroxide and no visible change was observed at first. On further bubbling, a white precipitate was observed.

(a) Explain the observation.

(b) Write equations for the reactions.

4.(a) Name two crystalline allotropes of carbon.

(b) State four differences between the allotropes you have named.

(c) Give two uses of each of the allotropes.

5. (a) Name one compound that when reacted with dilute hydrochloric acid can produce carbon dioxide.

(b) Excess carbon dioxide was passed through ice-cold sodium hydroxide solution.

(i) State what was observed.

(ii) Write equations for the reaction that took place.

6. (a) (i) Draw a labelled diagram of the set up of the apparatus that can be used to prepare a dry sample of carbon dioxide in the laboratory.

(ii) Write an equation for the reaction that took place.

(b) Burning magnesium was lowered into a jar of carbon dioxide.

(i) State what was observed.

(ii) Explain the observations.

- (c) Water was added to the product in (b) and the resultant mixture tested with litmus. State what was observed.
- (b) When a solution of sodium hydroxide was exposed to air, a white solid was formed on the surface.
- (i) Name the white solid.
- (ii) Write equation to show how the white solid is formed.
7. (a) Describe how sodium carbonate decahydrate, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$, were exposed to air for about 3 days.
- (i) State what was observed.
- (ii) Name the process that has taken place.
- (iii) Write an equation for the reaction that took place
- (b) Dilute hydrochloric acid was added to some crystals of sodium carbonate in a test tube.
- (i) Write equation for the reaction that took place.
- (ii) State the test for the gas evolved and state what is observed during the test.
- (iv) Write the equation for the reaction in c(ii).
8. (a) A compound of metal N was heated strongly until no further change took place. A reddish-brown residue was formed.
- (i) Identify metal N and write the formula of the carbonate.
- (ii) Give the name and formula of the residue.
- (iii) Write the equation for the action of heat on this carbonate.
- (iv) Write the equation for the reaction of the residue with dilute nitric acid.
- (b) Carbon dioxide was passed through calcium hydroxide solution for a long time.
- (i) State what was observed and write equations for the reaction that took place.
- (ii) Burning magnesium was plunged into a gas jar of carbon dioxide. Explain what was observed.
9. (a) Describe the structure of graphite.
- (b) Explain why graphite conducts electricity where as diamond does not.
- (c) State two uses of diamond.
- (c) Carbon monoxide was passed over strongly heated copper(II) oxide.
- (i) State what was observed.
- (ii) Write the equation for the reaction.
- (iii) Name any other oxide that shows similar reaction with carbon dioxide.

10. (a) Copper(II) carbonate was heated strongly until no further change.
- (i) State what was observed.
 - (ii) Write equation for the reaction.
 - (iv) Name the reagent which can be used to identify the gaseous product.
- (b) Excess dilute sulphuric acid was added to the residue in (a) and the mixture warmed.
- (i) State what was observed.
 - (ii) Write an equation for the reaction.
- 11.(a) Define the term allotropy.
- (b) Name the two allotropes of carbon.
 - (c) Give two other elements other than carbon that show allotropy and name their allotropes.
 - (d) Name one process in each case by which the concentration of carbon dioxide in the atmosphere is
- (i) Increased
 - (ii) decreased
12. (a) Describe how sodium carbonate crystals(washing soda) can be prepared.
- (b) State what is observed and write equation for the reaction that takes place when the following carbonates are heated:
- (i) Lead carbonate
 - (ii) Copper (II) carbonate
 - (iii) Zinc carbonate
 - (iv) Ammonium carbonate
- (d) Write equations for the reaction of the following hydrogen carbonates with hydrochloric acid.
- (i) Sodium hydrogen carbonate
 - (ii) Magnesium hydrogen carbonate
 - (iii) Calcium hydrogen carbonate

ORGANIC CHEMISTRY

Organic chemistry is the study of carbon compounds other than oxides of carbon, carbonates, hydrogen carbonates and carbides. Carbon is the main element in organic compounds. Carbon differs from other elements in one unique way – its atoms can join together and form long chains or rings.

HYDRO CARBONS

A hydrocarbon is a compound which contains carbon and hydrogen only. They are simple organic compounds.

Homologous series.

Organic compounds are divided into groups or families of compounds with similar properties.

A homologous series is a group of compounds of similar structure in which each member differs from the next by the presence of an additional – CH_2 group.

Characteristics of homologous series

Usually all the compounds in any one homologous series;

1. Can be prepared by similar methods.
2. Can be represented by a general formula, alkanes $\text{C}_n\text{H}_{2n+2}$, alkenes C_nH_{2n} , alkynes $\text{C}_n\text{H}_{2n-2}$.
3. Have similar chemical properties.
4. Show a gradual change of physical properties.
5. Differ from the next by $-\text{CH}_2$.

The three main homologous series of hydrocarbons are;

1. Alkanes
2. Alkenes
3. Alkynes

ALKANES.

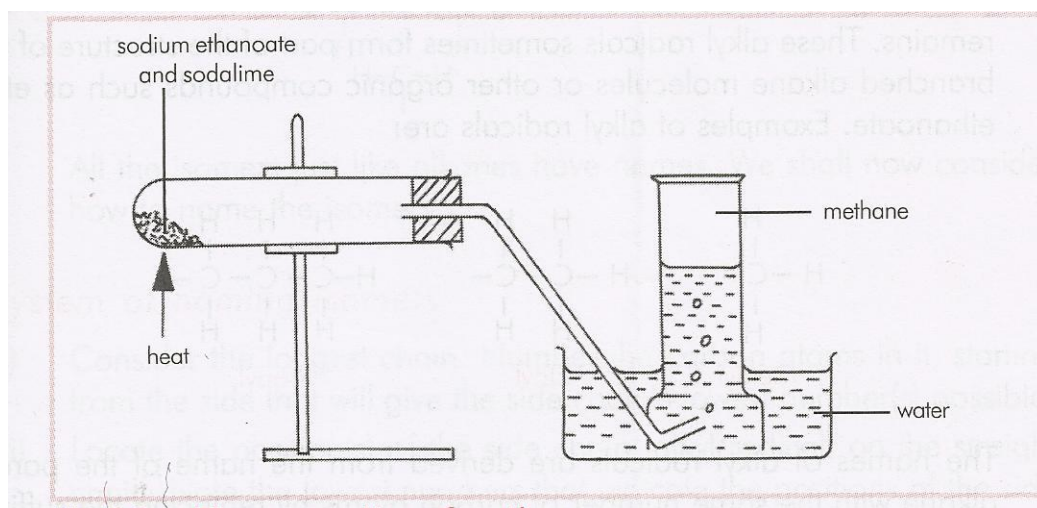
Alkanes are saturated hydrocarbons in which carbon atoms are bonded to others by single covalent bonds. They have a general formula $\text{C}_n\text{H}_{2n+2}$. Alkanes are normally very stable compounds and are known as paraffins. Paraffin means 'little affinity'.

Nomenclature

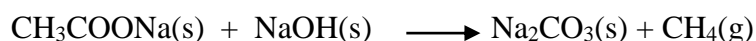
The first four alkanes have special names and all alkanes end with the suffix **-ane**. Alkanes with five carbon atoms onwards are named after polygons with the same number of sides as the carbon atoms in the alkane.

Formula	Name
1. CH ₄	Methane
2. C ₂ H ₆	Ethane
3. C ₃ H ₈	Propane
4. C ₄ H ₁₀	Butane

LABORATORY PREPARATION OF METHANE



A mixture of sodium ethanoate and soda lime is put into a hard glass tube and the mixture heated strongly. A colourless methane gas is evolved and collected over water by upward delivery because it's insoluble in water.



PHYSICAL PROPERTIES OF ALKANES

- The first four members are gases, C₁₂ – C₁₆ are liquids, the rest are waxy solids.
- They are insoluble in water and less dense than water.
- The melting point and boiling point increase gradually as the number of carbon atoms increase.

Physical properties of methane

- Methane is a colourless odourless gas.
- It is only slightly soluble in water.
- It has no effect on litmus.

- It is less dense than air.
- It is a stable compound and does not react with acids and alkalis.

CHEMICAL PROPERTIES OF ALKANES

Alkanes are generally inert and do not react with chemical reagents under ordinary conditions. However, they burn in air and chlorine.

Combustion:

In plentiful supply of oxygen, alkanes burn to form carbon dioxide and water with liberation of much heat.



In limited supply of oxygen, alkanes burn to form carbon monoxide and water with liberation of much heat.



Reaction with chlorine.

Methane reacts slowly with chlorine, at ordinary temperature, but explodes in bright sun light to form chloromethane.



Sources of Alkanes

1. Natural gas - This is a mixture of many gases but contains the largest percentage of methane gas which usually accompanies petroleum deposits. Natural gas is non poisonous and has no smell.
2. Fractional distillation of crude oil.

Crude Oil is heated until most of it vaporizes. The vapour passes into the bottom of a tall fractionating tower. This is divided into several compartments each cooler than the one below. Petrol, the most volatile fraction, condenses at the top, paraffin oil in the compartment below, then diesel and lubricating oils, the non volatile heavy oils run out at the bottom of the column.

USES OF PETROLEUM FRACTIONS

1. Natural gas

Its used as fuel for cooking and lighting.

2. Petrol(Naphtha)

- (i). It's used as motor and aviation fuel.

(ii). It's used as a solvent for grease stains and paints.

3 Paraffin oil (Kerosine)

It's used for lighting, heating, in diesel engines.

4. Diesel

It's used as fuel in diesel engines and in furnaces.

5. Lubricating oil

For lubrication and to make petroleum jellies e.g. Vaseline.

6. Paraffin wax.

Making candles and greases.

7. Bitumen

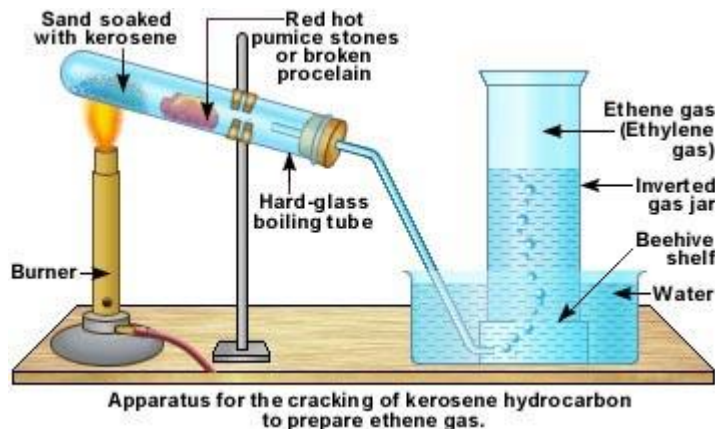
(i) To make protective coating for road surfaces and concrete roof tops.

(ii).As abinding agent for roofing sheet.

CRACKING OF OILS

Cracking is breaking down the large molecules of oils into smaller molecules of petrol and gases.

The gases are mainly alkenes.



TYPES OF CRACKING.

1 Thermal cracking

Gas oil is heated at about 500°C under pressure so that it remains a liquid which is fractionated to obtain petrol.

2 Steam cracking

Oil and steam are heated to about 900°C for less than a second, petrol, fuel oil, alkenes and hydrogen are formed.

3 Catalytic cracking

This uses aluminium (III) oxide and silicon (IV) oxide as catalysts and the process occurs at lower temperatures and pressures. It produces better quality petrol.

Uses of alkanes:

1. Alkanes are used as fuels e.g natural gas and coal gas.
2. Methane when heated to 1000°C decomposes to form hydrogen and carbon black (lamp black) which is used to make printer's ink, paints and carbon papers.
3. Bottled gas (liquefied propane and butane) is used for welding.

BIO GAS

Bio gas is formed by the anaerobic action of bacteria on cellulose and other organic matter. Bio gas contains 75% methane. Other gases include carbon dioxide, ammonia and hydrogen sulphide. The source of bio gas is animal wastes and plant wastes.

A bio gas generator consists of a container in which animal wastes or plants are mixed with a limited amount of water and then covered to exclude aerial oxidation. At temperature of $25-30^{\circ}\text{C}$, the anaerobic bacteria present decompose the wastes to form bio gas.

Bio gas is used for cooking and lighting purposes.

Advantages of bio gas production

- (i) Bio gas is cheap to produce.
- (ii) The solid bi-product is used as fertilizers since it contains high nitrogen content.

Disadvantages of bio gas production

- (i) Some of the gases contained in bio gas are air pollutants.
- (ii) Sulphur dioxide leads to formation of acid rain which results in damage to plants and aquatic organism.
- (iii) Hydrogen sulphide and ammonia contained in bio gas cause irritation of eyes.

ALKENES

Alkenes are unsaturated hydrocarbons which contain a double bond between two carbon atoms in the molecule. They have a general formula C_nH_{2n} where $n=2$ onwards. The alkenes are unsaturated hydrocarbons because the carbon atoms do not exert their normal valency.

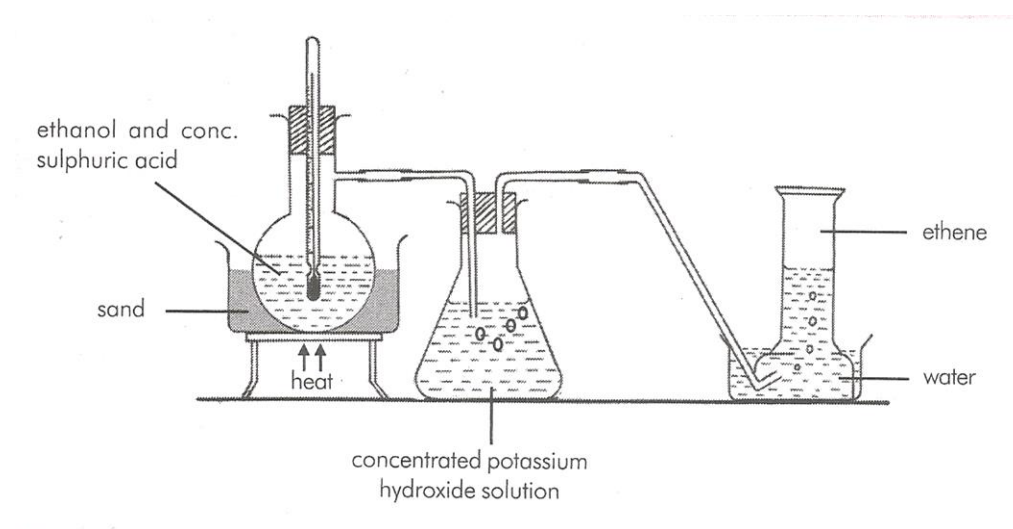
An unsaturated compound is one in which a carbon atom does not exert all its combining power with other atoms. This compound will possess a multiple bond. The molecules of alkenes are characterized by one carbon to carbon double bonds.

Nomenclature

The alkenes derive their names from alkanes with the same number of carbon atoms by replacing the suffix **-ane** in all alkanes by the suffix **-ene** for alkenes.

n	Formula	Name
2	C ₂ H ₄	Ethene CC
3	C ₃ H ₆	Propene
4	C ₄ H ₈	Butene
5	C ₅ H ₁₀	Pentene

PREPARATION OF ETHENE FROM ETHANOL.



Ethene is prepared by dehydration of ethanol by excess hot concentrated sulphuric acid. The mixture is heated on sand bath until the thermometer reading is about 180°C. Ethene is evolved and passed through a wash bottle containing concentrated potassium hydroxide solution to remove sulphur dioxide and carbon dioxide. Ethene is collected over water since it is insoluble in water.

Conditions for the reaction.

- Excess concentrated sulphuric acid
- Heating to a temperature of 180°C.

Equation.



Alternative method:

Ethene is prepared by dehydration of Ethanol.

Ethanol soaked in glass wool is heated strongly. It vaporizes and its vapour is passed over aluminium (III) oxide. Ethene is obtained and collected over water.

**PHYSICAL PROPERTIES OF ETHENE.**

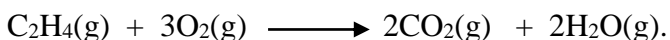
- Ethene is a colourless gas.
- It has a faint sweet smell.
- It's slightly soluble in water.
- It's density is about the same as that of air.
- It's neutral to litmus.

General characteristics of alkenes

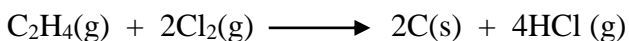
- The first 3 members of alkene series are gases, alkenes with 5-15 carbon atoms are liquids and the higher alkenes are solids.
- Alkenes are insoluble in water but soluble in organic solvents like ethanol.
- They have a characteristic smell and burn with a luminous flame.
- They are less volatile than alkanes, so their boiling points, melting points and densities are higher than those of corresponding alkanes.

CHEMICAL REACTIONS OF ETHENE**1 Combustion.**

Ethene burns in air with a bright smoky flame.



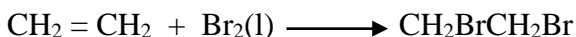
It also burns in chlorine with a red flame forming carbon.

**2 Addition reactions of Ethene****(i). Reaction with bromine**

Ethene is bubbled through bromine water in a test tube.

Observation: The reddish –brown bromine turns to a colourless, oily liquid.

Equation.

**(ii). Reaction with acidified potassium manganate (VII)**

Ethene is bubbled through acidified potassium manganate (VII) in a test tube.

Observation: The purple potassium manganate(VII) turns colourless.

Note: The reactions (i) and (ii) above are used as a test for Ethene (unsaturation)

USES OF ALKENES

Ethene

- Ethene is used to make polythene .
- Making cables.
- Ripening of Banana.
- Making Ethanol.

Propene is used to make plastics such as Perspex.

ALKYNES.

Alkynes are unsaturated hydrocarbons which contain one triple bond between two carbon atoms in its molecule. They have the general formula C_nH_{2n-2} where $n=2$ onwards.

Nomenclature

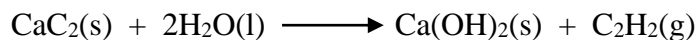
Alkynes are named after alkanes with the same number of carbon atoms by replacing the suffix-**ane** in the name of an alkane by **-yne** for the alkyne.

n	Formula	Name
2	C_2H_2	Ethyne
3.	C_3H_4	Propyne

PREPARATION OF ETHYNE.

Pour some sand into the flask and add some crystals of calcium carbide.

Drop cold water onto calcium carbide. Ethyne is evolved as a colourless gas and is collected over water by upward delivery.



Physical properties of ethyne

- It's a colourless, and odourless gas.

- It's slightly soluble in water.
- It has a sweet smell like that of garlic.
- It is less dense than air.

ALCOHOLS.

Alcohols are organic compounds which contain a hydroxyl group (-OH) joined to the hydrocarbon molecule. They have a general formula $C_nH_{2n+1}OH$.

Nomenclature.

Alcohols are named after corresponding alkanes by replacing the -e of the suffix -ane in alkanes with -ol.

n.	Formula	Name
1	CH_3OH	Methanol
2	C_2H_5OH	Ethanol
3	C_3H_7OH	Propanol

PREPARATION OF ETHANOL.

Ethanol is prepared by fermentation of glucose in the presence of yeast. Yeast contains an enzyme zymase which catalyses the decomposition of glucose into ethanol and carbondioxide. Fermentation is a process in which sugars are decomposed by enzymes into alcohol.



The ethanol obtained is fractionally distilled to concentrate it, and a distillate below 95°C is collected. The distillate is redistilled between 78 – 82°C. If required, dry, calcium oxide is added to remove all the water.

Manufacture of ethanol (locally)

1. From banana

The banana juice is extracted from ripe bananas by squeezing them using spear grass leaves. The juice is filtered to remove any solid impurities. The filtrate is then poured into a locally made wooden container where it is mixed with ground roasted sorghum. The container is covered and the mixture is allowed to ferment for two days. The resulting solution is crude ethanol, 'tonto'.

2. From millet

Millet is ground to form millet flour and the flour is mixed with water to form a paste which is covered in a container or buried in the ground for about 6days so that it can ferment. The fermented paste is removed and roasted to obtain malt. The malt is then dried in the sun. Dry germinated millet grains are ground and added to malt in a drum. A carefully determined amount

of water is then added to form a liquid mixture. The mixture is covered and stored in a warm place for 3 days to form an alcoholic drink, "marwa".

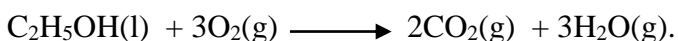
Physical properties of Ethanol.

- Ethanol is a colourless, mobile, hygroscopic liquid.
- It has a pleasant smell and burning taste.
- It has a density of 0.8g/cm^3 and mixes with water in any proportion.

CHEMICAL REACTIONS.

1 Combustion.

Ethanol burns in air with a blue flame to form carbon dioxide and water.



2 Dehydration.

Excess concentrated sulphuric acid at 170°C dehydrates ethanol and forms ethene.



Uses of ethanol

- It is used in the manufacture of perfumes, flavours, drugs and varnishes.
- It is used in thermometers measuring low temperatures.
- It is added to petrol and used as fuel.
- It is used in the manufacture of aftershaves and moisturising creams.
- It is used in the making of alcoholic drinks.

POLYMERIZATION

Polymerization is the joining together of small molecules to form a single large complex molecule. The small molecules that come together are called *monomers* and the large complex molecule is called *polymer*.

A polymer is a very large molecule with high relative molecular mass formed by the joining of simple molecules called monomers.

TYPES OF POLYMERIZATION

1. Addition polymerization

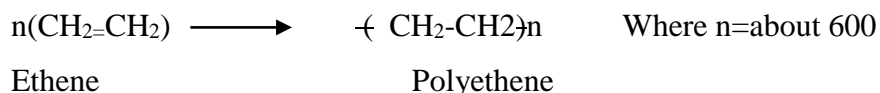
This type of polymerization involves the joining together of many small molecules to form a large molecule with no loss or gain of matter.

Examples of polymers formed by addition polymerization:

- Polyethene
- Polyvinyl chloride(PVC)
- Perspex
- Polystyrene

MANUFACTURE OF POLYETHENE

Ethene is heated to a temperature of 200⁰C under a pressure of 1500 atmospheres pressure with little oxygen as catalyst, white waxy solid polyethene forms.



PROPERTIES OF POLYETHENE

- It is a white waxy solid.
- It is less dense than water.
- It is tough and insoluble in all solvents at room temperature.
- It is a good electrical insulator.
- It does not rot, corrode or react with common chemicals.

USES OF POLYETHENE

- To make carrier bags, bottles, plates e.t.c.
- For making of over head and undersea telephone cables.
- For making of dustbins, toys, car carpets, bottles for milk.

2. Condensation polymerization:

This is the type of polymerization which involves the formation of a single large molecule by the combination of many small molecules and elimination of water.

Examples of polymers formed by condensation

1. Nylon

2. Bakelite
3. Terylene
4. Rubber.

TYPES OF POLYMERS

1. Natural polymers.

Natural polymers are polymers formed by condensation and exist in nature.

Examples and their uses

1. Starch – food for humans
 - Source of Ethanol by fermentation.
2. Silk – Making fabrics.
3. Cellulose – Making paper.
4. Cotton – making fabrics.
5. Rubber - making tyres, shoes.
6. Protein – human food.
7. Wool.

2 Synthetic polymers

Synthetic polymers are man – made polymers.

Examples and their uses

1. Polythene – manufacture of buckets, troughs, watering cans, pipes and bowls.
2. Polyvinyl chloride (PVC) – manufacture of insulating material used in wiring.
 - Making of car seat covers, hand bags and suitcase covers.
3. Perspex
 - making of wind screens for cars.
 - Making of windows and optical instruments.
4. Nylon
 - making fabrics, fishing lines and ropes.
5. Terylene
 - Making fabrics and conveyor belts.

VULCANISATION OF NATURAL RUBBER

Vulcanization of rubber is the process by which rubber is heated with sulphur. The sulphur atoms add to a few double bonds in rubber to form links which can no longer slide over each other easily making rubber hard.

Importance of vulcanization.

Vulcanization makes rubber strong, more elastic and heat resistant.

USES OF RUBBER

- Making of car tyres.
- Making of shoes.
- Making of gumboots.

PLASTICS.

Plastics are a group of materials produced from naturally occurring raw materials such as coal, petroleum and Silica. Plastics are materials that can easily be moulded.

TYPES OF PLASTICS

1 Thermo softening (thermo plastics)

Thermo softening plastics are plastics which soften when heated and harden when cooled.

Because they soften on warming, they can be moulded into different shapes.

Examples

- Polythene
- Perspex

2. Thermo setting plastics.

Thermosetting plastics are plastics which soften and decompose on heating.

Examples.

- Bakelite
- Commercial rubber.

Advantages and disadvantages of synthetic polymers over Natural polymers

Advantages

- They are cheap and attractive.
- Light and portable.
- Corrosion – resistant.

Disadvantages

- They pollute environment as they produce highly poisonous fumes when they burn.
- Some cannot be broken down by bacteria and other organisms (non biodegradable).

SOAP

Soap is sodium or potassium salt of long chain organic acid e.g stearic acid (NaSt). The chemical name of soap is sodium stearate.

MANUFACTURE OF SOAP.

Soap is manufactured by a process called *saponification*.

Oils or fats are heated by steam with sodium hydroxide solution. Sodium chloride is added to precipitate the soap.

The soap floats on top and can be removed and purified.

CLEANING ACTION OF SOAP.

Soap emulsifies grease into tiny droplets (micelles) in water. The polar end of soap dissolves in water and soap ions surround fat droplets. Non polar end of soap dissolves in grease that holds dirt on fabric and emulsifies it into a suspension. When washing is done, the oil is pulled off the fabric and floats free and dirt rinsed away by water.

Disadvantage of soap

Soap forms scum with hard water and a lot of it has to be used to produce sufficient lather for washing.

DETERGENTS: (Soap less soaps)

A detergent is any substance that facilitates the emulsification and removal of grease.

Soapless detergents are more effective than ordinary soap in hard water since they do not form scum.

Laboratory Preparation:

Concentrated sulphuric acid is added to vegetable oil in a test tube and the mixture stirred with a glass rod. The product is added to distilled water and stirred to remove excess acid. The solid left on decanting is washed with distilled water. Soapless detergents are manufactured in solid form (powders) or liquid form (washing liquids and shampoos). Some detergents contain phosphates to prevent scum formation.

Advantages of detergents

- They can be used with hard water without scum formation.
- They are better cleansing agents than ordinary soap.
- More economical than ordinary soap when used in laundry work.

Disadvantage

- Branched chain detergents are non biodegradable and pollute rivers.

Exercise

- (i) State what is meant by the term soap?
 - (ii) Briefly describe how a dry sample of soap flakes can be prepared in the laboratory.
 - A water sample X was boiled and on cooling was tested by shaking it with a known volume of soap solution. A greasy scum and a cloudy solution were observed.
 - (i) State what is meant by scum.
 - (ii) Write an ionic equation for the reaction leading to the formation of scum.
 - (i) Name one compound that can be used to stop scum formation in water sample X.
 - (ii) Write equation for the reaction that would take place when the compound named in c (i) is used.
 - Describe how soap can remove grease from fabrics when washed.
- Name the products of :
 - (i) Complete combustion of alkanes.
 - (ii) Incomplete combustion of alkanes.
 - Why is the incomplete combustion of alkanes dangerous to human beings?
 - Write the equation for the incomplete combustion of methane.
- (i) Explain what is meant by the term polymerization?
 - (ii) Name two naturally occurring polymers and two synthetic polymers.
 - Distinguish between a thermoplastic and a thermosetting plastic.
 - Explain the term cracking.
- (i) State the difference between fats and oils.
 - (ii) Give one example of each.
 - State what would be observed if soap solution was shaken with a solution containing calcium hydrogen carbonate.
 - Explain your answer in (c) .
 - State what would be observed if Soapless detergent was used in (c) .
 - (i) Give one disadvantage of Soapless detergents.
 - (ii) State the disadvantages of Soapless detergents.

4.(a) Beer or crude ethanol is manufactured by the process known as fermentation.

- (i) Explain what is meant by the term fermentation.
- (ii) Write equation for the reaction that takes place during fermentation.
- (iii) State whether the process of fermentation is exothermic or endothermic.
- (b) Describe briefly how in homes alcoholic drinks can be prepared from either ripe bananas or millet.
- (c) Draw a diagram of the apparatus that can be used to concentrate the alcohol produced in (b).

5. Glucose, $C_6H_{12}O_6$, can be converted to ethanol by a catalytic reaction caused by an enzyme produced from yeast.

(a) Name

- (i) the reaction in which yeast converts glucose to ethanol.
- (ii) the enzyme produced by yeast during the reaction.
- (b) Write the equation leading to the formation of ethanol.
- (c) Briefly describe how the ethanol produced can be concentrated.
- (d) State two uses of ethanol.

6.(a) Write the structural formula of

- (i) ethene
- (ii) Ethane
- (b) Name one reagent which can be used to distinguish between ethene and ethane.
- (c) State what would be observed if ethene was treated with the reagent you have named in(b).
- (d) Write equation for the polymerization of ethene.

7. Some carbon compounds form polymers.

- (a)(i) What is meant by the term polymer?
- (ii) Distinguish between synthetic polymer and natural polymer. Categorize silk and nylon as synthetic and natural polymer.
- (iii) State one use of each of the polymers.
- (b) State the disadvantages and advantages of synthetic polymers.
- (c) Name the product of polymerization of ethene and write equation leading to formation of the product.

8.(a) Ethene was bubbled into bromine.

- (i) State what was observed.
- (ii) Write an equation for the reaction that took place.

- (b) Name the products formed when ethene burns in a limited supply of oxygen.
- (c) State one environmental hazard associated with the products of the reaction in(b).

9.(a) Explain the following observations:

- (i) Water containing calcium hydrogen carbonate will not lather easily with soap unless the water is boiled prior to using soap.
- (ii) Water containing magnesium sulphate will not lather with soap even after boiling the water.
- (b) A gaseous hydrocarbon Y contains 20% of hydrogen by mass. 7.5g of Y occupy 5.6dm³ at s.t.p[1mole of gas at s.t.p occupies 22.4dm³].
 - (i) Calculate the empirical formula of Y.
 - (ii) Determine the molecular formula of Y.
- (c) Write
 - (i) The name of hydrocarbon Y
 - (ii) The structural formula of Y

10.(a) In the manufacture of soap, oil or fat is heated with sodium hydroxide solution. Saturated sodium chloride is then added.

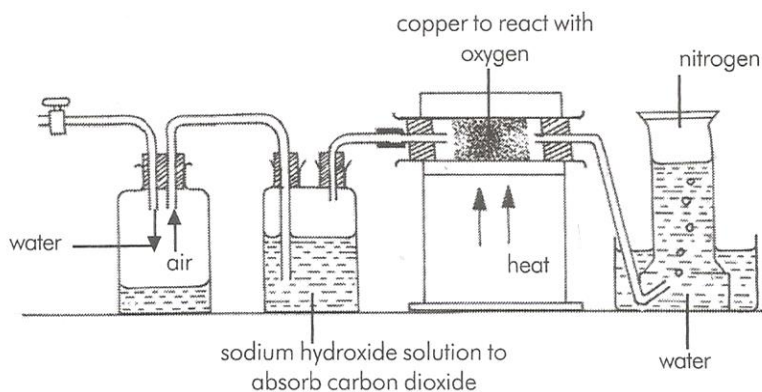
- (i) Name the process of making soap.
- (ii) What is the purpose of adding saturated sodium chloride solution?
- (iii) State the chemical nature of soap.
- (b) Explain what is meant by addition polymerization.

NITROGEN AND ITS COMPOUNDS.

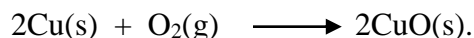
Nitrogen is the first member of group V of the periodic table. The atomic number of nitrogen is 7 and its simplest electronic configuration is 2:5. Nitrogen exhibits a covalency of three (since it requires three electrons to complete the outer most shell). Nitrogen gas occupies about 78% by volume of air.

Preparation of Nitrogen.

(a) From air

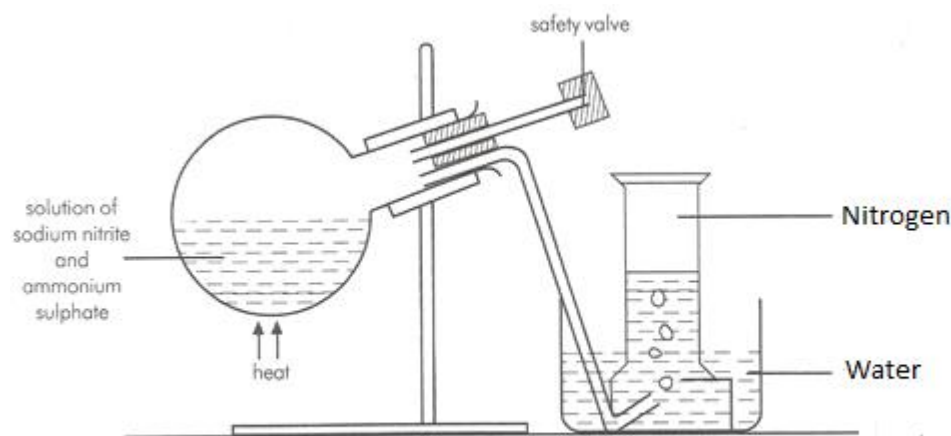


Tap water is used to displace the air upwards. Air is passed through concentrated solution of caustic soda (potassium hydroxide) in a wash bottle which absorbs all the carbon dioxide in it. The remaining gases, oxygen and nitrogen are passed over heated copper in a combustion tube which removes oxygen.



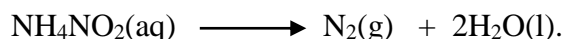
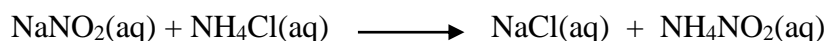
The remaining nitrogen is then collected over water.

(b) Laboratory preparation from ammonium nitrite.



A mixture of sodium nitrite (14g) and ammonium chloride (11g) are placed in a round bottom flask and water added (350cm³) and heated gently. As the solution becomes warm, rapid effervescence occurs and nitrogen evolved is collected over water.

This mixture forms ammonium nitrite which decomposes on heating to form nitrogen gas.



(c) Industrial preparation of Nitrogen.

Nitrogen is prepared on an industrial scale by fractional distillation of liquid air. Air is first compressed to about 200 atmospheres pressure, cooled, and allowed to escape from a small jet. Expansion cools the air further. The cooled air is allowed to flow by passing through tubes containing incoming compressed air. Successive coolings finally liquefy the air. Upon distillation of liquid air, Nitrogen (lower boiling point of -196°C) is evolved first leaving oxygen (higher boiling point, -183°C).

PROPERTIES OF NITROGEN

(a) Physical properties.

- (i). It's a colourless, odourless gas.
- (ii) It's insoluble in water at ordinary temperature.

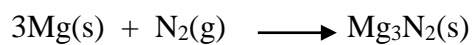
- (iii) It's slightly less dense than air.
- (iv) It's neutral to litmus.

(b) Chemical reactions of Nitrogen

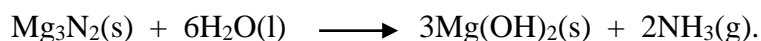
Nitrogen is unreactive (inert) under ordinary conditions due to the strong triple bond between its atoms which needs a lot of energy to break before the nitrogen atoms are free to react. However, it reacts under special conditions.

1. *Reaction with magnesium.*

Magnesium burns in nitrogen when heated to form magnesium nitride, a white solid.

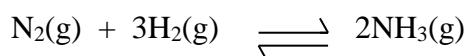


When water is added to magnesium nitride, ammonia is given off which turns red litmus to blue and magnesium hydroxide also forms.



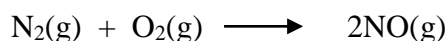
2. *Reaction with Hydrogen*

Nitrogen reacts with hydrogen when heated together in the presence of iron catalyst to form ammonia. This reaction is used in the Haber process for the manufacture of ammonia on an industrial scale.



3. *Reaction with oxygen*

Nitrogen reacts with oxygen in the presence of lighting or when sparked to form Nitrogen monoxide..



USES OF NITROGEN

- In the manufacture of ammonia by Haber process.
- It's used with argon to fill electric light bulbs.

- Liquid nitrogen is used as a refrigerant to preserve semen and biological specimens.
- Its essential for the growth of plants.

OXIDES OF NITROGEN

There are three main oxides of Nitrogen;

1. Dinitrogen oxide, N_2O
2. Nitrogen monoxide, NO
3. Nitrogen dioxide, NO_2

NITROGEN MONOXIDE

Nitrogen monoxide is prepared from moderately concentrated (50%) nitric acid and copper metal.

Moderately concentrated nitric acid is added on to copper turnings in a flask. Vigorous effervescence occurs and nitrogen monoxide evolved is collected over water because it is insoluble in water. The flask fills with brown fumes. This is because nitrogen monoxide produced reacts with air in the flask to form brown fumes of nitrogen dioxide. A blue solution of copper (II) nitrate is formed.



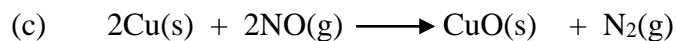
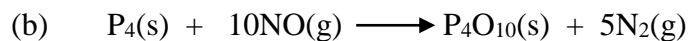
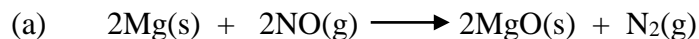
Properties of Nitrogen monoxide

Physical properties

1. it's colourless and odourless.
2. it's slightly soluble in water.
3. It's slightly denser than air (vapour density of NO is 15 and that of air is 14.4).
4. it's neutral to litmus.

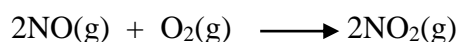
Chemical reactions.

1. It oxidizes Magnesium, Phosphorus and copper to their oxides and itself reduced to nitrogen gas.



Test for Nitrogen monoxide.

When exposed to air, it forms reddish brown fumes of Nitrogen dioxide.



Use: It's used in the manufacture of nitric acid.

NITROGEN DIOXIDE.

- (a). Its best prepared by heating lead(II) nitrate.

Put some Lead (II) nitrate crystals into a dry test tube and heat the crystals gently and then strongly. **Diag**

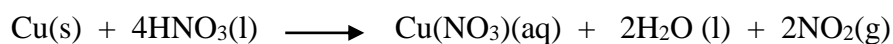
Observation.

Lead (II) nitrate crystals crackle, and melt to an orange liquid. It effervesces giving a reddish-brown gas. Reddish - brown solid when hot and yellow when cold remains in the tube.

Nitrogen dioxide is condensed in ice-salt mixture as a green liquid in the U-tube. The liquid in the U-tube is dinitrogen tetra oxide.



- (b) Nitrogen dioxide can also be prepared from copper and concentrated nitric acid in the cold.

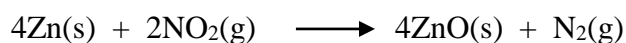
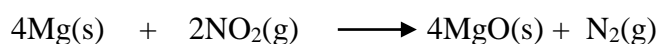


Physical properties of Nitrogen dioxide

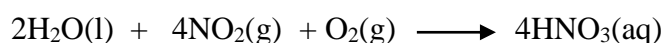
- Its reddish brown gas which condenses to a green liquid.
- It has pungent, irritating smell.
- It's poisonous and produces nitric acid which turns blue litmus to red when dissolved in water.
- Its soluble in water.
- Its denser than air .

CHEMICAL REACTIONS

- 1 It oxidizes red hot metals and itself reduced to nitrogen gas.



- 2 It reacts with water in the presence of excess air to form nitric acid as the only product.



3. It neutralizes alkalis forming a mixture of a nitrate, nitrite and water



4. It dissolves in water forming two acids, nitrous acid and nitric acid.

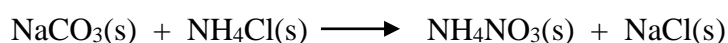


Use: It is used in the manufacture of nitric acid.

DINITROGEN OXIDE

It's prepared by heating sodium nitrate and ammonium chloride. This produces ammonium nitrate.

On heating ammonium nitrate, it melts and effervesces giving off dinitrogen oxide and steam.

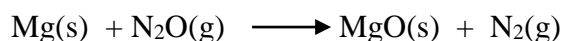


Physical properties.

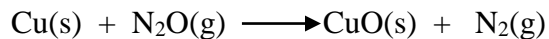
- Colourless gas with faint sweet sickly smell.
- Its neutral to litmus.
- Its slightly soluble in cold water.
- It's denser than air.
- It relights a glowing splint.

CHEMICAL REACTION

- 1 Burning magnesium decomposes the gas and continues to burn in it.



- 2 It oxidizes red hot copper and itself reduced to nitrogen gas.



TEST FOR DINITROGEN OXIDE.

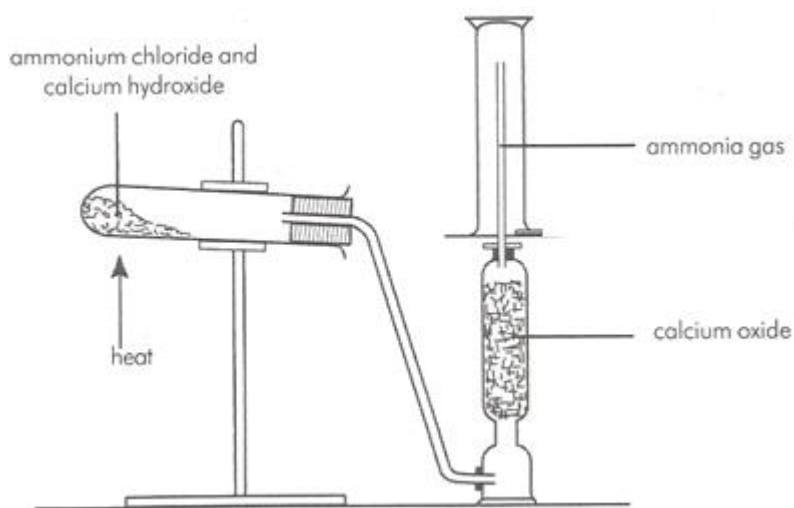
- 1 It relights a glowing splint.
- 2 Has faint sweet sickly smell. This test distinguishes it from oxygen.

Uses:

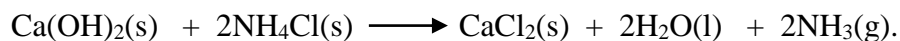
- 1 A mixture of dinitrogen oxide and oxygen is used as an anaesthetic in minor surgeries such as extraction of teeth.
- 2 Makes patients who breathe it to laugh hysterically hence the name “laughing gas”.

AMMONIA, NH₃:

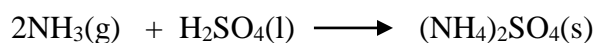
Laboratory preparation of ammonia



A ground mixture of calcium hydroxide and ammonium chloride is heated in a round flask. Ammonia gas is evolved and passed through calcium oxide to dry it. The dry ammonia is then collected by upward delivery (downward displacement of air) because it's less dense than air.



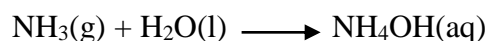
Ammonia gas is not dried using concentrated sulphuric acid because it reacts with sulphuric acid to form ammonium sulphate.



It also reacts with calcium chloride to form solid complex compounds.

Physical properties of Ammonia

- 1 Appearance: it's a colourless gas.
- 2 Density: It's less dense than air.
- 3 Solubility: It's highly soluble in water to form ammonia solution.



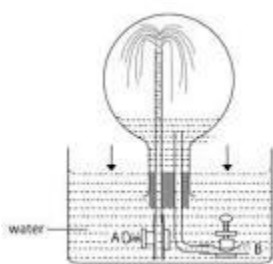
- 4 Smell: It has a pungent choking smell.

- 5 Effect on litmus: It turns red litmus paper to blue. It's the only alkaline gas.

SOLUBILITY OF AMMONIA.

The high solubility of ammonia is illustrated in the fountain experiment.

Fountain experiment.



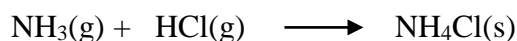
- Fill a large round bottom flask with dry ammonia.
- Fit the flask with rubber stopper carrying tubes fitted with clips. Place the mouth of the flask in water in a beaker.
- Open the clip attached at the end of along glass tube for a moment, close it and allow some water to run down into the round part of flask.
- As the ammonia in the flask dissolves, air pressure forces water up the tube and it enters as a fountain.

If little red litmus solution is added to the water in the trough, the litmus solution turns blue indicating that ammonia dissolves in water to form alkaline solution.

CHEMICAL REACTIONS OF AMMONIA

1 *Reaction with hydrogen chloride*

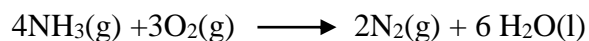
When a gas jar of ammonia is inverted over a gas jar of hydrogen chloride and the gases allowed to mix, dense white fumes are seen which settle as a white solid ammonium chloride on the sides of the gas jar.



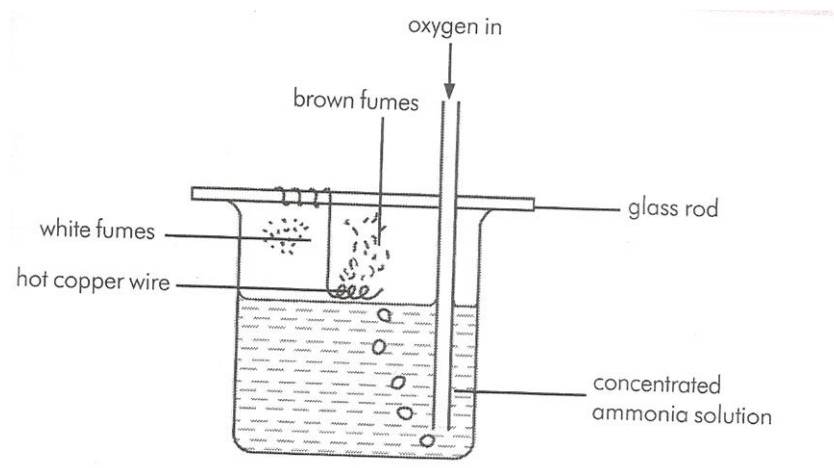
Reaction with oxygen (Combustion)

(i) In the absence of a catalyst

Ammonia burns with a green-yellow flame in air slightly enriched by oxygen to form nitrogen and water.



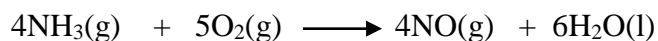
(ii) **In the presence of a catalyst**



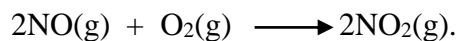
Ammonia gas in the presence of copper catalyst is oxidized to nitrogen monoxide and steam. The hot (platinum) or copper wire glows and brown fumes are evolved.

Explanation

Platinum (copper) glows because of the catalytic oxidation of ammonia by oxygen. The catalyst is platinum or copper



The nitrogen monoxide formed is oxidized to brown fumes of Nitrogen dioxide.



3 Ammonia as a reducing agent.

Reaction with heated Copper (II) Oxide.

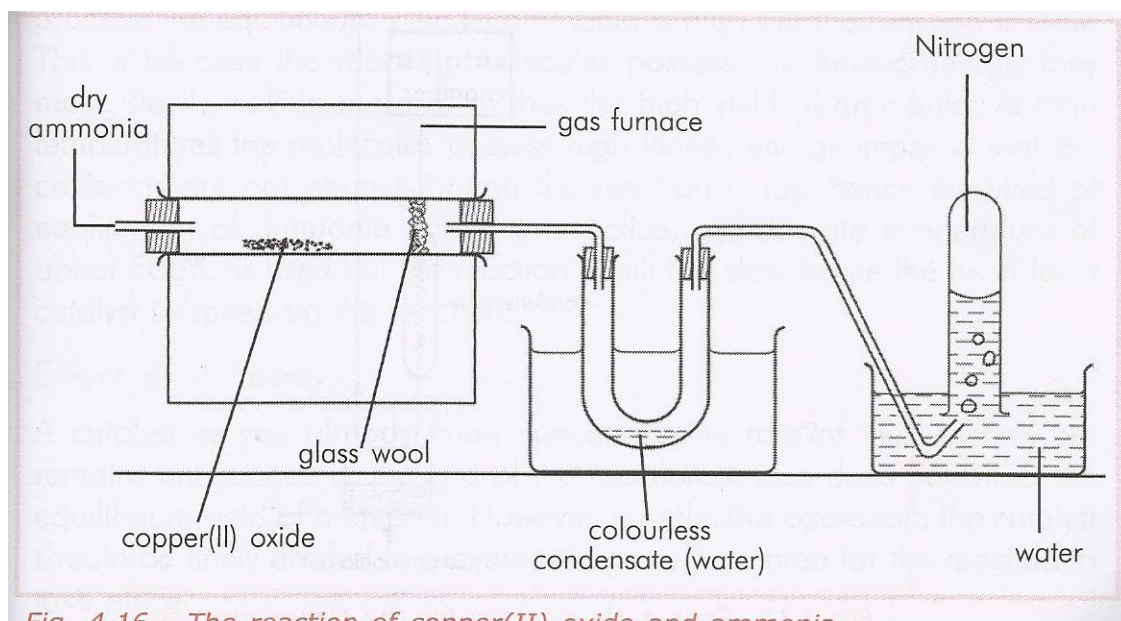


Fig. 4.16 The reaction of copper(II) oxide and ammonia

When ammonia gas is passed over heated copper (II) Oxide in a combustion tube, the black copper (II) oxide is reduced to a brown copper metal and itself oxidized to Water and nitrogen.

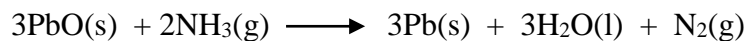


The water formed in the U-tube can be tested using anhydrous copper (II) sulphate which turns from white to blue.

Lead (II) Oxide can be used instead of copper (II) oxide.

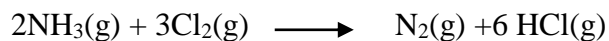
Observation:

When ammonia is passed over heated lead (II) oxide, the yellow solid turns to a grey solid.



4. Reaction with chlorine

Ammonia burns spontaneously in chlorine forming a mist of hydrogen chloride.



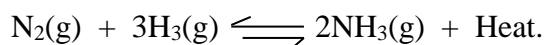
MANUFACTURE OF AMMONIA (HABER PROCESS)

Raw materials

- Nitrogen
- Hydrogen
- Finely divided iron catalyst

Process

- The gases nitrogen and hydrogen are cleaned (purified) to remove CO₂, SO₂ and water vapour which would poison the catalyst.
- Nitrogen gas is passed together with hydrogen over heated catalyst of finely divided iron impregnated with aluminium oxide at a temperature of 450°C and pressure 250 atmospheres.



- The ammonia produced is liquefied by cooling and compression.

Conditions for maximum yield of ammonia

- Temperature of about 450°C.
- High pressure of about 200 atmospheres.
- There must be a catalyst (finely divided iron usually, 'promoted' by alumina).

USES OF AMMONIA

1. In the manufacture of nitric acid.
2. In the manufacture of fertilizers e.g. ammonium sulphate.
3. Ammonia solution is used in laundry work for softening water.
4. Liquid ammonia is used in large- scale refrigerating plants, such as in ships and ware houses.
5. It's used in the production of nylon, plastics and glues.

Test for ammonia

Smell: It has a characteristic choking smell.

Action of litmus: Expose damp litmus paper to the gas. It is turned blue.

Effect of heat on ammonium salts

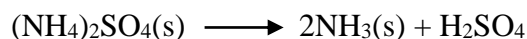
1. **Ammonium chloride**: Place a little ammonium chloride in a dry test tube and heat gently below the solid. It sublimes and a white sublimate of ammonium chloride is formed on the cooler upper part of the tube.



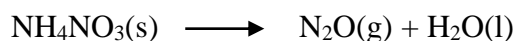
2. **Ammonium carbonate** decomposes to form ammonia, carbon dioxide and water.



3. **Ammonium sulphate** decomposes on heating forming ammonia and sulphuric acid. Ammonia is more volatile than sulphuric acid and escapes before the two can recombine.



4. **Ammonium nitrate** decomposes to form dinitrogen oxide and water.



5. **Ammonium nitrite** decomposes to form nitrogen and water.



FERTILIZERS

Fertilizers supply the required inorganic elements for the plants growth. Such important elements include mainly potassium, nitrogen and phosphorus.

Nitrogenous fertilizers such as sulphate of ammonia or ammonium nitrate are added to the soil to improve its fertility. The nitrate is the more useful compound because the sulphate can cause acidity of the soil. Phosphates are important fertilizers because they stimulate early growth of roots therefore early growth of crops. Potassium is needed by plants in order to absorb nitrogen, to produce sugar, starch, and to increase resistance to drought and disease. Potassium chloride and potassium sulphate (potash) are common fertilizers used.

Some fertilizers with their compositions

Fertilizer	Abbreviation	Composition
Nitrogen phosphorus,	NPK	17:17:17

potassium		
Calcium ammonium nitrate	CAN	27:0:0
Diammonium phosphate	DAP	18:52:0
Sulphate of ammonia	SOA	20:0:0

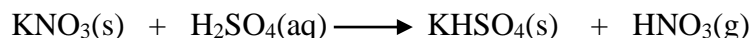
NITRIC ACID, HNO_3

Nitric acid is one of the common mineral acids which attacks most substances, including almost all the metals.

Laboratory preparation of nitric acid

Diag

Potassium nitrate crystals are placed in the bulb of the retort and concentrated sulphuric added through a thistle funnel. The retort is then heated gently. Potassium nitrate gradually dissolves and effervescence occurs. The brown fumes are condensed and collected in water cooled receiver as a yellow liquid.



Conditions for this reaction

- Sulphuric acid must be concentrated
- Gentle heating

Nitric acid prepared in the laboratory is yellow because it contains dissolved nitrogen dioxide. Pure nitric acid is colourless. The impurity can be removed by bubbling air through it.

INDUSTRIAL MANUFACTURE OF NITRIC ACID

Raw materials

- Air
- Ammonia
- Water

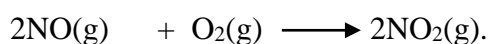
Nitric acid is manufactured through 3 stages:

- 1 By catalytic oxidation of ammonia.

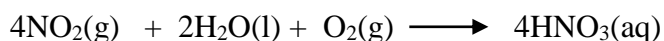
Ammonia is oxidized by passing it with excess air over platinum/ rhodium gauze catalyst to form nitrogen monoxide.



- 2 The nitrogen monoxide formed is rapidly cooled and combines with oxygen from excess air to form nitrogen dioxide.



- 3 Nitrogen dioxide in the presence of excess air is dissolved in hot water to form nitric acid.



REACTIONS OF NITRIC ACID

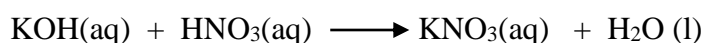
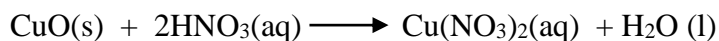
Nitric acid is a colourless, fuming liquid of density 1.5gcm^{-3} .

- (a) **As an acid**

Nitric acid is a very strong acid; it's almost completely ionized in dilute solution.



- 1 Nitric acid neutralizes bases forming metallic nitrates and water



- 2 Nitric acid liberates carbon dioxide in reactions with metallic carbonates.



- 0 Dilute nitric acid(1%) react with the more electropositive metals , liberating hydrogen



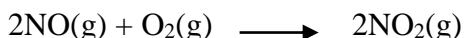
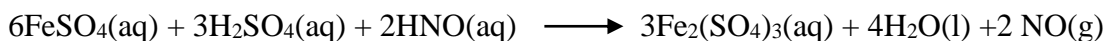
(a) As an oxidizing agent

1 *Reaction with iron (II) sulphate*

Iron (II) sulphate crystals are dissolved in dilute nitric acid. Little concentrated nitric acid is added and heated.

Observation

It oxidizes green iron (II) sulphate solution to a yellow solution of iron (III) sulphate and itself reduced to brown fumes of nitrogen dioxide .



Ionic equation

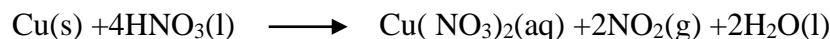


2 *Reaction with copper metal*

Concentrated nitric acid is added to copper turnings in a test tube.

Observation:

Vigorous effervescence occurs and dense brown fumes are given off and a green solution is formed. The test tube becomes hot as the reaction is exothermic.



Dilute nitric acid reacts with copper metal forming nitrogen monoxide and copper (II) nitrate solution.



3 *Reaction with hydrogen sulphide*

Hydrogen sulphide is passed through a fairly concentrated solution of nitric acid.

Observation: A yellow precipitate of sulphur is formed.



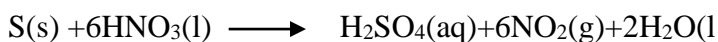
When concentrated nitric acid is used, a yellow precipitate is formed and brown fumes given off.



4 ***Reaction with non metals***

(i) Sulphur

Concentrated nitric acid oxidizes sulphur to sulphuric acid and itself reduced to brown nitrogen dioxide.



(ii) Carbon

It oxidizes carbon to carbondioxide and itself reduced to brown nitrogen dioxide



Uses of Nitric acid

- (i) Nitric acid is used in the manufacture of explosives such as trinitrotoluene (TNT).
- (ii) It is used in the manufacture of fertilizers such as sodium nitrate and ammonium nitrate.
- (iii) It is used in the manufacture of drugs.
- (iv) It is mixed with cellulose (straw or cotton) to form nitrocellulose and a solution of this pigment is used as cellulose paints.

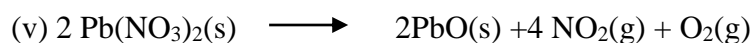
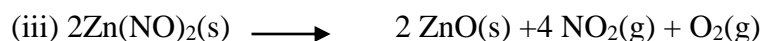
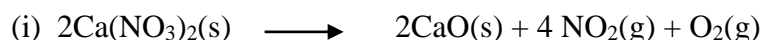
ACTION OF HEAT ON NITRATES

Nitrates decompose when heated. The products of decomposition depend on the position of the metal in the reactivity series.

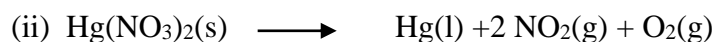
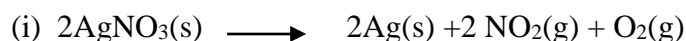
1. Nitrates of sodium and potassium melt when heated and then slowly decompose to form a pale yellow nitrite and a colourless oxygen gas.



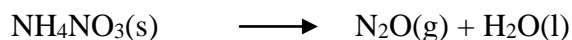
2. Nitrates of magnesium, Zinc, Calcium, Copper, lead decompose to form a metal oxide, nitrogen dioxide and oxygen.



3. Nitrates of mercury and silver decompose to form a metal, nitrogen dioxide and oxygen.



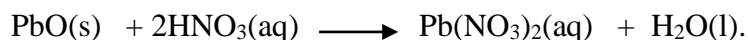
4. Ammonium nitrate decomposes to form dinitrogen oxide and water.



Preparation of Lead (II) nitrate starting from lead (II) oxide.

- Warm dilute nitric acid in a beaker.
- Add Lead (II) oxide to the acid until no more dissolves.
- Filter the excess lead (II) oxide and collect a colourless filtrate of lead (II) nitrate solution.

- Heat the filtrate to evaporate the water until crystals begin to form. Dip a glass rod in the beaker to see if crystals form on it.
- Allow the contents to cool and crystallize.
- The crystals are washed with little distilled water and dried between filter paper.



TEST FOR NITRATES IN SOLUTION

Brown ring test.

To the solution of the nitrate in a test tube, add fresh prepared iron (II) sulphate solution followed by concentrated sulphuric acid slowly down the side of the test tube.

Observation: A brown ring is formed where the two layers meet.

Exercise

1. (a). Lead (II) nitrate was heated until there was no further change
 - (i). what was observed
 - (ii). Write equation of the reaction that took place.
- (b). Dilute hydrochloric acid was added to a solution of lead (II) nitrate and the resultant solution warmed.
 - (i). State what was observed
 - (ii). What can you deduce from your observation.
2. (a) (i) Write equation to show how ammonia can be prepared from calcium hydroxide.
 - (ii) Name one substance that can be used to dry ammonia.
- (b) Ammonia was passed over heated copper (II) oxide.
 - (i) State what was observed.
 - (ii) Write equation for the reaction that took place.
3. Nitric acid is manufactured by catalytic oxidation of ammonia.

- (a) Name
- (i) two raw materials, other than ammonia that are used in the manufacture of nitric acid.
- (ii) the catalyst used
- (b) Write equation for the reaction between nitric acid and ammonia.
- (c) State one use of the product in (b).
4. (a) Draw a labelled diagram of apparatus that can be used to prepare ammonia in the laboratory.
- (b) Describe an experiment that can be carried out to show that ammonia is a soluble alkaline gas.
- (c) A copper foil was heated strongly and held over a concentrated solution of ammonia in a beaker. Oxygen was bubbled into the ammonia solution.
- (i) State what was observed.
- (ii) Explain the observation in c(i).
- (c) Ammonia reacts with lead (II) oxide according to the equation.
- $$2\text{NH}_3(\text{g}) + 3\text{PbO}(\text{s}) \longrightarrow \text{N}_2(\text{g}) + 3\text{Pb}(\text{s}) + 3\text{H}_2\text{O}(\text{l})$$
- Calculate the volume of ammonia at room temperature that would be required to completely react with 2.50g of lead (II) oxide. [Pb=207, O=16, 1 Mole of gas at room temperature occupies 24 dm³].
5. (a) A piece of burning magnesium was introduced into a jar of nitrogen.
- (i) State what was observed.
- (ii) Write an equation for the reaction which took place.
- (b) Water was added to the product of the reaction in (a) and the resultant mixture tested with litmus. State what was observed.
- (c) Name one other metal that reacts with nitrogen in a similar way to magnesium.
6. (a) (i) Name the raw materials used for the manufacture of ammonia.

- (iii) Write equation for the reaction leading to the formation of ammonia.
- (b) Explain how the formation of ammonia is affected by
- (i) Pressure
- (ii) Temperature
- (c) State another factor that affects the formation of ammonia.
- (d) Dry ammonia was passed over heated copper (II) oxide until there was no further change. State what was observed and explain your answer.
7. (a) Write equation to show the reaction between copper(II) hydroxide and dilute nitric acid.
- (b) Aqueous ammonia was added drop wise to a sample of copper(II) hydroxide until ammonia was in excess.
- (i) State what was observed.
- (ii) Write the formula of the final product.
- (c) Write equation for the reaction when copper(II) hydroxide is heated.
8. (a) Lead(II) nitrate was heated until there was no further change.
- (i) State what was observed.
- (ii) What do you deduce from your observation?
- (b) Dilute hydrochloric acid was added to a solution of lead(II) nitrate and the resultant solution warmed.
- (c) Describe how ammonia can be converted to nitric acid. Use equations to illustrate your answer.
9. (a) Starting from lead(II) oxide, describe how a dry sample of lead(II) nitrate can be prepared.
- (b) Describe what happens when lead(II) nitrate crystals are heated.

- (c) What would be observed and write equation for the reaction in each case when the following are heated:
- (i) Potassium nitrate.
 - (ii) Silver nitrate
 - (iii) Zinc nitrate.
10. (a) Describe how you would prepare nitric acid in the laboratory.
- (b) Write equation for the reaction of dilute nitric acid with the following:
- (i) Copper (II) oxide
 - (ii) Sodium carbonate
 - (iv) Magnesium
- (c) How does concentrated nitric acid react with:
- (i) Copper
 - (ii) Sulphur
 - (iii) Carbon
12. (a) What is a fertilizer?
- (b) Give three most common fertilizers useful to farmers.
- (c) Ammonium sulphate is fertilizer that increases acidity in soils. Calculate the percentage of nitrogen in ammonium sulphate(N=14,S=32,H=1,O=16).
13. (a) Describe how lead(II) nitrate crystals can be prepared starting from lead oxide.
- (b) Describe how you can show by chemical test that nitric acid contain a nitrate ion.
- (c) Write equation for the reaction between concentrated nitric acid and sulphur.

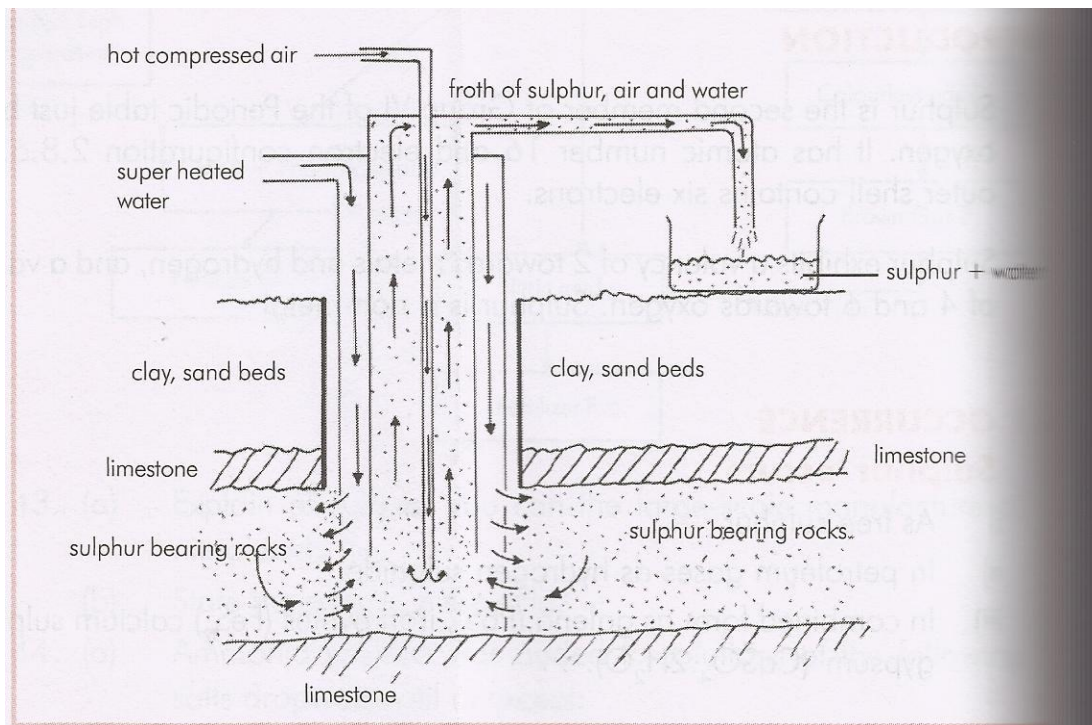
SULPHUR AND ITS COMPOUNDS

Sulphur is in group VI of the periodic table, with atomic number 16 and electronic configuration 2:8:6. It is a non metal with six electrons in the outer most shell.

Occurrence

- As free sulphur in underground deposits.
- Zinc blend (ZnS).
- In combined form as galena (PbS), iron pyrite (FeS_2), gypsum ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$).

Extraction of sulphur



Sulphur is extracted by the Frasch process.

In the Frasch process, the Frasch pump consisting of three concentric pipes are sunk deep to the sulphur deposit.

Super heated water at 170°C and under high pressure, 10 atmospheres is forced down the outer pipe and melts the sulphur.

Hot air compressed air (about 15 atmosphere) is forced down the narrow central pipe (smallest) pipe to force molten sulphur up the surface.

The molten sulphur and water are forced up the middle pipe and is led off to settling tanks where sulphur solidifies and separates from water.

ALLOTROPY.

Allotropy is the existence of an element in two or more different forms but in the same physical state. The different forms of the same element are called allotropes.

Allotropes of sulphur.

1 Rhombic sulphur (octahedral sulphur)

It's octahedral in structure and melts at 114°C. When heated, it is converted to monoclinic sulphur at 96°C. 96°C is the transition temperature for both rhombic and monoclinic sulphur. It has a density of 2.06gcm⁻³.

Preparation of rhombic sulphur

Shake some powdered sulphur with carbon disulphide in a test tube for some time.

Filter the contents in a dry beaker through dry filter paper and funnel. Tie the filter paper over the mouth of a beaker and pierce a few holes in it and leave to stand. The carbon disulphide evaporates leaving crystals of sulphur.

2 Monoclinic sulphur

It has a needle or pencil like structure. It has a melting point of 119°C. It has a density of 1.98gcm⁻³.

Preparation monoclinic.

Sulphur is heated slowly in an evaporating dish until all of it melts. The liquid is allowed to cool until a crust just form. The crust is pierced in 2 places and molten sulphur poured away. When the crust is lifted, brown needle-like crystals of monoclinic sulphur are seen.

Differences between Rhombic and monoclinic

Rhombic

- melting point 114°C
- Stable at temperatures below 96°C
- Density of 2.06g/cm³
- Octahedral in shape
- Yellow translucent crystals

Monoclinic

- melting point 119°C
- unstable at temperatures below 96°C
- Density is 1.98g/cm³
- needle shaped
- transparent brown crystals.

PROPERTIES OF SULPHUR

1 It's a yellow solid which is insoluble in water but soluble in organic solvents e.g carbon disulphide and methyl benzene.

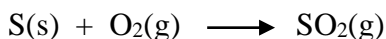
2 Combustion.

(a). When sulphur is heated in absence of air the following changes occur:

- It melts at about 115°C into amber coloured mobile liquid.
- It becomes darker in colour and suddenly at 160°C becomes very viscous.
- It gradually becomes more mobile again and very dark reddish brown in colour.
- It boils at 444°C giving off light brown sulphur vapour.

(b) Combustion of sulphur in plenty of air

Sulphur burns in air with a blue flame to form sulphur dioxide.

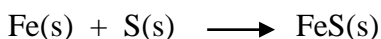


3 Sulphur reacts with metals when heated to form metal sulphide.

Examples

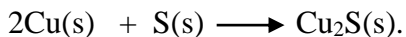
- (a) Mix iron fillings and sulphur in a glass tube and heat the mixture

Observation: The mixture glows and a black residue of iron (II) sulphide is left.



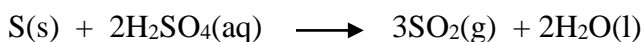
- (b) Mix copper and sulphur in a test tube and heat the mixture.

Observation: The copper foil or wire glows forming copper (I) sulphide.



4 *Action of acids on sulphur.*

Dilute acids do not act on sulphur. Its oxidized by hot concentrated sulphuric acid to sulphur dioxide.

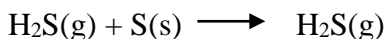


Sulphur is oxidized by hot concentrated nitric acid to sulphuric acid. Bromine can be used as a catalyst.



5. *Reaction with hydrogen gas*

Hydrogen gas is bubbled through molten sulphur, it forms hydrogen sulphide.



Uses of sulphur

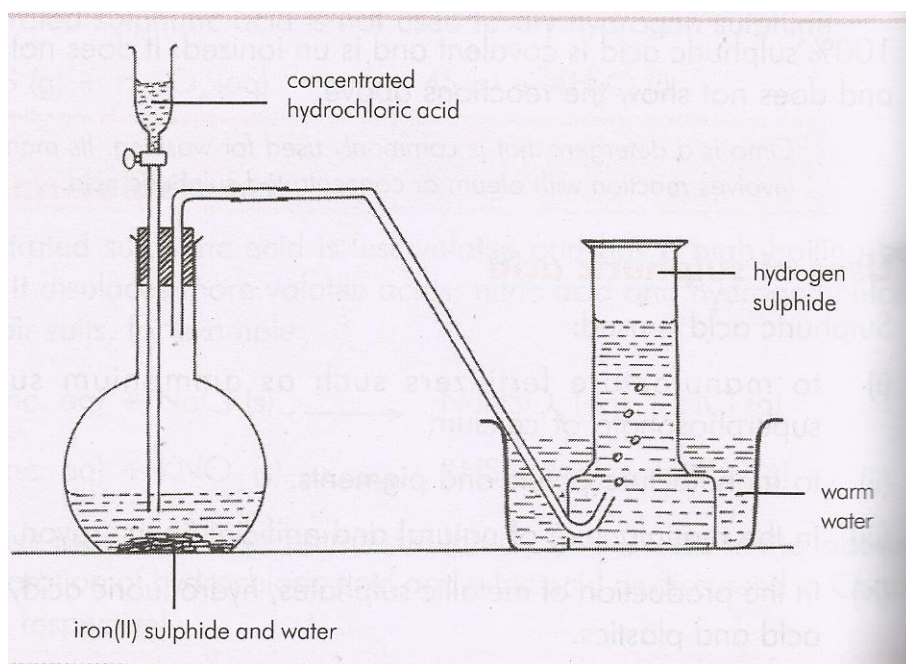
1. For the manufacture of sulphuric acid.
2. In making calcium hydrogen sulphite which is used as a bleacher of wool pulp in the manufacture of paper.
3. For vulcanization of rubber, a process that converts soft rubber into a hard, tough substance.
4. In the manufacture of dyes and fire works materials.

5. In the making of ointments and drugs.
6. For dusting vines and other trees to prevent the growth of fungus.

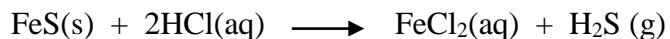
COMPOUNDS OF SULPHUR

HYDROGEN SULPHIDE, H_2S

Laboratory preparation.



Place some solid iron (II) sulphide in a flask and add some concentrated hydrochloric acid through a thistle funnel. Effervescence occurs and hydrogen sulphide gas produced is collected over warm water because it's soluble in cold water.



Note. Preparation of H_2S should be in a fume cupboard or outside the laboratory because it's very poisonous.

The gas can be dried by passing it through anhydrous calcium chloride and collected by downward delivery.

Test for Hydrogen sulphide

1. It turns moist Lead (II) ethanoate paper from white to black.
2. By distinct smell of rotten eggs.

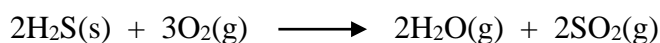
Physical properties

- It's a colourless gas with rotten egg smell.
- It's denser than air.
- It's fairly soluble in water.
- It's highly poisonous.

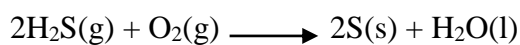
Chemical reactions

1 Combustion

Hydrogen sulphide burns with a blue flame in plenty of air to form sulphur dioxide and steam.



In limited supply of air, hydrogen sulphide burns to form a yellow deposit of sulphur.



2 As a reducing agent.

(a) *Reaction with Nitric acid*

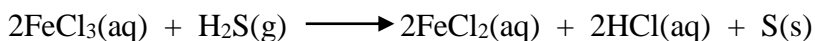
When hydrogen sulphide is bubbled through concentrated nitric acid,

Brown fumes of nitrogen dioxide are given off and a yellow deposit of sulphur. It reduces nitric acid to nitrogen dioxide and itself oxidized to sulphur.

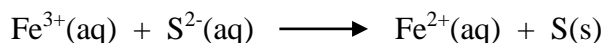


(b) *Reaction with Iron (III) chloride solution.*

Hydrogen sulphide reduces the yellow iron (III) chloride solution to a green iron (II) chloride solution and itself oxidized to hydrogen chloride, a yellow deposit of sulphur appears.



Ionic equation

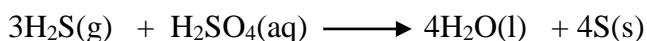


With acidified potassium permanganate – it reduces purple permanganate to colourless solution, and a yellow deposit of sulphur appears.

With acidified potassium dichromate – It turns potassium dichromate solution from orange to green leaving a yellow residue of sulphur.

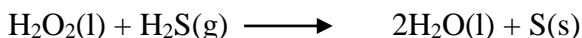
(c) ***Reaction with concentrated sulphuric acid***

Hydrogen sulphide reduces concentrated sulphuric acid to sulphur.



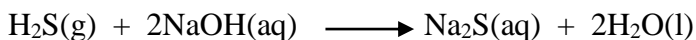
(b). ***Reaction with hydrogen peroxide***

Hydrogen sulphide reduces hydrogen peroxide to water and yellow precipitate of sulphur is formed.



3 ***As an acid***

Hydrogen sulphide acts as a weak dibasic acid it forms a normal salt and acid salt with sodium hydroxide solution.



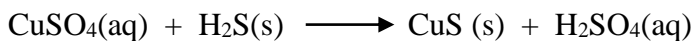
In excess H_2S sodium hydrogen sulphide is formed.



4 ***Reaction with metal salt solutions***

(a) ***Reaction with copper (II) sulphate***

When hydrogen sulphide is bubbled through a boiling solution of copper(II) sulphate, a dark brown precipitate is formed.



(b) When hydrogen sulphide is bubbled through a hot solution of lead (II) nitrate,

A black precipitate is formed.



Exercise

1.(a) Concentrated hydrochloric acid was added to iron sulphide in a test tube.

- (i) State what was observed.
 - (ii) Write equation for the reaction.
- (c) Hydrogen sulphide was burnt in limited air.
- (i) State what was observed.
 - (ii) Write the equation for the reaction.

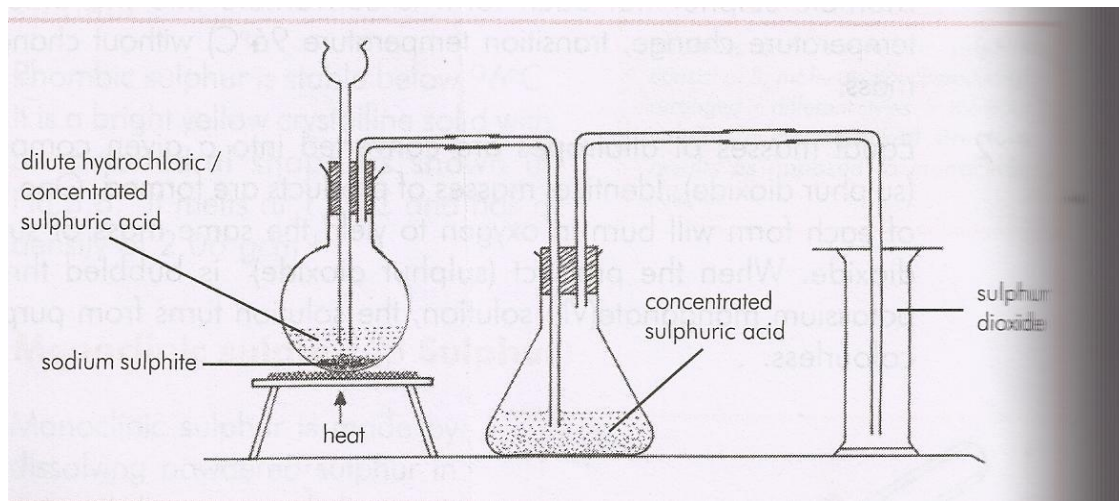
2.State what is observed and write equations for the reaction that takes place when the hydrogen sulphide is added to :

- (i) Iron(III) chloride solution.
- (ii) Concentrated sulphuric acid.
- (iii) Copper(II) sulphate solution.
- (iv) Lead(II) nitrate solution.

SULPHUR DIOXIDE, SO₂

Laboratory preparation

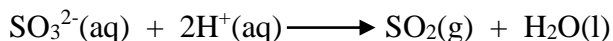
(a) From sodium sulphite



Concentrated sulphuric acid is added to sodium sulphite crystals in a flask and the mixture gently warmed. Sulphur dioxide is evolved and passed through concentrated sulphuric acid in a wash bottle to dry it. The gas is collected by down ward delivery because it's denser than air.



Ionic equation



(b) From copper turnings

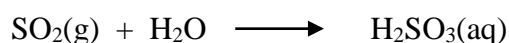
Copper turnings are slid into the flask and concentrated sulphuric added from the funnel.

The mixture is heated gently, effervescence occurs and sulphur dioxide is evolved. The gas is passed through concentrated sulphuric acid to dry it, and then collected by down ward delivery since it's denser than air.



Physical properties of sulphur dioxide

- it's a colourless gas with a characteristic irritating smell.
- It's fairly poisonous with an irritating smell.
- It's about 2 ½ times denser than air.
- It's soluble in water to form sulphurous acid.



Test for sulphur dioxide

- 1 By bubbling sulphur dioxide thru acidified potassium dichromate solution.

Observation.

The orange colour of potassium dichromate turns to green.

- 2 By bubbling sulphur dioxide through a solution of acidified potassium manganate (VII)

Observation

The purple colour of potassium manganate (VII) turns colourless.

SULPHUR DIOXIDE AS A REDUCING AGENT

Sulphur dioxide is a powerful reducing agent. In the presence of water, it reduces by supplying electrons to substances or removing oxygen from them.

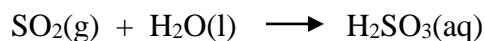
(a). *Bleaching action of sulphur dioxide*

When moist red or blue flowers are added to a gas jar of sulphur dioxide, the blue or red flowers are turned colourless due to reduction.

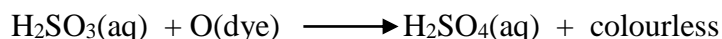
Explanation:

The bleaching is due to sulphurous acid formed by the reaction between water on moist flowers and sulphur dioxide.

The bleaching is due to sulphurous acid formed by the reaction between water on moist flowers and sulphur dioxide.



The sulphurous acid removes oxygen from the dye forming sulphuric acid.



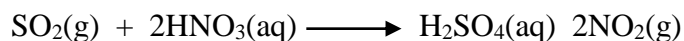
Note: This is bleaching by reduction, and the bleached compound when exposed to air, regains its colour. This explains browning of old news papers.

(b) ***Reaction with concentrated nitric acid***

When sulphur dioxide is bubbled through concentrated nitric acid in a boiling tube,

Brown fumes are evolved and the liquid becomes warm because the reaction is exothermic.

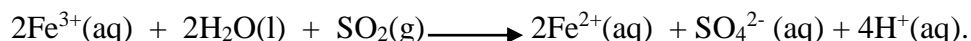
Sulphur dioxide reduces nitric acid to brown nitrogen dioxide and itself oxidized to sulphuric acid.



(c) ***Reaction with Iron (III) sulphate solution.***

When sulphur dioxide is bubbled through a brown solution of iron (III) sulphate, the brownish colour of the solution turns to pale green.

Sulphur dioxide reduces brown iron (III) sulphate to green iron (II) sulphate and itself oxidized to sulphuric acid.



(d) ***Reaction with potassium dichromate solution***

Acidified potassium dichromate solution is turned from orange to green. The green colour is due to the formation of chromium (III) sulphate.

(e) ***Reaction with potassium manganate (VII) solution.***

The purple potassium manganate (VII) solution turns colourless.

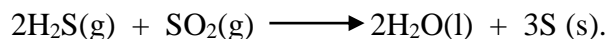
SULPHUR DIOXIDE AS AN OXIDISING AGENT

(a) *Reaction with hydrogen sulphide*

Add to a gas jar of sulphur dioxide little water and invert over a gas jar of hydrogen sulphide.

Observation

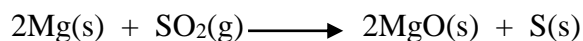
A yellow deposit of sulphur is produced. The dry gases do not react.



Note: H_2S is a more powerful reducer than Sulphur dioxide and takes up oxygen causing Sulphur to act as an oxidizer.

(b) *Reaction with magnesium*

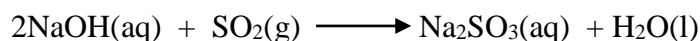
Magnesium continues to burn for a short time in a gas jar of sulphur dioxide to form a white powder of magnesium oxide and a yellow solid of sulphur.



Acidic Nature of Sulphur dioxide.

1. Sulphur dioxide being acidic in nature is absorbed by sodium hydroxide solution.

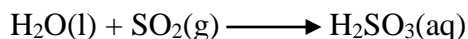
Sulphur dioxide dissolves in sodium hydroxide solution to form sodium sulphite



When sulphur dioxide is bubbled in excess, sodium hydrogen sulphite is formed.



2. Sulphur dioxide turns moist blue litmus paper red. It dissolves in water to form sulphurous acid.



Uses of sulphur dioxide.

1. It's mainly used in the manufacture of sulphuric acid.

2. It's used as a bleaching agent for wool, Silk, straw and sponges.
3. Its used as a preservative to food stuffs and fruit squashes eg orange juice.
4. As a fumigant for houses since the gas is poisonous, it can be used to kill germs in houses, clothing and grain.
5. For making calcium hydrogen sulphite used in the manufacture of paper.
6. Destruction of termites or white ants.

SULPHUR TRIOXIDE, SO₃

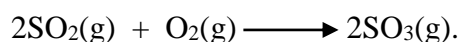
Sulphur trioxide is a white hygroscopic solid.

Preparation.

Its prepared by passing a mixture of dry sulphur dioxide and dry oxygen over heated plantinised asbestos (vanadium (V) oxide) as catalyst. The best temperature is 450 – 500°C.

Observation

Dense white fumes form and condense to white silky needles in a freezing mixture of ice and sodium chloride.



SULPHURIC ACID, H₂SO₄

Manufacture of sulphuric acid

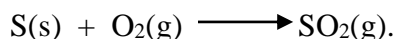
Sulphuric acid is manufactured by the contact process.

Raw materials

- 1 Sulphur dioxide
- 2 Air
- 3 Vanadium (v) oxide catalyst

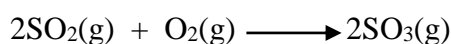
Process

- 1 Sulphur is burnt in air to form sulphur dioxide.

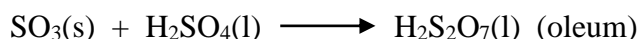


- 2 Sulphur dioxide is cleaned thoroughly so that the catalyst is not poisoned by dust and impurities.

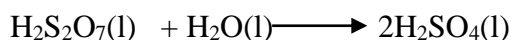
- 3 Sulphur dioxide is compressed and burnt in oxygen in the presence of vanadium (v) oxide catalyst to form sulphur trioxide.



- 4 Sulphur trioxide is dissolved in concentrated sulphuric acid to form oleum(fuming sulphuric acid).



- 5 Oleum is diluted with water to form concentrated sulphuric acid.



The conditions for the yield of SO₃ sulphuric acid.

- 1 increasing pressure since forward reaction occurs with decrease in volume
- 2 Increasing temperature (450°C) to increase rate of reaction
- 3 Use of vanadium (V) oxide catalyst to speed up the reaction.

PROPERTIES OF SULPHURIC ACID

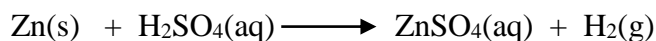
Sulphuric acid is a dense oily liquid.

(a) Dilute sulphuric acid – as an acid

Dilute sulphuric acid is a dibasic acid and is completely ionized in aqueous solution.



- (i). Reacts with metals to liberate hydrogen gas.



- (ii) Reacts with carbonates and hydrogen carbonates to liberate carbondioxide.



- (iii) Reacts with bases to form salt and water only.



- (b) **Concentrated sulphuric acid as an oxidizing agent**

Hot concentrated sulphuric acid is an oxidizing agent. It accepts electrons from reducing agents such as metals and non-metals.

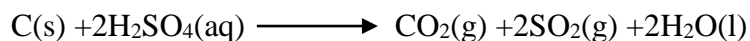
- (i) ***Reaction with copper metal***

Hot concentrated sulphuric acid oxidizes copper to a blue copper(II) sulphate and itself reduced to sulphur dioxide.



- (ii) ***Reaction with carbon (charcoal)***

Hot concentrated sulphuric acid oxidizes carbon to carbon dioxide and itself reduced to sulphur dioxide.



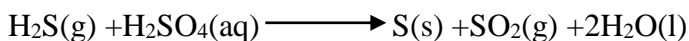
- (iii) ***Reaction with sulphur***

Hot concentrated sulphuric acid oxidizes sulphur to sulphur dioxide.



- (iv) ***Reaction with hydrogen sulphide***

Concentrated sulphuric acid oxidizes hydrogen sulphide to yellow sulphur and itself reduced to sulphur dioxide.



(c) **Concentrated sulphuric acid as a dehydrating agent**

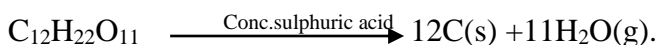
Concentrated sulphuric acid is a strong dehydrating agent. It can remove water or elements of water from many compounds.

(i) ***Dehydration of sugar***

Concentrated sulphuric acid is added to sugar covered with water in a beaker.

Observation

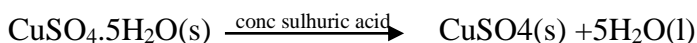
A black mass of charcoal forms filling the beaker. Steam is given off and the whole beaker becomes hot as the reaction is exothermic.



(ii) ***Dehydration of hydrated copper (II) sulphate crystals***

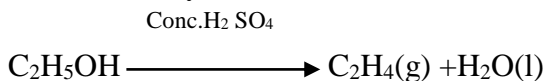
Copper (II) sulphate crystals are warmed with concentrated sulphuric acid.

Observation: The blue crystals change to a white powder of anhydrous copper(II) sulphate.



(iii) ***Dehydration of ethanol***

When concentrated sulphuric acid is heated with ethanol to a temperature of 170⁰C, ethanol is dehydrated to form ethene.

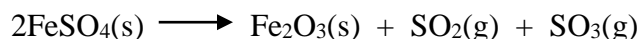


EFFECTS OF HEAT ON SULPHATES

- (a) Iron (II) sulphate, $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$ on gentle heating, it loses it's water of crystallization changing from green to yellow anhydrous solid.

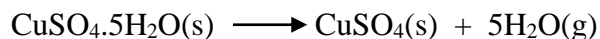


On strongly heating, it gives off SO_2 , white fumes of SO_3 , leaving reddish brown solid, Fe_2O_3 .

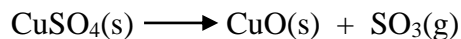


(b) Copper (II) sulphate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.

When heated, it loses water of crystallization and leaves a white anhydrous copper(II) sulphate.



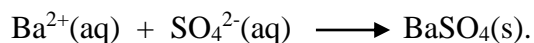
On strong heating, a black solid residue remains and white fumes are given off.



TEST FOR SULPHATES AND SULPHURIC ACID

To a solution containing sulphate (or sulphuric acid) is added Barium nitrate followed by dilute nitric acid.

Observation: A white precipitate is formed.



The white precipitate is due to the formation of insoluble barium sulphate.

Uses of sulphuric acid.

1. In the manufacture of fertilizers e.g Ammonium sulphate and super phosphate of calcium.
2. In the manufacture of paints and pigments.
3. In the manufacture of detergents and soap.
4. As an electrolyte in accumulators (Car batteries).
5. In the production of plastics.
6. In extraction of metals and pickling of metals.
7. In refining of petroleum oils.

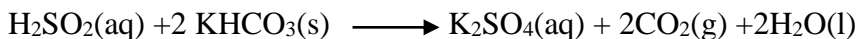
Exercise

- 1.(a) Describe how sulphur is extracted by the Frasch process.
- (b) Write equations to show how fuming sulphuric acid can be obtained from sulphur.
- (c) State what would be observed if concentrated sulphuric acid is added to sugar.
- 2.(a) Draw a labelled diagram to show how a dry sample of sulphur dioxide can be prepared in the laboratory.
- (b) Write the equation for the reaction that takes place in (a).
- (c) Describe a test that can be carried out to confirm the presence of sulphur dioxide.
- (d) Excess sulphur dioxide was bubbled through a solution of sodium hydroxide.

Write equation for the reaction that took place.

- (a) 25.0cm^3 of 0.1M sodium hydrogen carbonate solution reacted completely with 27.80cm^3 of sulphuric acid. Calculate the concentration of sulphuric acid in moles per litre.
3. (a) State what would be observed when
 - (i)Crystals of iron (II) sulphate, $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$, are heated strongly. Write equations for the reactions that occur.
 - (ii) Concentrated nitric acid is added to a solution of iron(II)sulphate.
- (f) Describe one test in each case that can be carried out to show that iron(II) sulphate crystals contain
 - (i) Water of crystallization
 - (ii) Iron(ii) ions
 - (iii) Sulphate ions.
- 0 (a) Sulphuric acid is a strong dibasic acid.
 - (i) Explain the terms strong acid and basicity.
 - (ii) Write an equation to show how sulphuric acid ionizes in water.

(b) Sulphuric acid reacts with potassium hydrogen carbonate according to the equation;



Calculate the volume of carbon dioxide produced at s.t.p when 20.0cm^3 of 0.5M sulphuric acid is reacted with excess potassium hydrogen carbonate.

(c) (i) State the conditions under which sulphuric acid reacts with potassium nitrate.

(ii) State what is observed and write an equation for the reaction.

(g) Explain what happens when a mixture of sulphur and concentrated nitric acid is heated.

1 Sulphuric acid reacts with ethanol to produce a colourless gas that decolourizes bromine water.

(a) Name the gas that is produced.

(b) State the conditions for the reaction.

(c) Write equation for the reaction.

(d) Give two large scale uses of the gas produced.

2 (a)(i) Name one substance that is reacted with hydrochloric acid to produce sulphur dioxide in the laboratory.

(ii) State the conditions for the reaction.

(iii) Name the substance that can be used to dry sulphur dioxide formed.

(iv) Write the equation for the reaction leading to the formation of sulphur dioxide.

(b) State what would be observed and explain what would happen if sulphur dioxide is passed through a solution containing

(i) Acidified potassium dichromate solution

(ii) A dye

(c) Briefly describe how sulphur dioxide can be converted to sulphuric acid. Your description should include equations and conditions for the reactions.

(e) Sulphur dioxide was passed into a beaker containing a red flower and water.

(i) State what was observed

(ii) Give a reason for your answer.

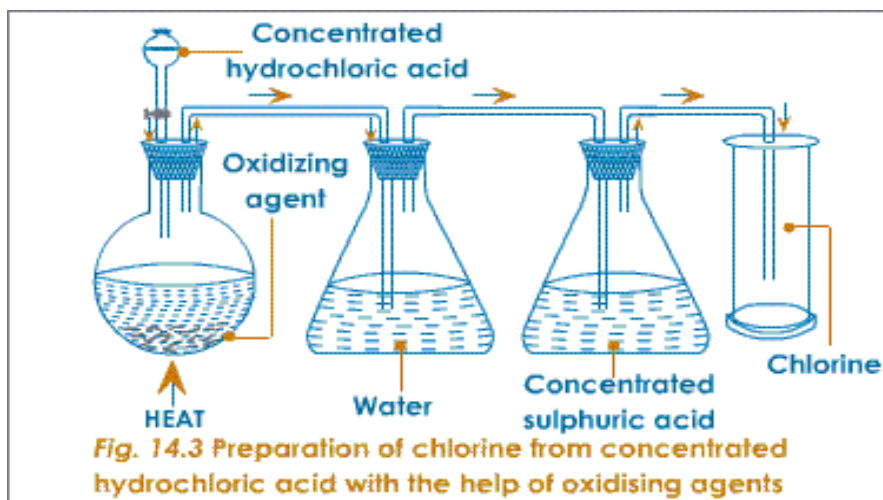
CHLORINE AND ITS COMPOUNDS

Chlorine is an element with atomic number 17. In the periodic table of elements it belongs to the halogens of group VII and has atomic mass 35.5. Chlorine has 17 protons, 17 electrons and 18 neutrons. Chlorine has a symbol Cl. The gas has a green yellow colour and is poisonous.

Laboratory preparation of chlorine

Chlorine gas can be prepared by oxidation of concentrated hydrochloric acid using:

(a) Manganese (IV) oxide.



A mixture of manganese (IV) oxide and concentrated hydrochloric acid in a flask is heated gently. A greenish- yellow gas is evolved. The gas is passed through a wash bottle containing water to remove hydrogen chloride fumes and then through concentrated sulphuric acid in a wash bottle to dry it.

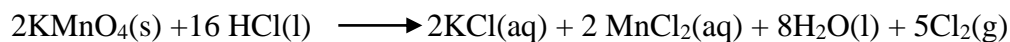
It is collected by down ward delivery because it is denser than air.



(b) Potassium manganate (VII)

Solid potassium manganate (VII) is placed in a flask and concentrated hydrochloric acid added.

The reaction occurs in the cold and a greenish –yellow gas is evolved. The gas is passed through a wash bottle containing water to remove hydrogen chloride fumes, then through concentrated sulphuric acid in a wash bottle to dry it.

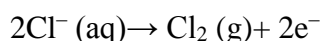


Industrial manufacture of chlorine

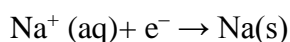
Chlorine is manufactured on an industrial scale by electrolysis of concentrated sodium chloride (Brine) in mercury cathode cell.

Reaction at electrodes

The reaction at anode (+) is:



The reaction at the mercury cathode is:



The sodium metal formed by this reaction dissolves in the mercury to form sodium amalgam. The sodium amalgam drops in water to form sodium hydroxide solution, hydrogen and mercury.

The sodium hydroxide solution is evaporated to dryness and the molten substance solidifies as pellets.

Physical properties of chlorine

- It is a greenish-yellow gas.
- It has a choking, unpleasant, irritating smell.
- It is very poisonous if inhaled to even a small extent.
- It is denser than air.
- It is fairly soluble in water forming hydrochloric acid and Hypochlorous acids.
- It bleaches damp litmus.

Chemical reactions of chlorine

1. Reaction with water

Chlorine reacts with water, forming chlorine water (a mixture of hydrochloric acid and Hypochlorous acid).



Effect of sun light on chlorine water

Pass chlorine gas into water in a beaker for some time until the water becomes yellow- green in colour. Fill a long tube with the chlorine water. Invert it in a beaker containing some of the chlorine water and expose the apparatus to bright sunlight.

Observation

After sometime, the gas collects in the boiling tube and when a glowing splint is applied, the gas relights a glowing splint showing it is oxygen.

- Formation of Hypochlorous acid and hydrochloric acid

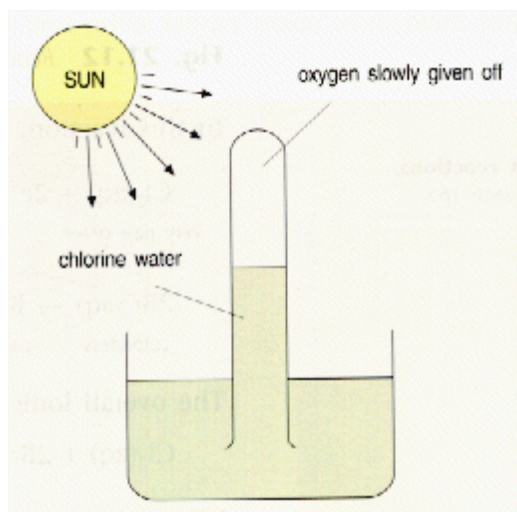


- Decomposition of Hypochlorous acid

Hypochlorous acid is unstable to heat (sun light) and decomposes to liberate oxygen gas.



Overall equation: $2\text{H}_2\text{O}(\text{l}) + 2\text{Cl}_2(\text{g}) \longrightarrow 4\text{HCl}(\text{aq}) + \text{O}_2(\text{g})$

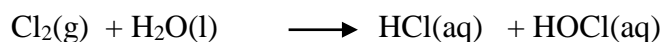


2. Bleaching action of chlorine

Pour a little litmus solution into a gas jar of chlorine.

Observation: The litmus solution immediately turns colourless.

Chlorine gas reacts with water forming hydrochloric acid and Hypochlorous acid.



The Hypochlorous acid is very reactive and readily gives up its oxygen to the dye to form a colourless compound.



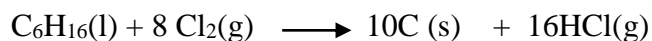
Note: Chlorine bleaches only in the presence of water by oxidation. Dry chlorine does not bleach.

3. Reaction with turpentine (C_6H_{16})

A piece of filter paper is dipped in warm turpentine and then dropped in a gas jar of chlorine.

Observation:

A black solid particle of carbon is formed. Chlorine removes hydrogen from turpentine leaving black carbon behind and hydrogen chloride.



3. Reaction with reducing agents

(i) *Reaction with hydrogen sulphide*

When a gas jar of hydrogen sulphide is inverted over a gas jar of chlorine, chlorine reacts with it to form a yellow precipitate of sulphur and white fumes of hydrogen chloride. Chlorine oxidizes sulphide ion in hydrogen sulphide to form yellow sulphur and hydrogen chloride.



Chlorine is reduced and hydrogen sulphide is oxidized to sulphur.

(ii) ***Reaction with iron(II) chloride solution***

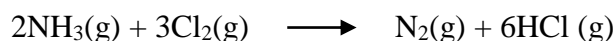
When a stream of chlorine is bubbled through a pale green solution of iron(II) chloride, the solution changes to yellow. Chlorine oxidizes iron (II) ions to iron (III) ions.



Ionic equation



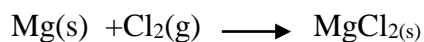
(iii) Ammonia is reduced by chlorine to form nitrogen gas and hydrogen chloride.



4. Reaction with metals

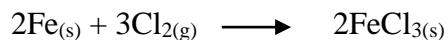
(i) ***Sodium and Magnesium***

These metals burn in chlorine forming white solids of sodium chloride and magnesium chloride.

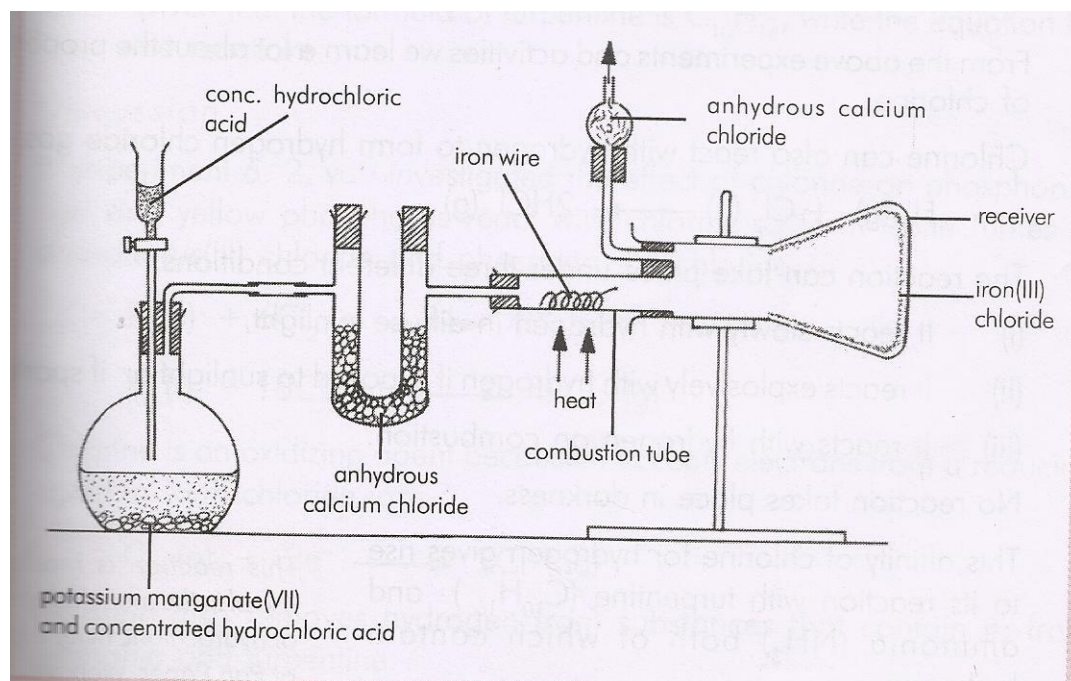


(ii) ***Reaction with iron wire***

A stream of dry chlorine is passed over heated coil of iron wire, it glows forming black crystals of iron (III) chloride .The crystals should be removed and placed in a desiccator as they are very deliquescent.



A tube containing anhydrous calcium chloride is connected to the bottle to prevent water from the atmosphere from entering the bottle as this would be absorbed by iron (III) chloride which is very deliquescent.

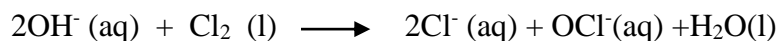


5. Action of chlorine on alkalis

Bubble chlorine gas slowly through cold dilute sodium hydroxide solution, the pale yellow solution of sodium chloride and sodium chlorate (I) is formed.



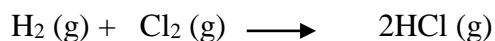
Ionic equation



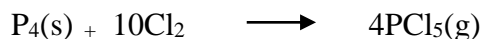
6. Reaction with non-metals

Chlorine reacts with non-metals to form chlorides.

- (i) Chlorine reacts with hydrogen in bright sunlight to form steamy fumes of hydrogen chloride. The greenish-yellow colour of chlorine gradually disappears.



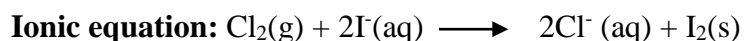
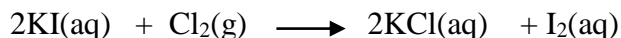
- (ii) Chlorine reacts with yellow phosphorous spontaneously in a gas jar to form white fumes of chlorides of phosphorous.



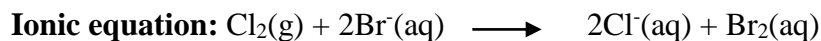
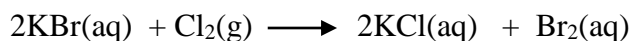
6. Displacement reaction of chlorine

Chlorine is higher in the reactivity series than bromine and iodine and therefore can displace them from solutions of their salts in water.

When chlorine is bubbled through a colourless solution of potassium iodide, the solution turns brown due iodine given off.



When chlorine is bubbled into a solution of potassium bromide, the colourless solution immediately turns red due to the formation of bromine water.



Test for chlorine

Chlorine is a greenish –yellow gas which rapidly bleaches damp litmus paper.

Since the gas is acidic, if damp litmus paper is used a red colour is often seen before it is bleached.

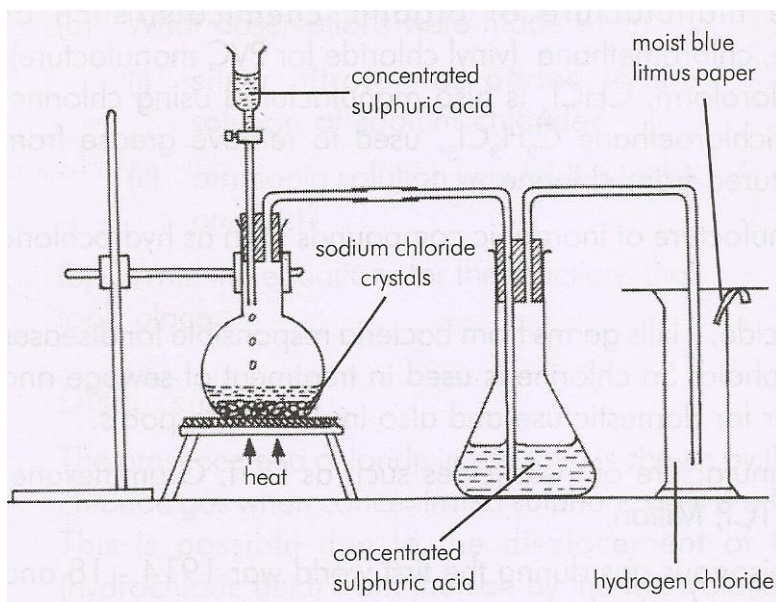
Uses of chlorine

- In the manufacture of hydrochloric acid.
- In the treatment of water to kill germs.
- In the manufacture of plastics such as polyvinyl chloride (PVC).

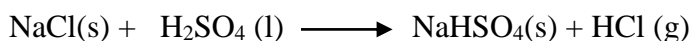
- In the manufacture of bleaching agents used in paper industry to bleach wood pulp.
- In the manufacture of weed killers.

HYDROGEN CHLORIDE

Preparation of hydrogen chloride gas

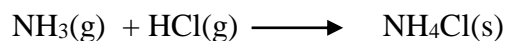


Sodium chloride is placed in a flask and concentrated sulphuric acid is added through a dropping funnel. Effervescence occurs and misty fumes are observed. The gas is passed through a wash-bottle containing concentrated sulphuric acid in a wash bottle to dry it, and collected by downward delivery, the gas being denser than air.

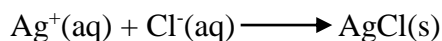


Test for hydrogen chloride

- It produces dense white fumes in the presence of ammonia.



- It produces a white precipitate of silver chloride in a drop of a solution of silver nitrate and nitric acid which is held on a glass rod in the gas.



- It forms misty fumes in moist air.

Physical properties of hydrogen chloride

- It is a colourless gas that fumes in moist or damp air.
- It turns damp blue litmus paper red showing it is acidic.
- It has a choking, irritating smell.
- It is very soluble in water. This high solubility can be demonstrated by the fountain experiment.
- It is denser than air.

Aqueous solution of hydrogen chloride is called hydrochloric acid.

HYDROCHLORIC ACID

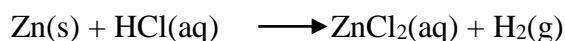
Hydrochloric acid is prepared by dissolving hydrogen gas in water using inverted funnel.

The filter funnel is used in passing hydrogen chloride into water in order to prevent the ‘sucking back’ of water from the beaker into the reaction vessel. If the gas is passed into water until no more gas is absorbed, the product is concentrated hydrochloric acid.

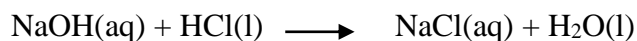
Diag

Properties of hydrochloric acid

1. It turns blue litmus paper red.
2. It reacts with certain metals (Mg, Zn, Fe) to liberate hydrogen.



3. It neutralizes bases to form salts and water only.



4. It liberates carbon dioxide with carbonates.



Uses of hydrogen chloride

1. In the manufacture of hydrochloric acid. The gas is dissolved in water.

2. In the manufacture of ammonium chloride. This is used in dry batteries as electrolyte.

Uses of hydrochloric acid

1. In pickling of metal surfaces to remove the surface oxide, rust, before the metal is galvanized.
2. To extract glucose from bones during the manufacture of glue and gelatin.
3. In the manufacture of plastics such as PVC (Polyvinylchloride).
4. It is used in removal (de-scaling) of rust from iron before galvanizing.
5. It is used in the manufacture of chlorine.

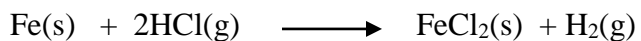
Properties of hydrogen chloride in water and in methylbenzene

Reaction	Solution of hydrogen chloride in water	Solution of hydrogen chloride in methylbenzene
Action on blue litmus paper	It turns blue litmus red	It does not react with litmus
Reaction with magnesium metal	Liberates hydrogen gas with metals	Does not form hydrogen with metals
Reaction with carbonates and hydrogen carbonates	Liberates carbon dioxide gas	Does not react with carbonates
Conduction of electricity.	Conducts electricity	Does not conduct electricity
Reaction with bases	Reacts to form a salt and water	Does not react with bases

Reactions of hydrogen chloride

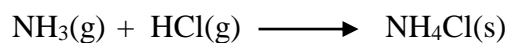
1. *Reaction with iron wire*

When dry hydrogen chloride is passed over heated iron wire, a white solid of iron (II) chloride is formed.



2. *Reaction with ammonia*

Hydrogen chloride reacts with ammonia to form white fumes of ammonium chloride.



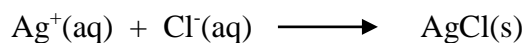
Test for chloride ion in solution

To a solution of a chloride, add dilute nitric acid followed by silver nitrate solution.

Observation

A white precipitate is formed. The white precipitate is due to the formation of insoluble silver chloride .

Equation



Exercise

1. (a) Describe how a pure dry sample of chlorine can be prepared in the laboratory starting from potassium manganate(VII). *(No diagram is required, you must include conditions and equations).*

(b) State what would be observed and write equations for the reactions that would occur if
 - (i) Chlorine was bubbled into an aqueous sodium hydroxide.
 - (ii) Burning magnesium was lowered into a jar of dry chlorine.
 - (iii) Chlorine was passed through a solution of potassium iodide.
(c) State two uses of chlorine.
2. (a) Write an equation to show how hydrogen chloride can be prepared from sodium chloride.

(b) Draw a labelled diagram to show how aqueous hydrogen chloride can be prepared in the laboratory.

(c) State what would be observed and write equation for the reaction that would take place when aqueous hydrogen chloride is reacted with:
 - (i) Solid calcium carbonate.
 - (ii) Magnesium
 - (iii) Silver nitrate solution.
(d) State why aqueous hydrogen chloride does not react with copper.
3. (a) Draw a well labelled diagram to show how a sample of dry hydrogen chloride can be prepared.

(b) Dry hydrogen chloride gas was passed over heated iron filings. Write equation for the reaction that took place.

(c) The solid product in (b) was dissolved in water and aqueous sodium hydroxide added to the resultant solution drop wise until in excess.

- (i) State what was observed.
- (ii) Write an ionic equation for the reaction.

(e) Name one reagent that can be used to test for

- (i) the cation formed in(d)
- (ii) the anion formed in (d)

In each case state what is observed when the reagent you have named is used.

3. (a) Chlorine can be prepared in the laboratory from hydrochloric acid.

- (i) Name the other reagent used in the preparation of chlorine.
- (ii) State the conditions for the reaction.
- (iii) Write an equation for the reaction which takes place between hydrochloric acid and the reagent you have named in(i).

(b) (i) Draw a labelled diagram to show the preparation of iron(III) chloride using chlorine.

(ii) State what would be observed during the reaction.

(iv) Write an equation leading to the formation of iron(III) chloride.

(c) (i) State what would be observed if aqueous ammonia was added to a solution of iron(III) chloride.

(iii) Write an ionic equation for the reaction in c (i).

4. (a) Describe how a dry sample of hydrogen chloride can be prepared from a named chloride.[no diagram is required].Your description should include the following:

- (i) Condition for the reaction
- (ii) Name of the drying agent
- (iii) Method of collection
- (iv) Equation for the reaction

- (b) Name the substance that is formed when hydrogen chloride is passed through water.
- (c) (i) Name one reagent that can be used to test for the presence of chloride ions in solution.
- (ii) State what would be observed if the reagent was added to the chloride solution.
- (c) Write an ionic equation to show the reaction between aqueous hydrogen chloride and calcium hydrogen carbonate solution.
- (d) 25.0cm^3 of a 0.2M lead(II) nitrate solution was shaken with excess aqueous hydrogen chloride. Lead(II) ions react with chloride ions according to the following equation:
- $$\text{Pb}^{2+}(\text{aq}) + 2\text{Cl}^{-}(\text{aq}) \longrightarrow \text{PbCl}_2(\text{s})$$
- Calculate the mass of lead (II) chloride formed.
5. (a) Draw a labelled diagram of the set up of apparatus you would use to prepare chlorine in the laboratory using potassium manganate (VII).
- (b) State what is observed when
- (i) a piece of yellow phosphorus is lowered in a jar of chlorine.
- (ii) Burning turpentine ($\text{C}_{10}\text{H}_{16}$) is lowered in a jar of chlorine.
- (iv) Chlorine is bubbled in a solution of potassium bromide.
- (c) A test tube filled with a solution of chlorine water was inverted over water and exposed to sun light. After a few days, a gas formed in the test tube.
- (i) Name the gas.
- (ii) Describe a simple test to identify the gas.
- (iii) Explain the reaction which leads to the formation of the gas.
6. (a) State what would be observed if chlorine is passed through
- (i) Litmus solution
- (ii) Aqueous potassium iodide
- (iii) Cold dilute sodium hydroxide
- (iv) Iron(II) chloride solution
- (b) Write equation for the reaction in a(iv).
- (c) Chlorine is a bleaching agent when in the presence of water.

- (i) Write an equation for the reaction between chlorine and water.
 - (ii) Using equations, explain the bleaching action of chlorine.
- (e) (i) State what would be observed if chlorine was bubbled through a solution of iron(II) sulphate.
- (iii) Write an ionic equation for the reaction between chlorine and iron (II) ions.
7. (a) Draw a labelled diagram to show how hydrochloric acid can be prepared in the laboratory.
- (b) Compare the chemical properties of hydrogen chloride in water and methyl benzene.
- (c) Dry hydrogen chloride was passed over heated iron wire in a combustion tube.
- (i) State what was observed in the combustion tube.
 - (ii) Write the equation for the reaction.
- (d) The product in (c) was dissolved in water and aqueous sodium hydroxide added dropwise until in excess.
- (i) State what was observed.
 - (ii) Write equation for the reaction that took place.
8. (a) Draw a labelled diagram to show how iron(III) chloride can be prepared in the laboratory.
- (b) Write equation for the reaction that occurs.
- (c) State what is observed and write equation for the reaction that occurs when chlorine reacts with;
- (i) Sodium
 - (ii) Magnesium
 - (iii) Ammonia

IONIC CHEMISTRY AND QUALITATIVE ANALYSIS

Qualitative analysis refers to the identification of elements (by detecting the ions present) in a substance. It involves carrying out specific tests, making observations such as colour changes, colour of precipitate, and gases evolved.

Identification of unknowns.

Testing for cations (positive ions)

1. Appearance (colour of substances)

Colour	Deduction
White solid/ colourless solution	Zn^{2+} , Al^{3+} , Pb^{2+} , NH_4^+
Green	Copper(II)ions, Cu^{2+} , Iron(II)ions, Fe^{2+}
Yellow/ brown	Iron(III)ions, Fe^{3+}
Blue	Copper(II)ions, Cu^{2+}

2. Action of heat (*observation of residue when a substance is heated*)

Observation	Residue	Cation
Yellow when hot, white when cold	ZnO	Zn^{2+}
Reddish-brown when hot, yellow when cold	PbO	Pb^{2+}
Black when hot and cold	CuO, FeO	Cu^{2+} , Fe^{2+}

3. Reaction with sodium hydroxide solution

To the solution containing the cation, add sodium hydroxide drop wise until in excess. If no precipitate forms, warm gently and test the gas evolved .

Observation	Deduction
A white precipitate soluble in excess	$\text{Pb}^{2+}, \text{Zn}^{2+}, \text{Al}^{3+}$
Pale blue precipitate insoluble in excess	Cu^{2+}
Dirty- green precipitate insoluble in excess	Fe^{2+}
Brown precipitate insoluble in excess	Fe^{3+}
On warming the solution with sodium hydroxide, a colourless gas is evolved which turns moist red litmus paper blue.	Ammonia gas hence NH_4^+

$\text{Pb}^{2+}, \text{Zn}^{2+}$ and Al^{3+} react with sodium hydroxide solution and dissolves in excess due to the formation of soluble complexes.

Lead(II) ions form plumbate ion, $\text{Pb}(\text{OH})_4^{2-}$

Zinc ions form, zincate ion, $\text{Zn}(\text{OH})_4^{2-}$

Aluminium ions form, aluminate ions, $\text{Al}(\text{OH})_4^-$

1. Reaction with ammonia solution

To the solution containing the cation, add aqueous ammonia drop wise until in excess.

Observation	Deduction
White precipitate forms, insoluble in excess	$\text{Pb}^{2+}, \text{Al}^{3+}$
White precipitate forms, soluble in excess to form a colourless solution	Zn^{2+}
Pale blue precipitate, soluble in excess to form deep blue solution	Cu^{2+}
Dirty-green precipitate, insoluble in excess	Fe^{2+}
Brown precipitate insoluble in excess	Fe^{3+}

Zn^{2+} and Cu^{2+} ions react with aqueous ammonia to form hydroxide which dissolves in excess ammonia due to formation of soluble complex.

Zinc ions form, tetra-amine zinc(II) ion, $\text{Zn}(\text{NH}_3)_4^{2+}$ (colourless)

Copper(II) ions form, tetra-amine copper(II) ion, $\text{Cu}(\text{NH}_3)_4^{2+}$ (Deep blue)

5. Confirmatory tests for cations

Reagent	Observation	Cation
Addition of potassium iodide solution	Yellow precipitate	Pb^{2+}
Addition of aqueous ammonia drop wise until in excess	White precipitate soluble in excess to form colourless solution	Zn^{2+}
Addition of aqueous ammonia drop wise until in excess	Pale-blue precipitate soluble in excess to form deep blue solution	Cu^{2+}
Addition of aqueous sodium hydroxide + heat	Colourless gas evolved that turned moist red litmus paper blue	NH_3 gas hence NH_4^+
Addition of aqueous ammonia drop wise until in excess	Dirty-green precipitate insoluble in excess	Fe^{2+}
Addition of aqueous ammonia drop wise until in excess	Brown precipitate insoluble in excess	Fe^{3+}

6. Test for anions (Negative ions)

Heating the solid

Test	Observation	Deduction
Heat the unknown solid strongly until no further change in a hard test	(a) colourless gas that turns lime water milky	CO_2 gas hence CO_3^{2-} or HCO_3^-

tube.		
	(b) colourless liquid condenses on cooler parts of the test tube	Water of crystallization
	(c) Sublimation, sublimate condenses on cooler part of test tube.	NH_4^+

Addition of acid

Test	Observation	Deduction
Solid+dilute HCl	Effervescence occurs and a gas that turns lime water milky evolved	CO_3^{2-} or HCO_3^-
Solid+conc. H_2SO_4 and gentle warming	Brown fumes	NO_3^-

7. Confirmatory tests for anions

Reagent	Observation	Deduction
Add 1cm^3 of freshly prepared iron(II)sulphate solution followed by 1cm^3 conc.sulphuric acid down the side of the test tube	Brown ring at the junction of the layers.	NO_3^-
Add silver nitrate solution followed by dilute nitric acid	White precipitate is formed $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$	Cl^-
Add Barium nitrate solution followed by dilute nitric acid	White precipitate $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$	SO_4^{2-}

(i) Add dilute nitric acid(or HCl) to the solution or solid	Effervescence of colourless gas that turns lime water milky	$\text{CO}_3^{2-}/\text{HCO}_3^-$
(ii) Add magnesium sulphate solution.	<ul style="list-style-type: none"> White precipitate $\text{Mg}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{MgCO}_3$	CO_3^{2-}
	<ul style="list-style-type: none"> No white precipitate $\text{Mg}^{2+}(\text{aq}) + \text{HCO}_3^-(\text{aq}) \rightarrow \text{Mg}(\text{HCO}_3)_2$	HCO_3^-

Ionic equations

Ionic equation involves those ions that change states during a reaction. There are three possibilities during ionic changes.

- Formation of a precipitate.
- Evolution of a gas.
- Formation of water.

When writing ionic equations, the knowledge of substances that contain free ions and those that do not is vital.

<i>Contain free ions</i>	<i>Do not contain free ions</i>
Metal salts in solution	Insoluble solids including precipitates
Ammonium salts in solution	Covalent compounds
Gases in solution	Uncombined elements
Dilute acids	Gases
Soluble metal alkalis	Water

Note: solids, liquids and gases do not contain free ions and do not ionize.

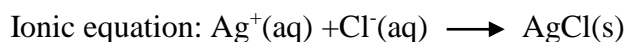
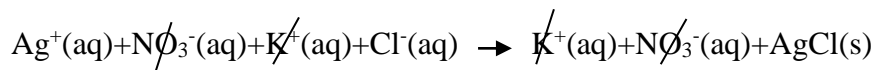
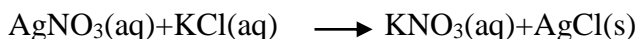
Steps when writing an ionic equation

- Write down the stoichiometric equation (normal chemical equation).

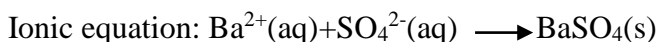
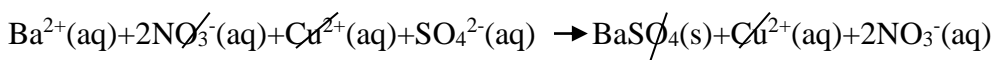
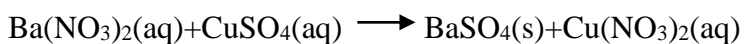
- (ii) Identify the substances that are ionic and ions that are free in solution. In aqueous solution, ions are completely separate from one another.
- (iii) Rewrite the equation in the ionic form.
- (iv) Cancel the “spectator ions” that appear on both sides of the equation as free ions.

1. ***The precipitation of an insoluble salt***

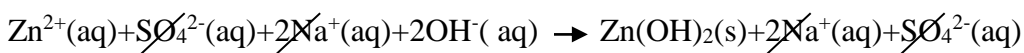
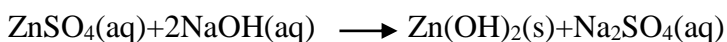
- (a) When silver nitrate solution is added to potassium chloride solution, a white precipitate is formed.



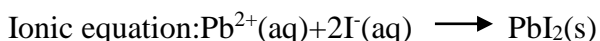
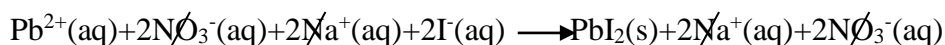
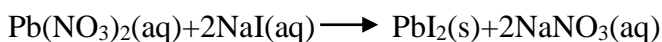
- (b) When barium nitrate solution is added to copper(II) sulphate solution, a white precipitate is formed.



- (c) When a solution of sodium hydroxide is added to a solution of zinc sulphate, a white precipitate is formed.

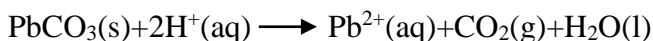
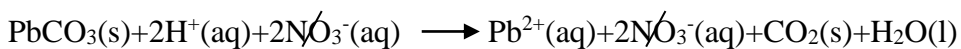
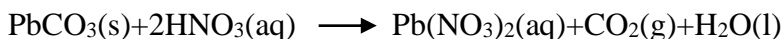


- (d) When sodium iodide solution is added to lead(II) nitrate solution, a yellow precipitate is formed.



1. ***Ionic equations involving solids***

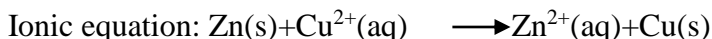
When dilute nitric acid is added to lead carbonate, aqueous lead(II)nitrate, carbon dioxide and water are formed.



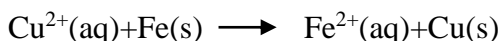
3. *Ionic equations for displacement reactions*

These are redox reactions which involve the displacement of one element from a solution of its compound by another element. A redox reaction is a reaction which involves both reduction and oxidation processes.

When zinc dust are added to copper(II)sulphate solutions, the solution turns from blue to colour less and a brown precipitate is formed.

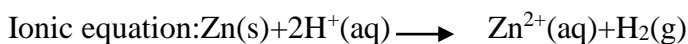


When iron filings are added to copper (II) sulphate solution, the solution turns from blue to pale green and a brown precipitate is formed.

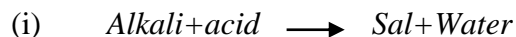


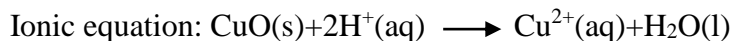
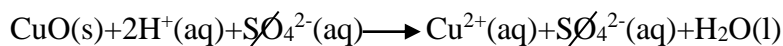
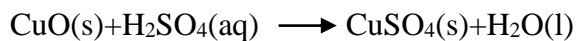
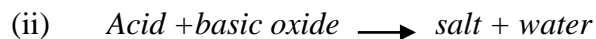
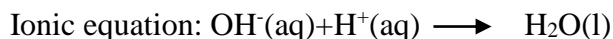
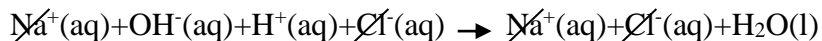
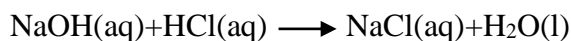
4. *Ionic equations for reaction between an acid with a metal*

When dilute sulphuric acid is added to zinc granules, hydrogen gas is evolved. Copper (II) sulphate solution may be added to speed up the reaction.

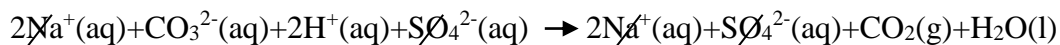
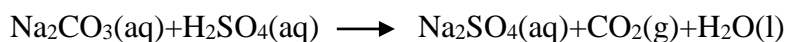


5. *Ionic equations for neutralization reactions*





6. ***Ionic equation for the reaction between metal carbonate and dilute acids***

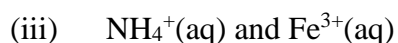
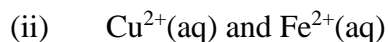
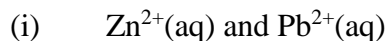


Exercise

1. Write ionic equations for the reactions between:

- A solution of barium chloride and a solution of magnesium sulphate.
- Potassium carbonate solution and dilute hydrochloric acid.
- Potassium iodide solution and chlorine gas.
- Magnesium and dilute nitric acid
- Calcium carbonate solid and dilute hydrochloric acid.

2. (a) Name one reagent that can be used to distinguish between each of the following pairs of cations. In each case state what would be observed if each cation is reacted with the reagent you have named.



(a) Name a reagent that reacts with $\text{CO}_3^{2-}(\text{aq})$ and $\text{SO}_4^{2-}(\text{aq})$ to show similar observations and another to distinguish the two anions. In each case state the observations made.

3. The table below shows some of the tests carried out on a green crystalline solid Q. The observations made were recorded as follows.

Test No.	Test	Observation
(i)	Q was heated strongly until no further change.	A colourless liquid condensed on cooler part of test tube. A colourless gas which turned aqueous potassium dichromate solution green was given out and a reddish-brown residue P was left.
(ii)	P was dissolved in dilute nitric acid and sodium hydroxide solution added to the aqueous solution.	A brown precipitate was formed.
(iii)	Chlorine was bubbled through an aqueous solution of Q.	The solution turned from green to yellow.

- (a) Identify substances P and Q.
(b) Name a substance that could be used to test for the colourless liquid.
(c) Write equations for the reactions that took place in tests (i), (ii) and (iii).
(d) Explain the reaction that took place in test (iii).
4. When aqueous ammonia was added drop wise until in excess to a solution containing a cation X, a white precipitate was formed which dissolved to give a colourless solution.
- (a) Identify X.
(b) Write the formula of the cation in the colourless solution.
(c) Write ionic equation for the reaction leading to the formation of the white precipitate.
(d) Name one other metal ion that when treated with aqueous ammonia would form a precipitate soluble in excess.
(e) State what would be observed if the metal ion you have named in (d) was treated with aqueous ammonia until in excess.

MOLE CONCEPT

DIFFUSION

Diffusion is the movement of molecules from a region of high concentration to a region of low concentration. This is due to the bombardment of the particles with the moving air molecules.

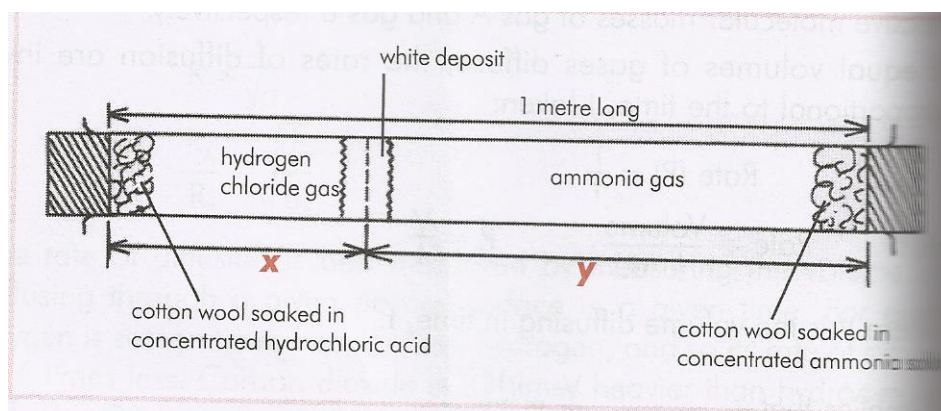
Diffusion in Gases

Place a drop of reddish-brown bromine liquid into a gas jar containing air and cover it.

The liquid bromine vaporizes and the gas jar gets filled with bromine vapour.

The bromine particles and particles of air have mixed thoroughly.

DIFFUSION OF AMMONIA AND HYDROGEN CHLORIDE

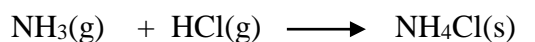


A piece of cotton wool soaked in concentrated hydrochloric acid is placed at one end A of a long glass tube and gives off hydrogen chloride fumes. The piece of cotton wool soaked in concentrated Ammonia is placed at the other end B and gives off ammonia fumes.

After 5 minutes, a white ring forms.

The white ring does not form immediately because the gas particles have to encounter air particles in the tube before they meet.

A white ring forms nearer the hydrochloric acid end A. this is because Ammonia particles and diffuses faster. The white ring is due to the formation of Ammonium chloride.



White ring

Diffusion in liquids.

- Pour some water into a beaker until it's half full.
- Drop in a crystal of potassium manganate(VII).

- Leave to stand for 10 minutes.

Observation.

The purple colour of potassium manganate (VII) spreads throughout the whole solution.

This shows that matter consist of particles.

Brownian motion.

The movement of small particles making up matter is best studied by the smoke cell experiment.

The smoke is sucked into the cell and illuminated. The movement of the smoke particles is then observed through a microscope.

Observation.

The smoke particles are seen in continuous rapid random movement. This type of movement is called Brownian motion.

THE MOLE

It's a unit of measurement based on a fixed number of particles. A mole of any substance is the amount of it which contain as many particles (atoms, molecules, ions, electrons) as there are carbon atoms in 12g of carbon 12.

The number of particles in one mole of any substance is 6.02×10^{23} . It's called Avogadro's constant denoted by L.

TERMS USED

1 Atomic mass: The mass of one mole of atoms of an element.

2 Relative atomic mass (R.A.M)

The mass of an atom of an element compared with $\frac{1}{12}$ th of the mass of carbon 12 atom.

3 Molar mass

This is the mass of one mole of a substance or compound. It's the sum of molar masses of the constituent atoms.

Units: grams.

4 Relative molecular mass (RMM)

This is the mass of one mole of a substance compared to $\frac{1}{12}$ th of the mass of carbon – 12 atom.

Units: No units

- 5 Formula mass. This is the mass of one mole of an ionic substance.

This is the mass of one mole of an ionic substance.

Unit: grams

- 6 Relative formula mass (RFM)

This is the mass of one mole of ionic compound compared with the mass of $\frac{1}{12}$ th of carbon – 12 isotope.

Units: No units.

CALCULATION OF MOLAR MASSES

When calculating molar mass, note the following:

- (i) You must know how to write the correct formula of substances.
- (ii) You must know how to use brackets to avoid wrong calculations.
- (iii) Hydrated compounds contain water of crystallization; correct knowledge of their formulae is vital.

Examples

- 1 Calculate the molar mass of calcium nitrate (Ca = 40, N = 14, O = 16)

$$\begin{aligned}\text{Ca}(\text{NO}_3)_2 &= 40 + (14 \times 2) + (16 \times 3 \times 2) \\ &= 164\text{g}.\end{aligned}$$

- 2 Calculate the molar mass of hydrated copper (II) sulphate crystals, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

(Cu = 64, S = 32, O = 16, H = 1)

$$\begin{aligned}\text{CuSO}_4 \cdot 5\text{H}_2\text{O} &= 64 + 32 + (16 \times 4) + 5(1 \times 2 + 16) \\ &= 250\text{g}.\end{aligned}$$

- 3 Calculate the molar mass of Ammonium phosphate (N = 14, H = 1, P = 31, O = 16, H = 1)

$$\begin{aligned}(\text{NH}_4)_3\text{PO}_4 &= (14 \times 3) + (1 \times 4 \times 3) + 31 + (16 \times 4) \\ &= 149\text{g}\end{aligned}$$

Exercise

- 4 Calculate the molar masses of the following.

- (i) Sodium sulphate (Na = 23, S = 32, O = 16)
- (ii) Potassium carbonate (K = 39, C = 12, O = 16)
- (iii) Lead(II) nitrate (Pb = 207, N = 14, O = 16)
- (iv) Ammonium carbonate (H = 1, N = 14, C = 12, O = 16)
- (v) Zinc sulphate (Zn = 65, S = 32, O = 16)

(vi) Magnesium nitrate (Mg=24,N=14,O=16,)

CALCULATION OF PERCENTAGE COMPOSITION.

Example

- 1 Calculate the percentage of Iron by mass in Iron(II) sulphate crystals, $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$.

(Fe = 56, S = 32, O = 16, H = 1)

$$\begin{aligned}\text{FeSO}_4 \cdot 7\text{H}_2\text{O} &= 56 + 32 + (16 \times 4) + 7(1 \times 2 + 16) \\ &= 278\text{g}.\end{aligned}$$

$$\% \text{ of Fe} = \frac{56}{278} \times 100\%$$

$$= 20.1\%$$

- 2 Calculate the percentage of oxygen in sodium carbonate crystals, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$

(Na = 23, C = 12, O = 16, H = 1)

$$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O} = (23 \times 2) + 12 + (16 \times 3) + 10(1 \times 2 + 16)$$

$$= 286\text{g}$$

$$\% \text{ of oxygen} = \frac{208}{286} \times 100\%$$

$$= 72.7\%$$

- 3 Calculate the percentage of Nitrogen in Ammonium sulphate (N= 14, H = 1, S = 32, O = 16)

$$(\text{NH}_4)_2\text{SO}_4 = (14 \times 2) + (1 \times 4 \times 2) + 32 + (16 \times 4)$$

$$28 + 8 + 32 + 64$$

$$= 132\text{g}$$

$$\% \text{ of Nitrogen} = \frac{28}{132} \times 100\%$$

- 4 Calculate the percentage of Aluminium in Aluminium sulphate (Al = 27, S = 32, O = 16)

$$\text{Al}_2(\text{SO}_4)_3 = (27 \times 2) + (32 \times 3) + (16 \times 4 \times 3)$$

$$= 342\text{g}.$$

$$\% \text{ of Al} = \frac{54}{342} \times 100\%$$

$$= 15.78\%$$

Exercise

Calculate the percentage of the following substances in their compounds:

- (i) Sulphur in hydrogen sulphide(S=32,H=1)
- (ii) Oxygen in glucose(C₆H₁₂O₆)
- (iii) Nitrogen in urea [(NH₂)₂CO] (N=14, H=1,C=12,O=16)
- (iv) Nitrogen in ammonium sulphate(N=14, H=1, S=32,O=16)

CALCULATING NUMBERS OF MOLES GIVEN MASES.

Masses and moles of elements or compounds are related by the formula

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

Examples.

- 1 Calculate the number of moles of 3.31g of lead nitrate (N = 14, O = 16, Pb = 207)

$$\text{Pb(NO}_3)_2 = 207 + (14 \times 2) + (16 \times 3 \times 2)$$

$$331\text{g}$$

331g of Pb(NO₃)₂ contain 1 mole

$$1\text{g of Pb(NO}_3)_2 \text{ contains } \left(\frac{1}{331}\right)\text{mole}$$

$$= 3.31\text{g of Pb(NO}_3)_2 \text{ contain } \frac{1}{331} \times 331$$

$$= 0.01\text{moles.}$$

- 2 Calculate the number of moles of 8.8g of carbondioxide (C = 12, O = 16)

$$\text{CO}_2 = 12 + (16 \times 2)$$

$$= 44\text{g.}$$

44g of CO₂ contain 1 mole

$$8.8\text{g of CO}_2 \text{ contain } \frac{1}{44} \times 8.8$$

$$= 0.2 \text{ moles of CO}_2$$

3. Calculate the number of moles 3.2g of copper (Cu = 64)

64g of Cu contain 1mole.

$$3.2\text{g of Cu contain } \frac{1}{64} \times 3.2$$

$$=0.05\text{m0les}$$

4. Calculate the number of moles of 4g of sodium hydroxide (Na=40,O=16,H=1)

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

$$= 40\text{g}$$

40g of NaOH contain 1mole.

$$\begin{aligned} 4\text{g of NaOH contain } \frac{1}{40} \times 4 \\ = 0.1\text{moles.} \end{aligned}$$

Exercise

Calculate the number of moles of the following:

- (i) 13.3g of ammonium sulphate (N=14,H=1,S=32,O=16)
- (ii) 4.8g of magnesium (Mg=24).
- (iii) 8.1g of aluminium sulphate (Al=27,S=32,O=16)
- (iv) 1.68g of potassium hydroxide (K=39,H=1,O=16)
- (v) 10.8g of silver (Ag=108)

CALCULATING MASSES GIVEN THE NUMBER OF MOLES

Examples

1. Calculate the mass of 0.025moles of sodium hydroxide (Na=23,O=16,H=1)

Molar mass of NaOH=23+16+1=40g

1mole of NaOH contain 40g

0.025moles contain 40×0.025

$$= 1.0\text{g}$$

2. Calculate the mass of 0.1 moles of sodium carbonate (Na=23,C=12,O=16)

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

1mole of Na_2CO_3 contain 106g

0.1mole of Na_2CO_3 contain 106×0.1

$$= 10.6\text{g}$$

3. Calculate the mass of 0.2 moles of lead (II) nitrate (Pb=207,N=14,O=16)

Molar mass of $\text{Pb}(\text{NO}_3)_2 = 207 + 14 \times 2 + 16 \times 3 \times 2 = 331\text{g}$

1mole of $\text{Pb}(\text{NO}_3)_2$ contain 331g

0.2moles of $\text{Pb}(\text{NO}_3)_2$ contain 331×0.2

$$= 66.2\text{g}$$

4. Calculate the mass of 0.02moles of zinc nitrate (Zn=65,N=14,O=16)

Molar mass of $\text{Zn}(\text{NO}_3)_2 = 65 + (14 \times 2) + (16 \times 3 \times 2)$

$$=189\text{g}$$

1 mole of $\text{Zn}(\text{NO}_3)_2$ contain 189g

0.02moles of $\text{Zn}(\text{NO}_3)_2$ contain 189×0.02

$$=3.78\text{g}$$

Exercise

Calculate the masses of the following:

1. 0.25moles of ammonium phosphate ($\text{N}=14, \text{P}=31, \text{O}=16, \text{H}=1$)
2. 0.02 moles of zinc chloride ($\text{Zn}=65, \text{Cl}=35.5$)
3. 0.4 Moles of sulphur($\text{S}=32$)
4. 0.001moles of Sodium sulphate ($\text{Na}=23, \text{O}=16, \text{S}=16, \text{O}=16$)

DETERMINATION OF FORMULA

Determination of formula of copper (II) oxide by reduction.

Diag

- ❖ Weigh a clean dry porcelain boat.
- ❖ Put some copper (II) oxide into a porcelain boat and weigh.
- ❖ Pass dry hydrogen gas through the combustion tube.
- ❖ Warm the copper (II) oxide for about 10 minutes.
- ❖ Allow the porcelain boat and its contents to cool.
- ❖ Remove and weigh the porcelain boat and its contents.

Note: The combustion tube must be sloping downwards so that the water formed collects in the cooler part of the tube and does not run back breaking the tube.

Sample results

1. Mass of porcelain boat =12.50g
Mass of porcelain boat + Copper (II) oxide=13.75g
Mass of porcelain boat + copper metal =13.50g
(b) Calculate the mass of oxygen and copper from the above results.
Mass of copper (II) oxide =13.75- 12.50
$$= 1.25\text{g}$$

Mass of copper metal=13.5-12.50
$$= 1.0\text{g}$$

$$\begin{aligned}\text{Mass of oxygen} &= 1.25 - 1.0 \\ &= 0.25\text{g}\end{aligned}$$

(c) Determine the formula of the oxide of copper (Cu=64, O=16)

	Cu	O
	1.0	0.25
Moles	$\frac{1.0}{64}$	$\frac{0.25}{16}$
	$\frac{0.015}{0.015}$	$\frac{0.015}{0.015}$
	1	1

Formula of the oxide of copper ***CuO***

1. Mass of crucible and lid=25.8g

Mass of crucible, lid and magnesium ribbon=28.2g

Mass of crucible, lid and magnesium oxide=29.7g

Calculate the formula of magnesium oxide (Mg=24, O=16)

$$\begin{aligned}\text{Mass of magnesium ribbon} &= 28.2 - 25.8 \\ &= 2.4\text{g}\end{aligned}$$

$$\begin{aligned}\text{Mass of magnesium oxide} &= 29.7 - 25.8 \\ &= 3.9\text{g}\end{aligned}$$

$$\begin{aligned}\text{Mass of oxygen} &= 3.9 - 2.4 \\ &= 1.5\text{g}\end{aligned}$$

	Mg	O
	2.4	1.5
Moles	$\frac{2.4}{24}$	$\frac{1.5}{16}$
	0.1	0.1
Ratio	$\frac{0.1}{0.1}$	$\frac{0.1}{0.1}$
	1	1

Formula of magnesium oxide ***MgO***

EMPIRICAL AND MOLECULAR FORMULAE

Empirical Formula

Empirical formula of a compound is the formula which shows the number of atoms and the simplest ratio in which they exist. It is also called the *simplest formula*.

Calculating empirical formula given masses in grams

Examples

1. A compound X consists of 4g of hydrogen and 64g of oxygen.

Calculate the empirical formula of compound X. (H=1, O=16)

Element	H	O
Mass	4	64
Moles	$\frac{4}{1}$	$\frac{64}{16}$
	4	4
Divide by		
Smallest	$\frac{4}{4}$	$\frac{4}{4}$
Ratio	1	1
Empirical formula	HO	

2. 1.5g of magnesium completely burns in oxygen to form 2.5g of magnesium oxide.

Calculate the empirical formula of magnesium oxide. (Mg=24, O=16)

Mass of oxygen = 2.5 - 1.5 = 1.0g

Mg	O
1.5	1.0
Moles $\frac{1.5}{24}$	$\frac{1.0}{16}$
0.0625	0.0625
$\frac{0.0625}{0.0625}$	$\frac{0.0625}{0.0625}$
Ratio 1	1

Empirical formula **MgO**

3. 20g of a compound of iron contains 14g of iron and the rest oxygen. Calculate the empirical formula of the compound of iron (Fe=56, O=16)

Mass of oxygen = 20 - 14 = 6.0g

Elements :	Fe	O
Mass	14	6.0
Moles	$\frac{14}{56}$	$\frac{6.0}{16}$
	0.25	0.375
	$\frac{0.25}{0.25}$	$\frac{0.375}{0.25}$
Ratio	1x2	1.5x2
	2	3

Empirical formula **Fe₂O₃**

4. A mass of 6.5g of metal X combined with oxygen to form an oxide of mass 8.1g. Determine the simplest formula of the oxide (X=65, O=16).

Mass of oxygen = 8.1 - 6.5 = 1.6g

	X	O
	6.5	1.6
Moles	$\frac{6.5}{65}$	$\frac{1.6}{16}$
	0.1	0.1
Ratio	$\frac{0.1}{0.1}$	$\frac{0.1}{0.1}$
	1	1

Simplest formula **XO**

Calculating Empirical formula given mass in percentages

Examples

1. A compound **K**, contains 43.4% by mass of sodium, 11.3% of carbon and 45.3% of oxygen. Calculate the empirical formula of the compound (Na=23, C=12, O=16)

Element:	Na	C	O
% mass:	43.4	11.3	45.3
Moles:	$\frac{43.4}{23}$	$\frac{11.3}{12}$	$\frac{45.3}{16}$
	1.886	0.94	2.83
Divide by	$\frac{1.89}{0.94}$	$\frac{0.94}{0.94}$	$\frac{2.83}{0.94}$
smallest			
Ratio	2	1	3

Empirical formula **Na₂CO₃**

2. A compound **R**, contains 28% by mass of iron, 24% of sulphur and 48% of oxygen.

Calculate the empirical formula of compound R (Fe=56, S=32, O=16)

Elements:	Fe	S	O
% mass	28	24	48
Moles	$\frac{28}{56}$	$\frac{24}{32}$	$\frac{48}{16}$
	0.5	0.75	3

Divide by

Smallest	$\frac{0.5}{0.75}$	$\frac{0.75}{0.75}$	$\frac{3}{0.75}$
	0.66	1	4
	1	1	4

Empirical formula **FeSO₄**

2. A hydrocarbon **Q**, consists of 80% carbon. Calculate the empirical formula of the hydrocarbon Q (H=1, C=12).

Mass of hydrogen = 100 - 80 = 20g

Elements:	C	H
% mass	80	20
Moles	$\frac{80}{12}$	$\frac{20}{1}$
	6.66	20

Divide by

Smallest	$\frac{6.66}{6.66}$	$\frac{20}{6.66}$
	1	3

Empirical formula **CH₃**

3. A compound **T**, consists of 39% by mass potassium, 1% hydrogen, 12% carbon and the rest oxygen. Calculate the empirical formula of T (K=39, H=1, C=12, O=16).

Mass of oxygen = 100 - 52 = 48%

Elements:	K	H	C	O
% mass	39	1	12	48
Moles	$\frac{39}{39}$	$\frac{1}{1}$	$\frac{12}{12}$	$\frac{48}{16}$

	1	1	1	3
Divide by				
Smallest	$\frac{1}{1}$	$\frac{1}{1}$	$\frac{1}{1}$	$\frac{3}{1}$
	1	1	1	3

Empirical formula **KHCO₃**

5. A compound **Y**, consists of 16.08% sodium, 4.2% carbon, 16.78% oxygen and 62.94% water.
(Na=23,H=1 ,O=16, C=12).

Elements:	Na	C	O	H ₂ O
% mass	16.08	4.2	16.78	62.94
Moles	$\frac{16.08}{23}$	$\frac{4.2}{12}$	$\frac{16.78}{16}$	$\frac{62.78}{18}$
	0.699	0.35	1.048	3.487
Ratio	$\frac{0.699}{0.35}$	$\frac{0.35}{0.35}$	$\frac{1.048}{0.35}$	$\frac{3.487}{0.35}$
	1.99	1	2.99	9.96
	2	1	3	10
Simplest formula	Na₂ CO₃.10 H₂O			

MOLECULAR FORMULAE

Molecular formula is the formula which shows the actual number of each kind of atom present in one mole of a substance. It can be obtained from empirical formula if the relative molecular mass is known.

$$\frac{\text{molecular Mass}}{\text{Empirical formula mass}} = n, \text{ a whole number.}$$

Examples

1. A compound **X**, consists of 4g of hydrogen and 64g of oxygen. The molecular mass of X=34(
H=1,O=16)

(a) Calculate the empirical formula of X

Elements:	H	O
Mass	4	64

Moles	$\frac{4}{1}$	$\frac{64}{16}$
	4	4
Ratio	1	1
Simplest formula	HO	

(d) Calculate the molecular formula of X

$$(\text{HO})_n = 34$$

$$1 \times n + 16 \times n = 34$$

$$n + 16n = 34$$

$$17n = 34$$

$$n = \frac{34}{17}$$

$$= 2$$

Molecular formula $(\text{HO})_2$



3. A compound **S**, contains 40% carbon, 6.67% hydrogen and the rest oxygen. The molar mass of the compound is 180.

(i). Calculate the empirical formula of the compound.

$$\text{Mass of oxygen} = 100 - 46.67 = 53.3\%$$

	C	H	O
	40	6.67	53.3
Moles	$\frac{40}{12}$	$\frac{6.67}{1}$	$\frac{53.3}{16}$
	$\frac{3.33}{3.33}$	$\frac{6.67}{3.33}$	$\frac{3.33}{3.33}$
Ratio	1	2	1

Empirical formula = CH_2O

(ii). Determine the molecular formula of the compound.

$$(\text{CH}_2\text{O})_n = 180$$

$$12n + 2n + 16n = 180$$

$$30n = 180$$

$$n = \frac{180}{30}$$

$$= 6.$$

Molecular formula = **C₆ H₁₂ O₆**

4. A carbohydrate **J**, contains 42.11% carbon, 51.46% oxygen and the rest hydrogen. The vapour density of the carbohydrate is 171. Calculate the molecular formula of the carbohydrate (H = 1, C = 12, O = 16).

Mass of hydrogen = 100 - 42.11 - 51.46 = 6.43g

C	H	O
42.11	6.43	51.46

Moles	$\frac{42.11}{12}$	$\frac{6.43}{1}$	$\frac{51.46}{16}$
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$\frac{3.51}{3.22}$	$\frac{6.43}{3.22}$	$\frac{3.22}{3.22}$
---------------------	---------------------	---------------------

1.1	1.99	1
-----	------	---

Ratio	1	2	1
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Empirical formula **CH₂O**

Molar mass = Vapour density \times 2 = 171 \times 2

$$(\text{CH}_2\text{O})_n = 342$$

$$12n + 2n + 16n = 342$$

$$30n = 342$$

$$n = \frac{342}{30}$$

$$= 11$$

$$\text{Molecular formula} = (\text{CH}_2\text{O})_{11}$$



Exercise

- 1 (a) When 1.34 g of copper was heated in air, 1.68g of an oxide was formed.
Calculate the empirical formula of the oxide.(Cu=63.5, O=16).
(b) To the oxide formed in (b) was added dilute hydrochloric acid.
(i) State what was observed.
(ii) Write equation for the reaction.
2. (a) Oxygen can be prepared in the laboratory from hydrogen peroxide and substance W. Identify substance W and state its role.
(b) Potassium chlorate decomposes on heating to give oxygen according to the following equation:

$$2 \text{KClO}_3(\text{s}) \longrightarrow 2 \text{KCl}(\text{s}) + 3 \text{O}_2(\text{g})$$

Calculate the volume of oxygen produced at room temperature when 10.6g of potassium chlorate was heated.[K=39, Cl=35.5, O=16, 1mole of gas at room temperature occupies 24cm³].
3. A compound Q of formula mass 60, contains carbon, 40%, hydrogen 6.7% and the rest being oxygen.
(a) (i) Calculate the empirical formula of Q.
 (H=1; C=12; O=16)
 (ii) Determine the molecular formula of Q.
 (b) Q dissolves in aqueous sodium hydrogen carbonate with effervescence.
 (i) Suggest the chemical nature of Q.
 (ii) Write an ionic equation for the reaction between Q and aqueous sodium hydrogen carbonate.
- 4 A compound P contains 28.1% iron, 35.5% chlorine, and 36.2% of water of crystallization by mass. Calculate the simplest formula of P (Fe=56; Cl=35.5; H=1; O=16).

MOLAR GAS VOLUME

Molar gas volume is the volume occupied by one mole of gas at s.t.p or room temperature. It's the same for all gases under the same conditions of temperature and pressure.

S.t.p refers to standard temperature and pressure.

Molar gas volume at s.t.p is 22.4dm^3 or 22400cm^3 .

Molar gas volume at room temperature is 24dm^3 or 2400cm^3 .

CALCULATING MOLES OF GAS GIVEN VOLUME AND MOLAR VOLUME.

1. Calculate the number of moles of Ammonia present in 3.6dm^3 at room temperature.

24dm^3 contain 1 mole of ammonia.

1dm^3 contain $\left(\frac{1}{24}\right)$ moles of NH_3

3.6dm^3 contain $\frac{1}{24} \times 3.6$ moles

= 0.15 moles of NH_3

2. Calculate the number of moles in 1.12dm^3 of hydrogen gas at s.t.p 22.4dm^3 contain 1 mole of H_2 .

1dm^3 contain $\left(\frac{1}{22.4}\right)$ moles of H_2

1.12dm^3 contain $\left(\frac{1}{22.4} \times 1.12\right)$ moles

= 0.05 moles.

3. Calculate the number of moles of ammonia gas in 560cm^3

22400cm^3 contain 1 mole of NH_3

1cm³ contain $\left(\frac{1}{22400}\right)$ moles of NH₃

560cm³ contain $\frac{1}{22400} \times 560$

= 0.025moles.

4. Calculate the number of moles in 600cm³ of hydrogen chloride gas at room temperature.

22400cm³ contain 1 mole of HCl

1cm³ contain $\left(\frac{1}{22400}\right)$ moles of HCl

600cm³ contain $\frac{1}{22400} \times 600$

=0.0267 moles.

CALCULATING MOLAR MASS GIVEN MASS AND VOLUME

Examples

1. The mass of a gas X which occupies 560cm³ at s.t.p is 1.10g

1cm³ of X contain $\left(\frac{1.10}{56}\right)$ g

22400cm³ of X contain $\frac{1.10}{56} \times 22400$

= 44g

- 2 calculate the volume that 8g of oxygen gas would occupy at s.t.p (O = 16, 1 mole of a gas occupies 22.4l at s.t.p).

32g of O₂ occupies 22.4l

1g of O₂ occupies $\left(\frac{22.4}{32}\right)$ l

8g of O₂ occupies $\frac{22.4}{32} \times 8$

$$= 5.6l$$

Note: the molar mass of the gas
 $= 2 \times \text{vapour density.}$

CALCULATING VOLUME OF GASES FROM EQUATIONS.

Examples

1. 42.0g of sodium hydrogen carbonate are heated strongly until no further change according to the following equation.



Calculate the volume of the gas evolved at s.t.p (H = 1, C = 12, O = 16, Na = 23, 1 mole of gas occupies 22.4dm³ at s.t.p).

$$\begin{aligned}\text{Molar mass of } 2\text{NaHCO}_3 \\ &= 2(23 + 1 + 12 + 16 \times 3) \\ &= 168\text{g}\end{aligned}$$

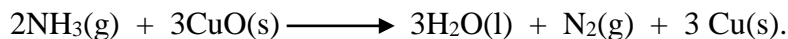
168g of NaHCO₃ liberates 22.4dm³ of CO₂

$$1\text{g of NaHCO}_3 \text{ liberate } \left(\frac{22.4}{168}\right) \text{ dm}^3$$

$$42.0\text{g of NaHCO}_3 \text{ liberate } \left(\frac{22.4}{168} \times 42.0\right) \text{ dm}^3$$

$$\text{Volume of CO}_2 = 5.6\text{dm}^3$$

- 2 Ammonia reacts with copper (II) oxide to form copper according to the following equation.



Calculate the volume of ammonia required to react with 6.0g of copper (II) oxide at s.t.p
(H = 1, Cu = 64, O = 16, N = 14, 1 mole of gas occupies 22400cm³ at s.t.p).

$$\begin{aligned}\text{Molar mass of CuO} &= (64 + 16)3 \\ &= 80\text{g} \times 3 \\ &= 240\text{g}\end{aligned}$$

$$\begin{aligned}\text{Volume of NH}_3 &= 2 \times 22400 \\ &44800\text{cm}^3\end{aligned}$$

240g of CuO require 44800cm³ of NH₃

$$1\text{g of CuO requires } \left(\frac{44800}{240}\right)\text{cm}^3 \text{ of NH}_3$$

$$\begin{aligned}6.0\text{g of CuO require } &\frac{44800}{240} \times 6.0 \\ &= 1120\text{cm}^3\end{aligned}$$

- 3 Calcium carbonate reacted with dilute hydrochloric acid according to the equation



Calculate the volume of carbondioxide evolved at room temperature when 15.9g of calcium carbonate was reacted with excess acid (Ca = 40, C = 12, O = 16, 1 mole of gas at room temperature occupies 24dm³)

$$\begin{aligned}\text{Molar mass of CaCO}_3 &= 40 + 12 + 16 \times 3 \\ &= 100\text{g}.\end{aligned}$$

100g of CaCO₃ produce 24dm³ of CO₂

$$1\text{g of CaCO}_3 \text{ produce } \left(\frac{24}{100}\right)\text{dm}^3 \text{ of CO}_2$$

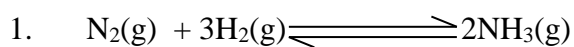
15.9g CaCO_3 produce $\frac{24}{100} \times 15.9$

Volume of $\text{CO}_2 = 3.74\text{dm}^3$.

CALCULATIONS OF VOLUMES FROM GAY LUSSAC'S LAW

When gases react, they do so in volumes that bear a simple ratio to each other and to the gaseous products provided all volumes are measured under the same conditions of temperature and pressure.

Example

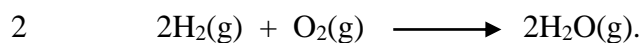


Calculate the volume of hydrogen required to react completely with 12cm^3 of Nitrogen gas.

	N_2	H_2
Ratio	1	3
Vol.	12	x

$$\text{Volume of } \text{H}_2 = 12 \times 3$$

$$= 36\text{cm}^3$$



A volume of 50cm^3 of hydrogen were mixed with oxygen and the mixture was sparked.

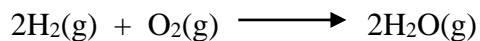
Calculate the volume of oxygen that reacted.

H_2	O_2
2	1
50	x

$$\text{Volume of oxygen} = \frac{50}{2}$$

$$= 25\text{cm}^3.$$

3. 50cm^3 of oxygen are exploded with 50cm^3 of hydrogen. What is the volume of the residual gas at room temperature?

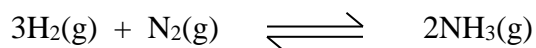


$$\text{Volume of O}_2 \text{ reacted} = 25\text{cm}^3$$

$$\text{The residual volume} = 50 - 25$$

$$= 25\text{cm}^3$$

- 4 Under a certain temperature and pressure, hydrogen reacted with Nitrogen according to the equation.



Calculate the volume of Nitrogen required to react with 150cm^3 of hydrogen under the same temp and pressure



$$\text{Volume of N}_2 = \frac{150}{3}$$

$$= 50\text{cm}^3$$

CALCULATION OF MASS FROM EQUATIONS.

1. Zinc carbonate decomposes according to the following equation when heated



Calculate the mass of Zinc Oxide when 2.5g of zinc carbonate is heated (Zn = 65, O = 16, C = 12)

$$\begin{aligned}\text{Molar mass of ZnCO}_3 &= 65 + 12 + 16 \times 3 \\ &= 125\text{g}\end{aligned}$$

$$\begin{aligned}\text{Molar mass of ZnO} &= 65 + 16 \\ &= 81\text{g}\end{aligned}$$

125g of ZnCO_3 produce 81g of ZnO .

1g of ZnCO_3 produce $\left(\frac{81}{125}\right)$ g of ZnO

$$\begin{aligned}2.5\text{g of ZnCO}_3 \text{ produce } &\frac{81}{125} \times 2.5 \\ &= 1.62\text{g}.\end{aligned}$$

2. Calcium carbonate reacts with hydrochloric acid according to the equation.



Calculate the mass in grams of carbondioxide liberated when 20g of calcium carbonate is completely reacted with hydrochloric acid (Ca=40, C=12, O=16)

$$\begin{aligned}\text{Molar mass of CaCO}_3 &= 40 + 12 + 16 \times 3 \\ &= 100\text{g}\end{aligned}$$

$$\begin{aligned}\text{Molar mass of CO}_2 &= 12 + 16 \times 2 \\ &= 44\text{g}\end{aligned}$$

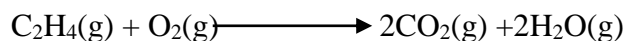
100g of CaCO_3 liberates 44g of CO_2

1g of CaCO_3 liberate $\left(\frac{44}{100}\right)$ g of CO_2

$$20\text{g of CaCO}_3 \text{ liberate } \left(\frac{44}{100} \times 20\right) \text{ g of CO}_2$$

$$= 8.8\text{g}$$

3. Ethene burns in oxygen according to the equation:



Calculate the mass of each of the products when 14g of ethene is completely burnt in oxygen (C=12, H=1).

$$\text{Molar mass of C}_2\text{H}_4 = 12 \times 2 + 1 \times 4$$

$$= 28\text{g}$$

$$\text{Formula mass of } 2\text{CO}_2 = 2(12 + 16 \times 2)$$

$$= 88\text{g}$$

$$\text{Formula mass of } 2\text{H}_2\text{O} = 2(1 \times 2 + 16)$$

$$= 36\text{g}$$

28g of C₂H₄ produce 36g of H₂O

$$14\text{g of C}_2\text{H}_4 \text{ produce } \left(\frac{36}{28} \times 14\right) \text{ g of H}_2\text{O}$$

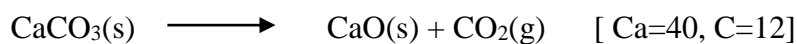
$$= 18\text{g of water}$$

28g of C₂H₄ produce 88g of CO₂

$$14\text{g of C}_2\text{H}_4 \text{ produce } \left(\frac{88}{28} \times 14\right) \text{ g of CO}_2$$

$$= 44\text{g of carbon dioxide}$$

4 Calcium carbonate decomposes according to the following equation:



Calculate the mass of calcium carbonate required to produce 28g of calcium oxide.

$$[\text{Ca}=40, \text{C}=12]$$

$$\text{Molar mass of CaCO}_3 = 40 + 12 + 16 \times 3$$

$$= 100\text{g}$$

$$\text{Molar mass of CaO} = 40 + 16$$

$$= 56\text{g}$$

$$56\text{g of CaO require } 100\text{g of CaCO}_3$$

$$1\text{g of CaO requires } \left(\frac{100}{56}\right)\text{g of CaCO}_3$$

$$28\text{g of Ca O require } \left(\frac{100}{56} \times 28\right)\text{g of CaCO}_3$$

$$= 50\text{g}$$

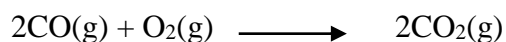
Exercise

1. Copper (II) nitrate decomposes on heating according to the following equation:



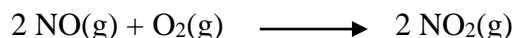
- (i) Calculate the mass of copper(II) oxide formed when 6.4g of copper(II) nitrate was heated.
- (ii) Calculate the total volume of the gaseous products formed, at s.t.p on heating the same mass of copper (II) nitrate. (Cu=64; N=14; O=16; 1 mole of gas at s.t.p occupies 22.4 dm³).

2. Carbon monoxide reacts with oxygen according to the equation below:



Calculate the volume of oxygen remaining after 10cm³ of carbon monoxide is mixed with 45cm³ of oxygen.

3. Nitrogen monoxide reacts with oxygen according to the following equation:



Calculate the volume of nitrogen dioxide formed when 15cm³ of nitrogen monoxide reacts with excess oxygen.

4. Propane burns in air according to the following equation:



Calculate the mass of oxygen required for complete combustion of 4.4g of propane(C=12; H=1; O=16).

MOLAR SOLUTIONS

A molar solution is a solution containing one mole of solute in one litre of solution. A molar solution of any compound contains one mole of that compound in one litre(1000cm³) of a solution.

Common terms used:

2. Concentration; concentration refers to the amount of solute dissolved in a specific volume of solvent.
3. Concentration in grams per litre is the mass of a substance dissolved in 1 litre of solution.
4. Standard solution is a solution whose concentration is known.
5. Molarity is the number of moles of a solute contained in one litre of solution.

The concentration of the solution is often referred to as molarity.

Units of molarity: moles per litre (mol/ l).It are also denoted as *M*.

Molar mass, concentration in grams per litre and molarity are related by:

$$\text{Molarity} = \frac{\text{concentration g/l}}{\text{molar mass}}$$

Calculating molarity given mass and volume

Examples.

- 1 Calculate the molarity of a solution made by dissolving 0.56g of potassium hydroxide to make 200cm³ of solution(K=39,O=16,H=1).

Molar mass of KOH=39+16+1

=56g

56g of KOH contain 1 mole

0.56g of KOH contain $\left(\frac{1}{56} \times 0.56\right)$ moles

=0.01moles

200cm³ of solution contain 0.01moles of KOH

1000cm³ of solution contain $\left(\frac{0.01}{200} \times 1000\right)$ M

=0.05M.

2. Calculate the molarity of a solution prepared by dissolving 4.0g of sodium hydroxide to make 250cm³ of solution(Na=23,H=1,O=16).

Molar mass of NaOH=23+16+1

=40g

40g of NaOH contain 1 mole

4.0g of NaOH contain $\left(\frac{1}{40} \times 4.0\right)$ moles

=0.1moles.

250cm³ contains 0.1moles of NaOH

1000cm³contain $\left(\frac{0.1}{250} \times 1000\right)$ Moles

=0.4M

9. 10.6g of anhydrous sodium carbonate was dissolved in water to make 500cm³ of solution. Calculate the molarity of this solution (Na=23,C=12,O=16).

Molar mass of Na₂CO₃=23x2 +12+16x3

=106g

106g of Na₂CO₃ contain 1 mole

10.6g of Na₂CO₃ contain $\left(\frac{1}{106} \times 10.6\right)$ moles

$$=0.1\text{moles}$$

500cm³ of solution contain 0.1moles of Na₂CO₃

1000cm³ of solution contain $\left(\frac{0.1}{500} \times 1000\right)$ moles of Na₂CO₃

$$= 0.2\text{M}$$

CALCULATING MOLARITY GIVEN MOLES AND VOLUME

Examples.

- 1 Calculate the molarity of a solution containing 0.05moles in 80cm³

80cm³ contain 0.05 moles of solution

1cm³ contain $\left(\frac{0.05}{80}\right)$ moles

1000cm³ contain $\frac{0.05}{80} \times 1000$

$$= 0.625\text{M}$$

- 2 Calculate the molarity of a solution containing 0.01moles in 250cm³.

250cm³ contain 0.01 moles of solution

1cm³ contain $\left(\frac{0.01}{250}\right)$ moles

1000cm³ contain $\left(\frac{0.01}{250} \times 1000\right)$ moles

$$= 0.04\text{M}$$

- 3 Calculate the molarity of a solution containing 0.2moles in 25.0cm³.

25.0cm³ contain 0.2 moles of solution

1cm³ contain $\left(\frac{0.2}{25.0}\right)$ moles

1000cm³ contain $\frac{0.2}{25.0} \times 1000$

$$= 0.8\text{M}$$

4. Calculate the molarity of a solution containing 0.4 moles in 500cm³.

500cm³ contain 0.4 moles of solution

1cm³ contain $\left(\frac{0.4}{500}\right)$ moles

1000cm³ contain $\left(\frac{0.4}{500} \times 1000\right)$ moles

$$= 0.8\text{M}$$

CALCULATING NUMBER OF MOLES GIVEN MOLARITY AND VOLUME.

Examples.

- 1 Calculate the number of moles in 25.0cm³ of 0.2M sodium hydroxide solution.

1000cm³ contain 0.2 moles of NaOH

1cm³ contain $\left(\frac{0.2}{1000}\right)$ moles

25.0cm³ contain $\left(\frac{0.2}{1000} \times 25.0\right)$ moles

$$= 0.005 \text{ moles.}$$

- 2 Calculate the number of moles in 30.0cm³ of 0.4M Sulphuric acid

1000cm³ contain 0.4 moles of H₂SO₄

1cm³ contain $\left(\frac{0.4}{1000}\right)$ moles of H₂SO₄

30.0cm³ contain $\left(\frac{0.4}{1000} \times 30.0\right)$ moles

$$= 0.012 \text{ moles of H}_2\text{SO}_4$$

- 3 Calculate the number of moles in 15.20cm³ of 0.1M sodium carbonate.

1000cm³ contain 0.1 moles of Na₂CO₃

1cm³ contain $\left(\frac{0.1}{1000}\right)$ moles Na₂CO₃

$$15.20\text{cm}^3 \text{ contain } \left(\frac{0.1}{1000} \times 15.20\right) \text{ moles}$$

$$= 0.00152 \text{ moles.}$$

4. Calculate the number of moles in 20.0cm^3 of 0.25M hydrochloric acid.

$$1000\text{cm}^3 \text{ contain } 0.25 \text{ moles of HCl}$$

$$1\text{cm}^3 \text{ contain } \left(\frac{0.25}{1000}\right) \text{ moles HCl}$$

$$20.0\text{cm}^3 \text{ contain } \left(\frac{0.25}{1000} \times 20.0\right) \text{ moles HCl}$$

$$= 0.005 \text{ moles}$$

CALCULATING CONCENTRATION IN GRAMS PER LITRE GIVEN MOLARITY

Examples

- 1 Calculate the concentration in grams per litre of 0.25M Sodium sulphate ($\text{Na} = 23$, $\text{S} = 32$, $\text{O} = 16$)

$$\text{Molar mass of Na}_2\text{SO}_4 = 23 \times 2 + 32 + 16 \times 4.$$

$$= 142\text{g.}$$

$$1 \text{ mole contain } 142\text{g of Na}_2\text{SO}_4$$

$$0.25\text{moles contain } \frac{142}{1} \times 0.25$$

$$= 35.5\text{g l}^{-1}$$

- 2 Calculate the concentration in grams per litre of 0.2M sulphuric acid ($\text{H}=1$, $\text{S} = 32$, $\text{O} = 16$).

$$\text{Molar mass of H}_2\text{SO}_4 = 1 \times 2 + 32 + 16 \times 4.$$

$$= 98\text{g.}$$

$$1 \text{ mole contain } 98\text{g of H}_2\text{SO}_4$$

$$0.2\text{moles contain } \frac{98}{1} \times 0.2$$

$$= 19.6\text{g l}^{-1}$$

- 3 Calculate the concentration in grams per litre of 0.1M sodium carbonate (Na = 23, C = 12, O = 16)

Molar mass of $\text{Na}_2\text{CO}_3 = 23 \times 2 + 12 + 16 \times 3$.

$$= 106\text{g}.$$

1 mole contain 106g of Na_2CO_3

$$0.1\text{mole contain } \frac{106}{1} \times 0.1$$

$$= 10.6\text{g l}^{-1}$$

CALCULATING CONCENTRATION IN GRAMS PER LITRE GIVEN MASS AND VOLUME

Example

- 1 Calculate the concentration in grams per litre of a solution made by dissolving 4.0g of sodium hydroxide to make 250cm^3 of solution.

250cm^3 of solution contain 4.0g of NaOH

$$1000\text{cm}^3 \text{ of solution contain } \frac{4.0}{250} \times 1000\text{g}$$

$$= 16\text{g l}^{-1}$$

2. Calculate the concentration in grams per litre of a solution made by dissolving 3.5g of KOH to make 500cm^3 of solution

500cm^3 of solution contain 3.5g of KOH

$$1000\text{cm}^3 \text{ of solution contain } \frac{3.5}{500} \times 1000\text{g of KOH}$$

$$= 7.0\text{g l}^{-1}$$

Exercise

1. Calculate the concentration in grams per litre of the following solution:

(i) 4.0g of sodium hydroxide in 500cm³ of solution.

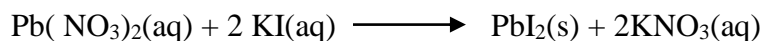
(ii) 0.5g of sodium sulphate in 400cm³ of solution.

(iii) 2.0g of sodium carbonate in 200cm³ of solution.

Calculation of mass of precipitates

Example

1. A solution of lead (II) nitrate reacted with potassium iodide solution according to the equation:



138.3g of lead (II) iodide was precipitated. Calculate the mass of potassium iodide that was required to complete the precipitation. [K=39; I=127; Pb=207]

Formula mass of 2KI = 2(39+127)

$$= 332\text{g}$$

Molar mass of PbI₂ = 207 + 2(127)

$$= 461\text{g}$$

$$\text{Moles of PbI}_2 = \frac{138.3}{461}$$

$$= 0.301\text{moles}$$

Ratio KI : PbI₂ = 2 : 1

Moles of KI = 2 × 0.301

$$= 0.602$$

Mass of KI = 332 × 0.602

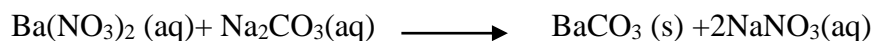
$$= 199.86\text{g}$$

2. Calculate the mass of barium carbonate that will be precipitated when 5.22g of barium nitrate solution are reacted with excess sodium carbonate solution. (C = 12, N = 14, O = 16, Ba = 137)

$$\begin{aligned}\text{Molar mass of Ba(NO}_3)_2 &= 137 + (14 \times 2) + (16 \times 3 \times 2) \\ &= 261 \text{ g}\end{aligned}$$

$$\begin{aligned}\text{Molar mass of BaCO}_3 &= 137 + 12 + (16 \times 3) \\ &= 197 \text{ g}\end{aligned}$$

$$\begin{aligned}\text{Moles of Ba(NO}_3)_2 &= \frac{5.22}{261} \\ &= 0.02 \text{ moles}\end{aligned}$$



$$\text{Mole ratio of Ba(NO}_3)_2 : \text{BaCO}_3 = 1 : 1$$

$$\text{Moles of BaCO}_3 = 0.02 \text{ moles}$$

$$\begin{aligned}\text{Mass of BaCO}_3 &= 0.02 \times 197 \\ &= 3.94 \text{ g}\end{aligned}$$

3. A mass of 11.2g of iron fillings were dropped into a solution of copper (II) sulphate solution. Calculate the mass of copper that would be formed. (Fe =56, Cu =64)

$$\begin{aligned}\text{Moles of Fe} &= \frac{11.2}{56} \\ &= 0.2 \text{ moles}\end{aligned}$$



$$\text{Mole ratio of Fe: Cu} = 1: 1$$

$$\text{Moles of Cu} = 0.2 \text{ moles}$$

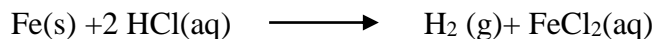
$$\begin{aligned}\text{Mass of Cu} &= 0.2 \times 64 \\ &= 12.8 \text{ g}\end{aligned}$$

4. 50.0cm³ of 3M hydrochloric acid was added into 10g of impure iron fillings. Calculate the percentage purity of iron. (Fe = 56)

1000cm³ contain 3 moles of HCl

1cm³ contain ($\frac{3}{1000}$) moles of HCl

50.0cm³ contain $\frac{3}{1000} \times 50.0$
= 0.15 moles



Mole ratio of Fe : HCl = 1:2

Moles of Fe = $\frac{1}{2} \times 0.15$ moles
= 0.075 moles

Mass of Fe = 0.075 x 56
= 4.2g

% of iron = $\frac{4.2}{10} \times 100$
= 42%

CALCULATING MOLAR MASS AND RELATIVE ATOMIC MASS.

Example.

- 1 Calculate the molar mass of a solution of 0.5M M₂CO₃ made by dissolving 5.02g to make 1 litre of solution.

0.5moles of solution contain 5.02g of M₂CO₃

1mole of solution contain $\frac{5.02}{0.5}$ g of M₂CO₃
= 10.04g

- 2 10.6g of X₂CO₃ was dissolved to make 1 litre of solution. The molarity of the solution was 0.1M. Calculate

(i). the molar mass of X₂CO₃.

0.1moles of solution contain 10.6g of

1mole of solution contain $\frac{10.6}{0.1}$ of X₂CO₃.

$$= 106\text{g.}$$

(ii). the atomic mass of x

$$\text{X}_2\text{CO}_3 = 106$$

$$2x + 12 + 16 \times 3 = 106$$

$$2x + 60 = 106$$

$$2x = 106 - 60$$

$$X = 23.$$

3 0.025M solution of a metal hydroxide, MOH contain 1gram per litre of solution.

Calculate:

(i) The molar mass of MOH

0.025moles contain 1g of MOH

$$\begin{aligned} 1 \text{ mole contain } \frac{1}{0.025} \text{ g of MOH} \\ = 40\text{g.} \end{aligned}$$

(ii) the atomic mass of M.

$$\text{MOH} = 40$$

$$M + 16 + 1 = 40$$

$$M = 40 - 17$$

$$M = 23.$$

MOLARITY AND CALCULATIONS INVOLVING SOLUTIONS.

Examples.

- 1 25.0cm³ of sodium hydroxide solution reacted completely with 20.0cm³ of 0.1M hydrochloric acid. Calculate the molarity of sodium hydroxide solution



1000cm³ of solution contain 0.1 moles HCl

$$20.0\text{cm}^3 \text{ of solution contain } \frac{0.1}{1000} \times 20$$

$$= 0.002 \text{ moles of HCl.}$$

Mole ratio of NaOH : HCl = 1:1

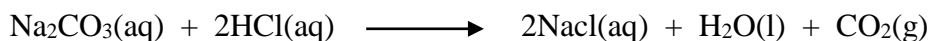
Moles of NaOH = 0.002

25.0cm³ of NaOH contain 0.002 moles.

$$1000\text{cm}^3 \text{ of NaOH contain } \frac{0.002}{25} \times 1000$$

$$= 0.08\text{M}$$

- 2 25.0cm³ of 0.1M Hydrochloric acid react with 20.0cm³ of sodium carbonate solution. Calculate the molarity of sodium carbonate.



1000cm³ of solution contain 0.1 moles HCl

$$25.0\text{cm}^3 \text{ of solution contain } \frac{0.1}{1000} \times 25.0 \text{ moles of HCl}$$

$$= 0.0025 \text{ moles of HCl.}$$

Mole ratio of Na₂CO₃ : HCl = 1:2

Moles of Na₂CO₃ = $\frac{1}{2} \times 0.0025$

$$\text{Moles of Na}_2\text{CO}_3 = 0.00125$$

20.0cm³ of Na₂CO₃ contain 0.00125 moles.

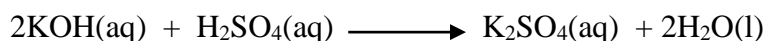
1000cm³ of Na₂CO₃ contain $\frac{0.00125}{20} \times 1000$ moles

$$= 0.0625\text{M}$$

- 3 25cm³ of 0.05M sulphuric acid required 35cm³ of potassium hydroxide solution.

Calculate;

- (a) The molarity of the alkali



1000cm³ of solution contain 0.05 moles H₂SO₄

25cm³ of solution contain $\frac{0.05}{1000} \times 25$ moles of H₂SO₄

$$= 0.00125 \text{ moles}$$

Mole ratio of KOH: H₂SO₄ = 2:1

Moles of KOH = 2 x 0.00125

$$= 0.0025.$$

35cm³ of KOH solution contain 0.0025 moles

1000cm³ of KOH solution contain $\frac{0.0025}{35} \times 1000$ moles.

$$= 0.07\text{M}$$

- (b) The concentration of the alkali in grams per litre

Molar mass of KOH = 39 + 16 + 1

$$= 56\text{g}$$

1 mole of KOH contain 56g

0.07 moles of KOH contain $\frac{56}{0.07}$

$$= 4.0\text{g l}^{-1}$$

CALCULATION OF THE MOLE RATIO.

Examples

- 1 25.0cm^3 of 0.02M hydrochloric acid required 15.0cm^3 of 0.032M alkali solution for complete neutralization.

(a) Calculate;

(i) the number of moles of the acid

1000cm^3 of solution contain 0.02M HCl.

25.0cm^3 of solution contain $\frac{0.02}{1000} \times 25.0$

$= 0.0005$ moles.

(ii) the number of moles of alkali

1000cm^3 of solution contain 0.032moles alkali

15.0cm^3 of solution contain $\frac{0.032}{1000} \times 15.0$

$= 0.00048$ moles.

(b) Determine the mole ratio between the acid to alkali.

HCL : Alkali

0.0005 : 0.00048

$\frac{0.0005}{0.00048}$: $\frac{0.00048}{0.00048}$

1 : 1

2 20cm³ of 0.1M acid HX reacts completely with 5.0cm³ of 0.2M sodium carbonate solution.

(i) the number of moles of the acid

1000cm³ of solution contain 0.1 moles of HX

20cm³ of solution contain $\left(\frac{0.1}{1000} \times 20\right)$

= 0.002 moles

(ii). The number of moles of sodium carbonate

1000cm³ of solution contain 0.2 moles Na₂CO₃

5.0cm³ of solution contain $\left(\frac{0.2}{1000} \times 5.0\right)$

= 0.001 moles

(b) Determine the mole ratio between the acid and sodium carbonate.

HX : Na₂CO₃

$\frac{0.002}{0.001}$: $\frac{0.001}{0.001}$

2 : 1

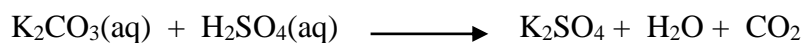
Mole ratio 2: 1

CALCULATION OF PERCENTAGE PURITY.

Example

- 1 28.0g of impure potassium carbonate was dissolved in water to make 1 litre of solution. 20.0cm³ of this solution required 32.0cm³ of 0.1M sulphuric acid solution for complete reaction.

Calculate the percentage purity of potassium carbonate (K = 39, C = 12, O = 16)



1000cm³ of solution contain 0.1mole H₂SO₄

32.0cm³ of solution contain $\frac{0.1}{1000} \times 32.0$

$$= 0.0032 \text{ moles}$$

Mole ratio $\text{K}_2\text{CO}_3 : \text{H}_2\text{SO}_4 = 1:1$

Moles of K₂CO₃ = 0.0032.

20.0cm³ contain 0.0032 moles of K₂CO₃

1000cm³ contain $\left(\frac{0.0032}{20} \times 1000\right)$ moles

$$= 0.16\text{M}.$$

Molar mass of K₂CO₃ = $2 \times 39 + 12 + 16 \times 3$

$$= 138\text{g}.$$

1 mole contain 138g of K₂CO₃

0.16moles contain $\frac{138}{1} \times 0.16$

$$= 22.08\text{g}.$$

$$\% \text{ purity} = \frac{\text{mass of pure carbonate}}{\text{Given mass}} \times 100$$

$$= \frac{22.08}{28.0} \times 100$$

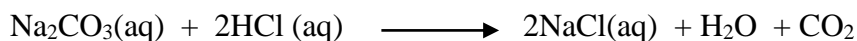
$$= 78.8\%$$

2. 6.4g of an impure sample of anhydrous sodium carbonate was dissolved in water and the solution made up to 500cm³.

25.0cm³ of this solution required 24cm³ of a 0.2M hydrochloric acid solution for complete reaction.

Calculate the;

- (i). Number of moles of the acid that reacted.



1000cm³ of solution contain 0.2moles HCl.

$$24\text{cm}^3 \text{ of solution contain } \left(\frac{0.2}{1000} \times 24 \right) \text{ moles HCl}$$

$$= 0.0048\text{moles}$$

- (ii). Number of moles of the carbonate that reacted

$$\text{Mole ratio } \text{Na}_2\text{CO}_3 : \text{HCl} = 1:2$$

$$\text{Moles of Na}_2\text{CO}_3 = \frac{1}{2} \times 0.0048$$

$$= 0.0024.$$

- (iii) Percentage of the carbonate in the sample

25cm³ of Na₂CO₃ contain 0.0024 moles.

$$500\text{cm}^3 \text{ of Na}_2\text{CO}_3 \text{ contain } \frac{0.0024}{25} \times 500$$

$$= 0.048 \text{ moles.}$$

$$\text{Molar mass of Na}_2\text{CO}_3 = 23 \times 2 + 12 + 16 \times 3$$

$$= 106\text{g.}$$

1 mole contain 106g of Na_2CO_3

$$0.048 \text{ moles contain } \frac{106}{1} \times 0.048$$

$$= 5.088\text{g}$$

$$\% \text{ of } \text{Na}_2\text{CO}_3 = \frac{5.088}{6.4} \times 100$$

$$= 79.5\%$$

3. A mass of 5.0g of a mixture of sodium nitrate and anhydrous sodium carbonate were dissolved in water and the solution made up to 250cm^3 . 25.0cm^3 of this solution required 20.0cm^3 of 0.2M hydrochloric acid for complete neutralization. Calculate the percentage by mass of anhydrous sodium carbonate in the mixture.

(Na = 23, O = 16, C = 12)



1000cm^3 of solution contain 0.2 moles HCl

$$20.0\text{cm}^3 \text{ of solution contain } \left(\frac{0.2}{1000} \times 20.0 \right) \text{ moles}$$

$$= 0.004 \text{ moles.}$$

Mole ratio of Na_2CO_3 : HCl = 1:2

$$\text{Moles of } \text{Na}_2\text{CO}_3 = \frac{1}{2} \times 0.004$$

$$= 0.002 \text{ moles}$$

25.0cm^3 of Na_2CO_3 contain 0.002 moles

$$250\text{cm}^3 \text{ of } \text{Na}_2\text{CO}_3 \text{ contain } \frac{0.002}{0.02} \times 250$$

$$= 0.02 \text{ moles}$$

$$\text{Molar mass of } \text{Na}_2\text{CO}_3 = 23 \times 2 + 12 + 16 \times 3$$

$$= 106\text{g}$$

1 mole contains 106g of Na_2CO_3 .

$$0.02 \text{ moles contain } \frac{106}{1} \times 0.02$$

$$= 2.12\text{g}.$$

$$\% \text{ of } \text{Na}_2\text{CO}_3 = \frac{2.12}{5.0} \times 100$$

$$= 42.4\%.$$

CALCULATION OF NUMBER OF MOLES OF WATER OF CRYSTALLISATION.

Example

1. 8.58g of sodium carbonate, $\text{Na}_2\text{CO}_3 \cdot \text{XH}_2\text{O}$, was dissolved in water to make 250cm^3 of solution. 25.0cm^3 of this solution required 15.0cm^3 of 0.4M hydrochloric acid solution for complete reaction. Calculate the number of moles of water of crystallization in the compound. (H = 1, C = 12, Na = 23, O = 16)



1000cm^3 of solution contain 0.4moles HCl.

$$15.0\text{cm}^3 \text{ of solution contain } \left(\frac{0.4}{1000} \times 15.0 \right) \text{ moles of HCl}$$

$$= 0.006 \text{ moles}$$

$$\text{Mole ratio of } \text{Na}_2\text{CO}_3 : \text{HCl} = 1:2$$

$$\text{Moles of } \text{Na}_2\text{CO}_3 = \frac{1}{2} \times 0.006$$

$$= 0.003\text{moles}$$

25.0cm^3 of Na_2CO_3 contain 0.003 moles.

$$1000\text{cm}^3 \text{ of } \text{Na}_2\text{CO}_3 \cdot \text{XH}_2\text{O} \text{ contain } \frac{0.003}{25.0} \times 1000$$

$$= 0.12\text{M}$$

250cm³ of Na₂CO₃ contain 8.58g

1000cm³ of Na₂CO₃ contain $\left(\frac{8.58}{250} \times 1000\right)$ g

$$= 34.32\text{g/l}$$

0.12 moles of Na₂CO₃ contain 34.32g

1 mole of Na₂CO₃ . xH₂O contain $\frac{34.32}{0.12}$

$$= 286\text{g.}$$

Na₂CO₃ . xH₂O = 286

106 + 18x = 286

$$18x = 286 - 106$$

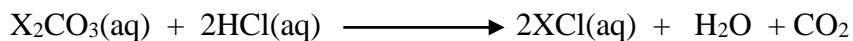
$$18x = 180$$

$$x = \frac{180}{18}$$

$$x = 10.$$

2. 3.43g of a metal carbonate, X₂CO₃.nH₂O was dissolved in water to make 250cm³ of solution. 25.0cm³ of this solution required 30.0cm³ of 0.08M hydrochloric acid for complete reaction. Calculate the value of n in X₂CO₃.nH₂O.

(X=23, C =12, O = 16, H =1).



1000cm³ of solution contain 0.08 moles HCl

30.0cm³ of solution contain $\left(\frac{0.08}{1000} \times 30.0\right)$ moles of HCl

$$= 0.0024\text{moles}$$

Mole ratio of X₂CO₃ : HCl = 1 : 2

$$\text{Moles of } X_2CO_3 = \frac{1}{2} \times 0.0024.$$

$$= 0.0012 \text{ moles}$$

$$25.0 \text{ cm}^3 \text{ of } X_2CO_3 \cdot nH_2O \text{ contain } 0.0012 \text{ moles}$$

$$1000 \text{ cm}^3 \text{ of } X_2CO_3 \cdot nH_2O \text{ contain } \frac{0.0012}{25} \times 1000 \text{ moles}$$

$$= 0.048 \text{ M}$$

$$250 \text{ cm}^3 \text{ of } X_2CO_3 \cdot nH_2O \text{ contain } 3.43 \text{ g}$$

$$1000 \text{ cm}^3 \text{ of } X_2CO_3 \cdot nH_2O \text{ contain } \frac{3.43}{250} \times 1000$$

$$= 13.72 \text{ g/l}$$

$$0.048 \text{ moles contain } 13.72 \text{ g}$$

$$1 \text{ moles contain } \frac{13.72}{0.048}$$

$$= 286$$

$$X_2CO_3 \cdot nH_2O = 286.$$

$$23 \times 2 + 12 + 16 \times 3 + 18n = 286$$

$$106 + 18n - 106 = 286 - 106$$

$$18n = 180$$

$$n = \frac{180}{18}$$

$$\text{the value of } n = 10$$

- 3 When 3.22g of a hydrated sodium sulphate, $Na_2SO_4 \cdot nH_2O$ was heated until there was no further change, 1.42g of the residue remained.

- (a). Determine the value of n in $n Na_2SO_4 \cdot nH_2O$

$$\text{Mass of water} = 3.22 - 1.42$$

$$= 1.8\text{g}$$

Molar mass of $\text{Na}_2\text{SO}_4 = 142$

Molar mass of water = 18

	Na_2SO_4	$n\text{H}_2\text{O}$
	1.42	1.8
Moles	$\frac{1.42}{142}$	$\frac{1.8}{18}$
	0.01	0.1
Ratio	$\frac{0.01}{0.01}$	$\frac{0.1}{0.01}$
	1	10

$$n = 10$$

(b) Calculate the percentage of water of crystallization

$$\frac{1.8}{3.22} \times 100$$

$$= 55.9\%.$$

Exercise

1. A mass of 5.0g of a mixture of sodium nitrate and anhydrous sodium carbonate were dissolved in water and the solution made up to 250cm^3 . 25.0cm^3 of this solution required 20.00cm^3 of 0.2M hydrochloric acid solution for complete reaction. Calculate the percentage by mass of anhydrous sodium carbonate in the mixture. (Na = 23, C = 12, O = 16).
2. A solution of ethanedioic acid, $\text{H}_2\text{C}_2\text{O}_4 \cdot x\text{H}_2\text{O}$ was made by dissolving 10g of acid crystals to make 500cm^3 of solution. 25.0cm^3 of this solution required 15.90cm^3 of 0.5M potassium hydroxide solution for complete reaction. Calculate the number of moles of water of crystallization in the acid crystals. (H = 1; C = 12; O = 16)

3. When excess dilute hydrochloric acid was added to 7.8g of a mixture of calcium carbonate and calcium sulphate 896cm^3 of carbon dioxide measured at s.t.p was evolved.

(a) Write an equation for the reaction that took place

(b) Determine the

(i) Mass of calcium carbonate in the mixture.

(ii) Percentage of calcium carbonate in the mixture.(C=12; O=16,Ca=40,1 mole of gas at s.t.p occupies 22.4dm^3).

4. 25.0cm^3 of 0.1 M hydrochloric acid react with 20.0cm^3 of sodium hydroxide solution.

Calculate the molarity of sodium hydroxide solution.

5. 100cm^3 of hydrochloric acid dissolves 3g of magnesium ribbon. Calculate the molarity of the acid.(Mg=24)

6. 25.0cm^3 of a solution containing 5.3g per litre of a metal carbonate M_2CO_3 , neutralized 25cm^3 of 0.1M hydrochloric acid. Calculate the;

(i) molarity of the carbonate solution.

(ii) formula mass of the metal carbonate.

(iii) Relative atomic mass of the metal, M in the metal carbonate.

7. 25cm^3 of impure sodium hydroxide solution containing 5.0g of impure sodium hydroxide per litre, completely reacted with 20cm^3 of 0.1M hydrochloric acid. Calculate the

(i) Molarity of sodium hydroxide solution.

(ii) Percentage impurity in the sample of sodium hydroxide.

8. 20cm^3 of 0.12M potassium hydroxide required 24cm^3 of a solution of a dibasic acid H_2X containing 1.225g per 250cm^3 of solution. Calculate the formula mass of the acid.

9.(a) 3.4g of a mixture containing calcium carbonate and calcium sulphate were reacted with excess hydrochloric acid and 448cm^3 of carbon dioxide measured at s.t.p was evolved.

Write equation(s) for the reaction(s)

(b) Calculate the

(i) Mass of calcium carbonate in the mixture

(ii) Percentage of calcium carbonate in the mixture.

10. In an experiment, 25cm^3 of sodium hydroxide solution were completely neutralized by 25cm^3 of 0.2M sulphuric acid.

(a) Write an equation for the reaction between sodium hydroxide solution and sulphuric acid.

(b) Calculate the number of moles of

(i) Sulphuric acid that reacted

(ii) Sodium hydroxide that reacted

(c) Calculate the concentration in

(i) Moles per litre of sodium hydroxide solution.

(ii) Grams per litre of sodium hydroxide solution.

11.5.2g of zinc granules were dropped into 16.0g of a solution of copper(II) sulphate solution.

(a) Write the equation for the reaction.

(b) Calculate :

(i) the mass of zinc sulphate that will be formed.

(ii) the mass of copper that precipitates out from the solution.

(iii) the number of moles of copper(II) sulphate that remained un reacted. (O=16; S=32; Cu=64; Zn=65).

EFFECTS OF AN ELECTRIC CURRENT ON SUBSTANCES.

Common terms used.

1. **Electrolysis:** The decomposition of an electrolyte by the passage of an electric current through it.
2. **Electrolyte:** A compound which when in solution or molten state conducts electricity and is decomposed by it.

Types of electrolytes

(b).Strong electrolytes: these are electrolytes which ionize completely in solution.

Examples

Salts- NaCl, KCl, MgSO₄

Mineral acids – HCl, H₂SO₄, HNO₃

Caustic alkalis: NaOH, KOH

(b).Weak electrolytes: these are electrolytes which ionize only slightly while in solution, producing a mixture of molecules and ions.

Examples

- Ammonia solution
- Most organic acids e.g. Ethanoic acid
- Water

3. **Non electrolytes:** these exist only in form of molecules and are not capable of ionizing in molten form and in solution.

Examples

- Sugar
- Ethanol

- Methyl benzene

Non electrolytes are substances which when in solution or molten state do not conduct electric current.

4. Conductors: These are substances which allow passage of an electric current through them.

Metals and graphite conduct electricity due to the presence of mobile electrons in their structures.

Ionic compounds conduct electricity because of free mobile ions which carry charge.

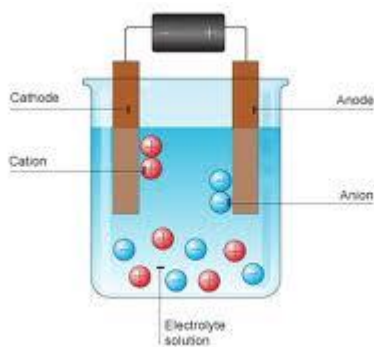
Differences between a conductor and an electrolyte.

Conductor	Electrolyte
Conducts electricity by movement of electrons.	Conducts electricity by movement of ions.
It is not decomposed by electric current.	It is decomposed by passage of an electric current.
They are mainly metallic solids except graphite.	They are ionic compounds.

5. Non conductors (insulators): these are substances that do not allow electricity to pass through them e.g glass, dry wood, rubber and plastic. They are used for making handles of electrical appliances and as cover for power cables and telephone lines.

6. Electrodes: these are two pieces of carbon (graphite) or metal by which current enters or leaves an electrolyte.

Diag.



Anode-a positive electrode at which electrons enter the external circuit.

Cathode-a negative electrode at which electrons leave the external circuit and enter the electrolyte.

MIGRATION OF IONS

Ions are electrically charged particles. The particles can be either positively charged (cations) or negatively charged (anions).

Migration of ions is the movement of ions towards electrodes during the passage of an electric current.

Negative ions (anions) move towards the anode. At anode, anions lose electrons to become neutral atoms. Therefore, the reaction at the anode is oxidation reaction.

Positive ions (cations) move towards the cathode. At the cathode, cations gain electrons to become neutral atoms. Therefore, the reaction at the cathode is reduction.

Discharge

This is the formation of neutral atoms at electrodes as the ions gain or loses electrons.

Preferential discharge: This is the choice of ion to be discharged at each electrode when there are more than one type of positive or negative ions at the respective electrodes.

Cations



Anions





Note: The nitrate (NO_3^-), Sulphate(SO_4^{2-}) and carbonate(CO_3^{2-}) ions are never discharged at anode. Hydroxyls (OH^-) are preferentially discharged instead.

Factors affecting preferential discharge

When two or more ions of similar charges are present in solution e.g Na^+ and Ag^+ or OH^- and SO_4^{2-} , one is preferentially selected for discharge. The preferential discharge depends on the following factors:

1. Position in the electrochemical series.

The ion lower in the electrochemical series is preferentially discharged. This is because they are less stable as ions.

2. Concentration

This factor affects only anions. The anion present in a much greater concentration is discharged in preference to the ones less concentrated.

3. Nature of electrodes

There are two types of electrodes.

- Inert electrodes (graphite or platinum) which do not take part during electrolysis.
- Active electrodes (metallic e.g copper) normally take part during electrolysis by “going into solution”. When used, they affect the reactions at electrodes.

Some metal electrodes undergo oxidation more readily than anions in solution e.g copper. When used anode dissolves instead of liberating a gas.

HALF EQUATIONS

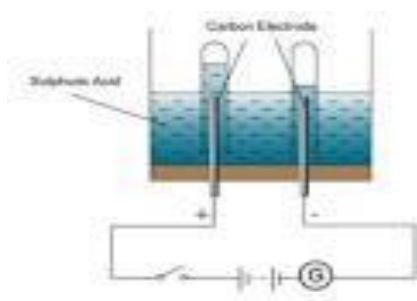
These are equations showing the reactions taking place separately at the electrodes.

Overall equation: An equation that combines the half equations at the electrodes thus representing the overall redox (reduction-oxidation) reaction.

Role of water in electrolysis

- It breaks down crystal lattice of an ionic compound and ions become mobile.
- It reacts with some molecular compounds to form ions.
- It provides H^+ ions and OH^- ions in solution.

Electrolysis of dilute sulphuric acid (Acidified water)



Ions present:

From sulphuric acid (H_2SO_4) are 2H^+ and SO_4^{2-}

From water (H_2O) are H^+ and OH^-

Reactions at electrodes:

Reaction at anode:

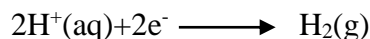
Observation: Effervescence occurs and colourless gas evolved.



Reaction at cathode:

Observation: Effervescence occurs and a colourless evolved.

Hydrogen ions migrate to the cathode, gain electrons and become hydrogen gas.



Overall equation:



Note: The volume of hydrogen gas produced at the cathode is twice the volume of oxygen gas produced at the anode, because oxygen is more soluble in water than hydrogen gas and decreases in volume.

PH; no change in PH of the solution.

Electrolysis of concentrated sodium chloride solution (brine)

Ions present:

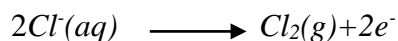
From sodium chloride (NaCl) are Na^+ and Cl^-

From water (H_2O) are H^+ and OH^-

Reactions at electrodes:

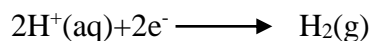
Reaction at anode:

Observation: Effervescence occurs and greenish-yellow gas evolved.



Reaction at cathode

Observation: Effervescence occurs and a colourless gas is evolved.



PH: the resulting solution has a PH greater than 7 due to accumulation of hydroxyl ions.

Electrolysis of concentrated solution of copper (II) chloride

(Using copper cathode and graphite anode)

Ions present:

From copper (II) chloride (CuCl_2) are Cu^{2+} and Cl^-

From water (H_2O) are H^+ and OH^-

Reaction at electrodes:

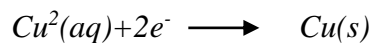
Reaction at anode:

Observation: Effervescence occurs and a greenish-yellow gas is evolved.

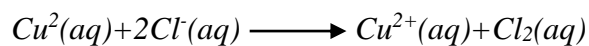


Reaction at cathode:

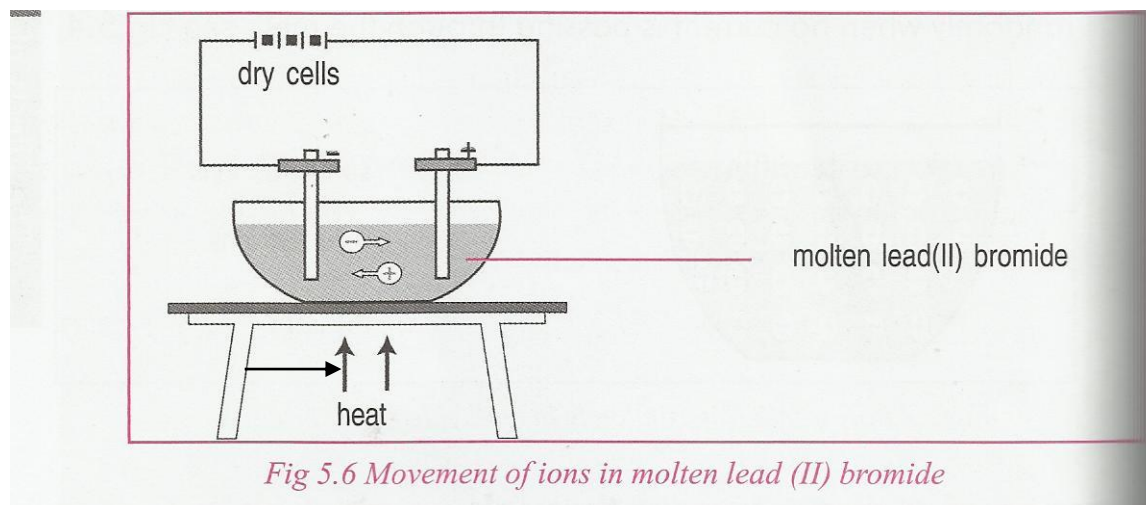
Observation: A brown solid is deposited.



Overall equation:



Electrolysis of molten lead (II) bromide



When molten lead (II) bromide is in solid form, it does not conduct electricity because there are no ions in it which would conduct electricity. When molten the ions are free to move which conduct electric current and the bulb lights.

Ions present

From molten lead (II) bromide(PbBr_2) are Pb^{2+} and 2Br^-

Reaction at electrodes:

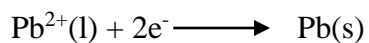
Reaction at anode:

Observation: The reddish-brown fumes are formed at the anode.

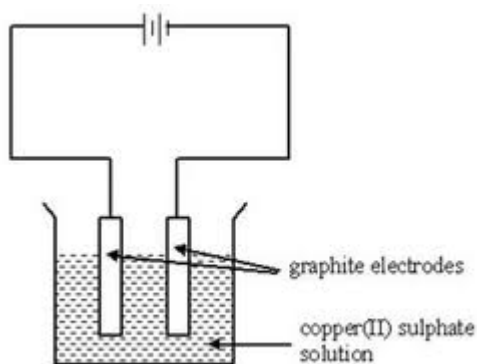


Reaction at cathode

Observation: Shiny-grey beads are formed at the cathode.



Electrolysis of copper (II) sulphate solution using graphite electrodes



Ions present:

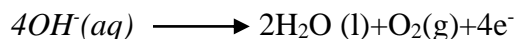
From copper (II) sulphate(CuSO_4) are Cu^{2+} and SO_4^{2-}

From water (H_2O) are H^+ and OH^-

Reaction at electrodes:

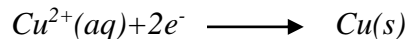
Reaction at anode

Observation: Effervescence occurs and a colourless gas is evolved.



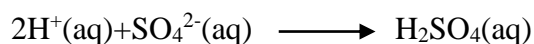
Reaction at cathode:

Observation: brown solid is deposited at the cathode.



The blue solution gradually turns colourless due to discharge of copper(II) ions.

PH: the final solution becomes acidic due to undischarged ions.



Electrolysis of copper (II) sulphate solution using copper anode and graphite cathode.

Ions present:

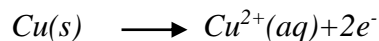
From copper (II) sulphate(CuSO_4) are Cu^{2+} and SO_4^{2-}

From water (H₂O) are H⁺ and OH⁻

Reaction at electrodes:

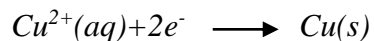
Reaction at anode

Observation: The anode dissolves to form copper (II) ions. the anode decreases in mass(size).



Reaction at cathode:

Observation: brown solid is deposited at the cathode.



The blue colour of the electrolyte remains the same because copper(II) ions from the solution are not discharged.

The table below gives examples of electrolytes and the products at each electrode when each is electrolysed.

Electrolyte	electrode	At anode	At cathode
Dilute hydrochloric acid	platinum	1 volume of chlorine	1 chlorine
Dilute sulphuric acid	platinum	1 volume of oxygen	2 volumes of hydrogen
Copper chloride	carbon	Chlorine	Copper deposited
Molten lead(II) bromide	carbon	Bromine	Lead

Laws of electrolysis

The laws of electrolysis were stated by Faraday. According to his laws the amount of the substance produced during electrolysis depends on:

- (i) time of passing a steady current.
- (ii) Magnitude of the steady current.

(iii) Charge on the ion of the element.

Faraday's first law

It states that the mass of a substance produced at an electrode during electrolysis is proportional to the quantity of electricity passed through it.

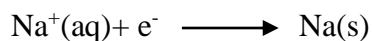
Quantity of electricity = Current x time

$$Q = It$$

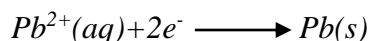
Faraday

This is the quantity of electricity required to deposit one mole of a substance from an ion with a single charge.

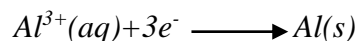
One mole of sodium ions require one faraday (1 mole of electrons) to discharge at the cathode.



One mole of lead ions require 2 faradays (2 moles of electrons) to discharge at the cathode.



One mole of aluminium ions require 3 faradays (3 moles of electrons) to discharge at cathode.



One mole of electrons is equivalent to one Faraday. One Faraday = 96500 Coulombs.

$$1\text{e}^- = 1\text{F} = 96500\text{C}$$

Faraday's second law

It states that the mass of a substance deposited at the electrodes is inversely proportional to the charge on its ion.

Example 1: Calculating quantity of electricity used

A current of 3A was passed through a solution of zinc sulphate for 10 minutes.

Calculate the quantity of electricity used.

$$Q=It$$

$$Q=3 \times 10 \times 60$$

$$=1800C$$

Example2: Calculating the charge on the metal

When a current was passed through a solution of silver nitrate and a solution of metal Y, 4.32g of silver and 2.24g of Y were formed respectively. Calculate the charge on Y (Ag=108, Y=112)

108g of silver contains 1mole.

$$4.32g \text{ of silver contain } \frac{1}{108} \times 4.32$$

$$=0.04 \text{ moles}$$

112g of Y contains 1mole

$$2.24g \text{ of Y contain } \frac{1}{112} \times 2.24$$

$$=0.02 \text{ moles}$$

Mole ratio Ag :Y

$$0.04 \quad 0.02$$

$$\frac{0.04}{0.02} \quad \frac{0.02}{0.02}$$

$$2:1$$

The charge on Y is 2+. This is because the number of moles deposited is inversely proportional to the number of charges on ion.

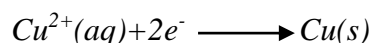
Example 3. Calculating the relative atomic mass

When a current of 0.45A was passed through a solution of copper (II) sulphate for 25minutes, 0.222g of copper were deposited. Calculate the relative atomic mass of copper

$$Q=It$$

$$=0.45 \times 25 \times 60$$

$$=675\text{C}$$



2moles of electrons are required to produce 1mole of Cu.

$$2 \times 96500 = 193000\text{C}$$

675C deposit 0.222g of copper

$$193000\text{C deposit } \left(\frac{0.222}{675} \times 193000 \right) \text{ g of Cu}$$

$$=63.5$$

Example 4: Calculating mass of substance deposited at cathode.

A concentrated solution of copper (II) chloride was electrolysed using copper cathode and graphite anode. State what was observed at the;

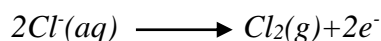
(i) **Anode**

Effervescence occurs and a greenish-yellow gas evolved.

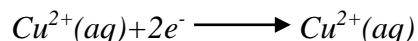
(ii) **Cathode**

Brown solid is deposited

(iii) Write equation for the reaction at anode electrode



(iv) Calculate the mass of the product at the cathode when 3240C of electricity is passed through the electrolyte ($\text{Cu}=64, 1\text{Faraday}=96500\text{C}$).



2moles of electrons are required to produce 1mole of copper.

$$2 \times 96500 = 193000\text{C}$$

193000C produce 64g of Cu

$$3240C \text{ produce } \left(\frac{64}{193000} \times 3240 \right) g$$

$$= 1.07g \text{ of Cu.}$$

Example 5: Calculating volume of gas liberated at electrodes.

1. A current of 6.0A was passed for 3 minutes during electrolysis of acidified water. Calculate the volume of oxygen gas evolved at s.t.p during the process of electrolysis (1 mole of gas occupies 22.4dm³ at s.t.p, 1 Faraday = 96500C).

$$Q = It$$

$$6 \times 3 \times 60$$

$$= 1080C$$



moles of electrons are lost to produce 1 mole of oxygen

$$4 \times 96500 = 386000C$$

386000C liberate 22.4dm³ of O₂

$$1080C \text{ liberate } \left(\frac{22.4}{386000} \times 1080 \right) dm^3$$

$$= 0.063 dm^3 \text{ of oxygen}$$

2. A current of 4.2A was passed for 90 minutes between graphite electrodes immersed in copper (II) sulphate solution at room temperature (1 Faraday = 96500C, O = 16, 1 mole of gas at room temperature occupies = 24dm³).

- (i) Calculate the mass of oxygen gas liberated.
- (ii) Calculate the volume of oxygen gas liberated at room temperature.

Applications of electrolysis

1. Purification of metals such copper

diag

In copper refining the impure copper is made the anode while a pure strip of copper is made the Cathode. The electrodes are immersed in copper (II) sulphate solution.

The impure copper anode dissolves on passing electric current and forms copper ions in solution.

The ions are discharged and deposited at the cathode as pure copper.

Impurities either dissolve in electrolyte or are deposited as sludge around the anode.

At anode: $Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^{-}$

At cathode: $Cu^{2+}(aq) + 2e^{-} \longrightarrow Cu(s)$

Uses of copper

1. Electrical wiring.
2. Making kettles, boilers, ornaments.
3. Making alloys e.g. brass and bronze.
4. Making coins since it is not easily attacked by air.
5. Making water pipes and boilers.

2. Electroplating

This the process of coating one metal with another metal by means of electrolysis. In

electroplating the metal to be plated is made the cathode and the plating metal is made the anode.

Electroplating is used to improve the beauty of an article (silver), prevent material from corrosion (silver and chromium), give a material hard surface (chromium), give thickness to worn parts of machinery (nickel).

3. Extraction of metals

More reactive metals such as sodium, potassium and aluminium are extracted by electrolysis of their molten compounds.

ELECTROCHEMISTRY

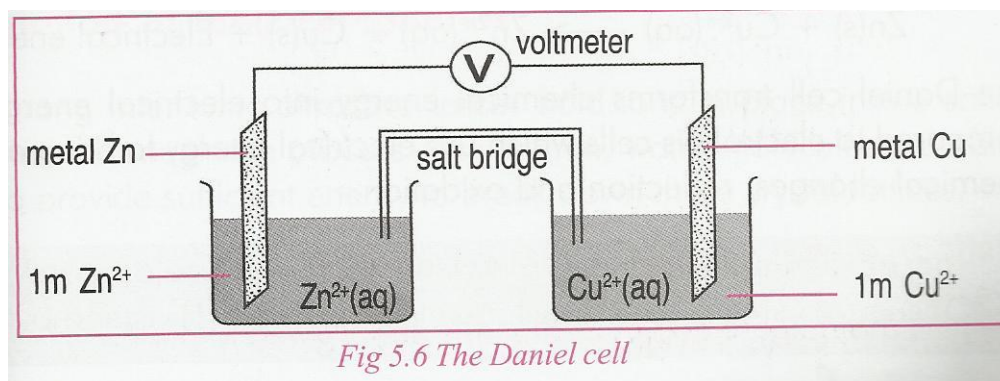
Electrochemical cells (Primary cells)

These are cells designed to produce electricity by chemical reactions. They basically consist of two half-cells of different elements. Each element is dipped in a solution of its own ions and is called an electrode.

The element that gives off its electron more readily (more reactive element) is the cathode. The electrode that accepts electrons more readily (less reactive element) is the anode.

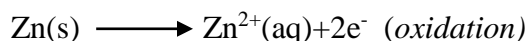
The Daniell cell

The Daniell cell consists of zinc rod dipped in zinc sulphate solution (zinc half-cell) and copper rod dipped in copper (II) sulphate solution (copper half-cell), the solutions being separated by porous partition (salt bridge). The porous partition completes the circuit without allowing the solutions to mix. The metal rods are connected by use of a wire. The metal wire forms external circuit along which the liberated electrons flow from zinc to copper. These electrons are accepted by the copper ions of copper (II) sulphate solution to form copper metal.



Reaction at the zinc half-cell (cathode)

Zinc rod dissolves forming zinc ions in solution.



- Excess electrons are deposited on the zinc electrode making it the negative terminal.
- However, the zinc is the anode since this is where oxidation takes place.

The zinc half cell can be represented as $\text{Zn(s)} / \text{Zn}^{2+}(\text{aq})$.

Reaction at the copper half-cell (anode)

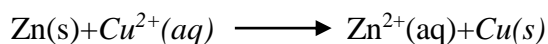
- Copper (II) ions are discharged at the copper electrode.

$$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Cu(s)} \quad (\text{reduction})$$
- The copper electrode is the anode (positive terminal).
- The copper electrode is the cathode since reduction takes place at this electrode.

The copper half cell can be represented as $\text{Cu}^{2+}(\text{aq}) / \text{Cu(s)}$

Overall cell reaction

The overall cell reaction is derived by adding the two half equations. The electrons on either sides of the two equations cancel out.



The cell convention for the overall reaction is



The Daniel cell changes chemical energy into electrical energy compared to secondary cells which use electrical energy to bring about chemical changes.

Cell notation (cell diagram)

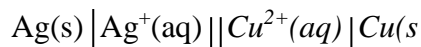
The cell representation shows the sequence in which the reactions proceed.



The two vertical lines(\parallel) represent a salt bridge that links the two cell electrolytes. The vertical line between Zn(s) and $\text{Zn}^{2+}(\text{aq})$ represents a phase boundary between the zinc electrode and its aqueous ions. The vertical line between Cu(s) and $\text{Cu}^{2+}(\text{aq})$ represents the phase boundary

between the copper electrode and the copper ions in the electrolyte. In the cell representation, the cathode is always on the left –hand side and anode on the right –hand side.

The cell diagram for an electrochemical cell is shown below:



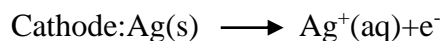
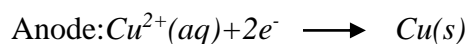
(i) Name two substances that could be used as electrodes.

Silver nitrate solution and copper (II) sulphate solution.

(ii) State which electrode is the anode.

Copper electrode

(iii) Write equation for the reaction at



(iv) Write equation for overall equation



Practical uses of electrochemical cells

1. They are used in making dry cells.
2. Car batteries are made up of a set of electrochemical cells.

Exercise

1.(a) Draw a diagram of a Daniell cell consisting of a zinc rod dipped in zinc sulphate and a copper rod dipped in copper sulphate solution. The solutions separated by a porous wall and the rods connected by a wire.

(b) Indicate

- (i) the charges on each electrode
- (ii) the direction of electron movement in the wire.

(c) Write

(i) equations for the reactions at each electrode.

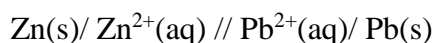
(ii) an equation for the overall reaction.

2. During electrolysis of acidified water using graphite electrodes, gas Y was produced at anode and gas X was liberated at cathode.

(a) (i) Name the gases X and Y that are evolved during electrolysis.

(ii) Give equation for the reaction occurring at anode.

3. The cell convention for an electrochemical cell is shown below.



(a) Name the two substances that could be used as electrolytes.

(b) State which one of the electrodes is the anode.

(c) Write equation for the reaction at

(i) The anode

(ii) The cathode

3 A current of 0.25A was passed through copper(II) sulphate solution for 40 minutes .Calculate the

(a) quantity of electricity used.

(b) quantity of electricity required to deposit one mole of copper.

(c) mass of copper deposited at the cathode.

(d) moles of copper deposited.[Cu=64,1F=96500C].

(b) (i) Calculate the volume of gas X produced when a current of one ampere flows for 10 minutes through the electrolyte.(1F=96500C, 1 mole of gas occupies 24000 cm³ at room temperature).

(ii) State two industrial applications of electrolysis.

4 The cell diagram for an electrochemical cell is shown below.



(a) Name the substance that could be used as an electrolyte at the

(i) Anode

(ii) cathode

(b) Write the half-cell equation for the reaction at the

(i) Anode

(ii) Cathode

(c) Write equation for the overall cell reaction.

EXTRACTION OF METALS

Metals are generally reactive elements, so they occur as compounds (ores) containing low percentage of the metals.

The method of extraction depends on the reactivity of the metal.

Highly reactive metals such as sodium and aluminium are extracted by electrolysis of their molten chlorides or oxides. This is because they cannot be reduced by common reducing agents such as carbon.

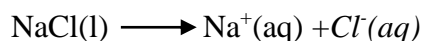
Less reactive metals such as iron, zinc, copper and lead are extracted by reducing their oxides using a more reactive element, usually carbon.

Extraction of sodium

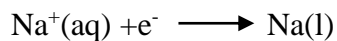
The main ores of sodium are sodium carbonate, sodium nitrate and sodium chloride (common salt). Sodium metal is extracted from its main ore sodium chloride.

The process

The solid sodium chloride is poured into the Down's cell and calcium chloride added to lower the melting point of sodium chloride from 800°C to 600°C . Electric current is then passed through the electrolyte where it melts to form sodium ions and chloride ions.

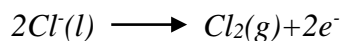


The sodium ions move to the iron cathode and are discharged by gain of electron.



The molten sodium is collected under dry nitrogen in iron tank to prevent it from reacting with air.

The chloride ions move to the carbon anode and are discharged by loss of electrons.



The iron cathode and carbon anode are separated by iron gauze diaphragm which ensures that chlorine does not react with hot sodium.

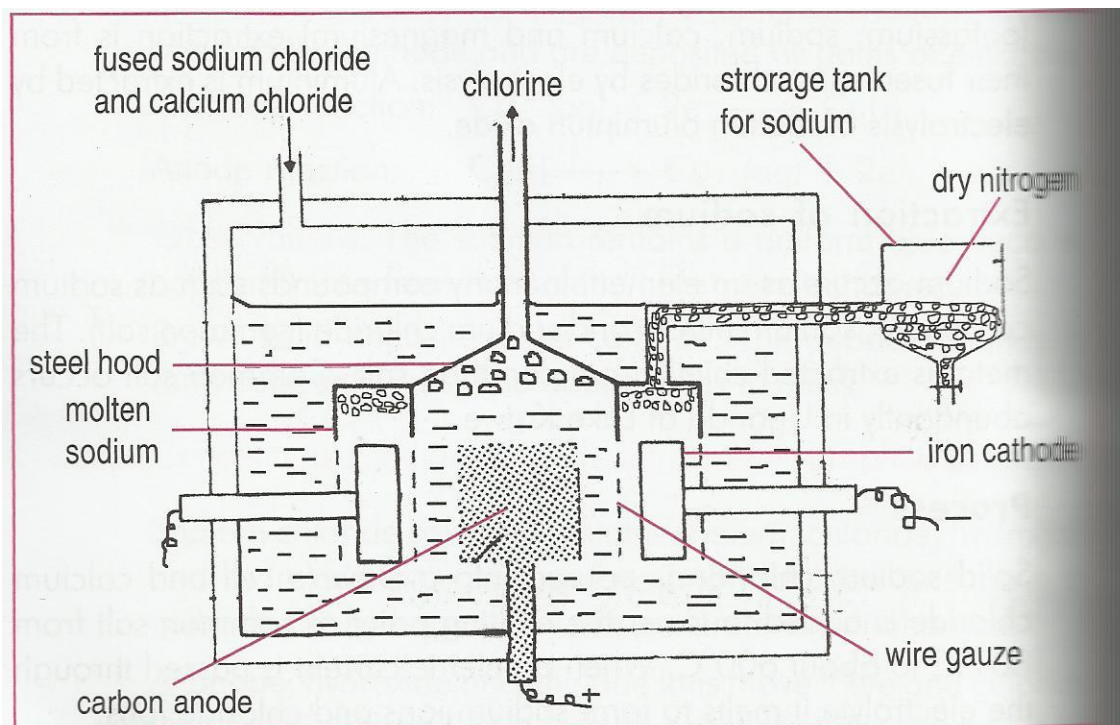


Fig 5.10 Extraction of sodium by Down's cell

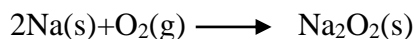
Physical properties of sodium

- It is a white-silvery shiny metal. It rapidly tarnishes in air due to oxidation.
- It is a relatively soft metal that can be cut with a metal.
- It is less dense than water and hence floats on water when reacting.
- It conducts both heat and electricity.

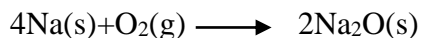
Chemical reactions of sodium

1. Reaction with air.

Sodium reacts with air when heated and burns with a yellow flame to form a yellow solid of sodium peroxide.



Sodium burns in limited air to form a white solid of sodium oxide.



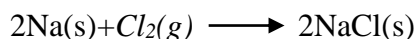
2. Reaction with water

Sodium floats and darts on the surface of water as it melts giving off hydrogen gas.



3. Reaction with chlorine

When heated in chlorine it burns brightly forming white fumes of sodium chloride..



Uses of sodium

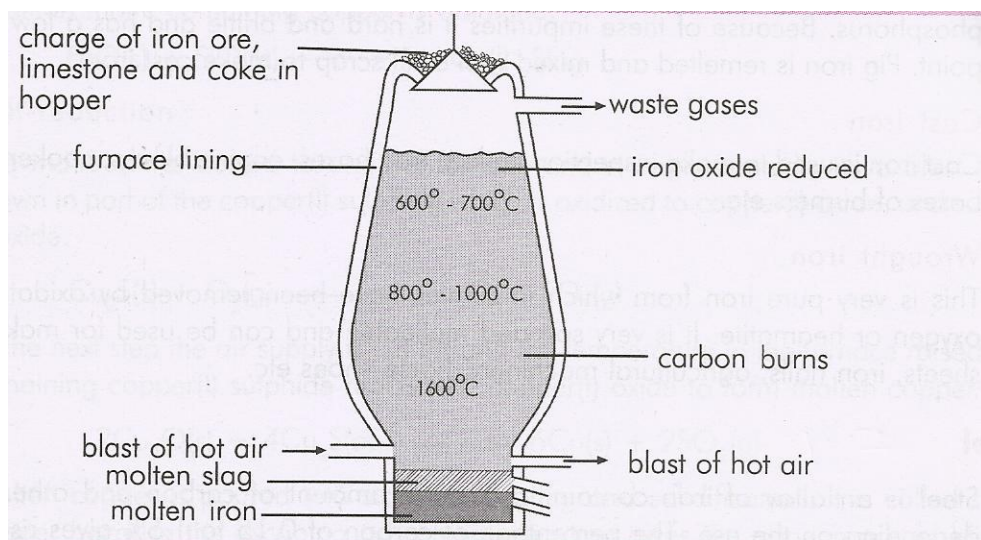
- It is used in vapour lamps producing deep yellow light for street lighting.
- Sodium –potassium alloy is used as a coolant in nuclear reactors.
- Sodium-lead alloy is used in the manufacture of antiknock additives for petrol.
- Sodium is used in the manufacture of sodium cyanide for gold extraction.
- Sodium is used in the manufacture of sodium peroxide.

Extraction of Iron

The main ores of iron are;

- Haematite, Fe_2O_3
- Magnetite, Fe_3O_4
- Siderite(Spathic iron), FeCO_3
- Iron pyrite, FeS_2

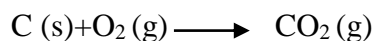
Blast furnace



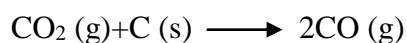
Iron is extracted by reduction from its main ore Haematite, Fe_2O_3 in a blast furnace lined with fire bricks. The raw materials, Haematite, coke and limestone are fed into the furnace from the top.

The role of coke

At the bottom of the furnace, coke burns in hot air to produce carbon dioxide.

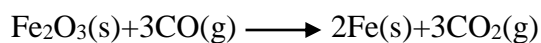


The carbon dioxide produced is reduced by unburnt coke in the middle of the furnace to produce carbon monoxide.



The role of Haematite

The carbon monoxide produced reduces the Haematite to iron.

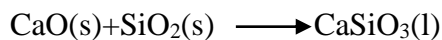


The role of limestone

Lime stone is used in the removal of impurities. Limestone decomposes in the heat to form quick lime (calcium oxide).



The quick lime combines with silica (SiO_2) to form slag (calcium silicate).



The slag sinks to the bottom and floats on iron, protecting the iron from being re-oxidized hot air.

Physical properties of iron

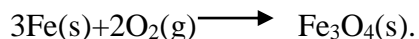
- It is silver-grey when freshly cut.
- It is malleable and ductile.
- It has high melting point of about 1525°C .
- It is a good conductor of heat and electricity.

Chemical reactions of iron

1. Reaction with air

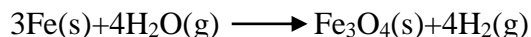
When exposed in moist air, it readily rusts forming a reddish-brown solid called hydrated iron(III) oxide, $\text{Fe}_2\text{O}_3 \cdot x \text{H}_2\text{O}$.

Iron filings of iron wire, when heated in oxygen, burn to form a blue-black solid of tri-iron tetraoxide, Fe_3O_4 .



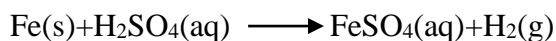
2. Reaction with water

Iron does not react with water, but reacts with excess steam to form a black solid of tri iron tetra oxide and giving off hydrogen.

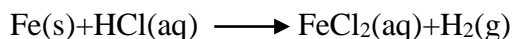


3. Reaction with acids

- Iron reacts with dilute sulphuric acid to form a green solution of iron (II) sulphate.

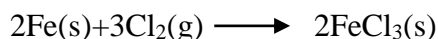


- Iron reacts with dilute hydrochloric acid to form a green solution of iron(II) chloride and hydrogen gas.



4. *Reaction with chlorine.*

When iron wire is heated in a stream of dry chlorine, it glows forming a black sublimate.



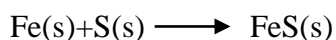
5. *Reaction with hydrogen chloride*

When dry hydrogen chloride is passed over heated iron wire, a white anhydrous iron(II) chloride is formed.



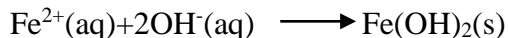
6. *Reaction with sulphur*

When a mixture of iron filings and sulphur is heated in a hard glass tube, the mixture glows red hot and a black solid of iron(II) sulphide is produced.



Tests for iron (II) and iron (III) ions.

1. When a solution of sodium hydroxide or ammonia solution is added to a solution of iron salt, a dirty green precipitate insoluble in excess indicates the presence of Fe^{2+} ion. The dirty green precipitate is iron(II) hydroxide.



2. If a red-brown precipitate insoluble in excess alkali is formed, it indicates the presence of Fe^{3+} ion.

Uses of iron

- For making of railway lines.
- Making saw blades
- Making roofing nails
- Making roofing sheets
- Construction of mechanical structures such as bridges, car bodies and frames.

Alloys

An alloy is a uniform mixture of the metal with one or more other elements.

Alloys and their uses

Alloy	Main components	Uses
1.Solder	Lead and tin (low melting point)	<ul style="list-style-type: none"> • Making electrical contacts. • Joining pipes.
2.Brass	Copper and zinc (great resistance to corrosion)	In electrical connection,jewellery,screws,light Bulb caps and condenser tubes.
3.Bronze	Copper and tin	<ul style="list-style-type: none"> • Making copper coins • Making clock springs • Making electrical contacts.
4.Steel	Iron and carbon	<ul style="list-style-type: none"> • Making car bodies • Making girders • Making railway lines • In building industry.
5.Duralumin	Aluminium, copper and magnesium (vey light and	<ul style="list-style-type: none"> • Making parts of air craft • Making light weight machinery • Making utensils

6.Stainless steel	corrosion resistance) Iron, chromium and nickel (high resistance to corrosion)	<ul style="list-style-type: none"> • Making window frames and panels. <p>Making cutlery, kitchen utensils and razor blade.</p>
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Exercise

5. (a) (i) Name one ore of each of the following metals: Sodium and iron.
- (i) Briefly describe how sodium and iron are extracted from their ores.
- (c) State the conditions under which sodium and iron can react with water. Write equations for the reaction in each case.

- 2.(a) One of the ores from which iron can be extracted is siderite, FeCO_3 .

Name and write the formulae of two other ores from which iron can be extracted.

- (b) Out line the process by which iron metal is obtained from one of the ores you have named in (a). Write equations for the reactions that take place.

- (c) Iron rusts when exposed to moist air. Give three methods by which iron can be prevented from rusting.

3. (a) Name one ore from which sodium can be extracted.

- (b) Describe how sodium is extracted from the ore you have named in (a) .Your description should include the following:

- (i) Names of materials used as the electrodes.
- (ii) Equations for the reactions hat take place at electrodes.
- (iii) Method of collecting the sodium produced.
- (c) A piece of sodium was heated and plunged into a jar of chlorine.
 - (i) State what was observed.
 - (ii) Write the equation for the reaction which took place.

(d) When aqueous silver nitrate was added to a solution of the product in (d) ,a white precipitate was formed. Write an ionic equation for the reaction that took place.

4. (a) Name one ore of iron and write it's formula.

(b) During the extraction of iron, limestone and coke are added into a blast furnace.

Explain the role of

(i) Coke

(ii) Limestone

Use equations to illustrate your answer.

(c) Write equation for the reaction leading to the formation of iron(II) sulphate.

(d) Iron (II) sulphate was heated strongly.

(i) State what was observed.

(ii) Write the equation for the reaction.

5. (a)(i)What is an alloy.

(ii) Give an example of an alloy.

(iii) State the composition of the alloy you have named in(a).

(b) State two uses of the alloy in a(ii).

ENERGY CHANGES IN CHEMICAL AND PHYSICAL PROCESSES

Energy is the capacity of overcoming resistance or doing work. It is a property of matter, it can do work and can be obtained from work. Chemical changes are normally accompanied by energy changes. The most common form of energy change in chemical reactions is the heat change.

Endothermic and exothermic reactions.

1. ***Endothermic reactions.*** These are reactions during which heat is absorbed from the surroundings. The temperature of the surroundings decrease. The products are at a higher energy levels than the reactants and the enthalpy change is positive ($+\Delta H$).

Examples of endothermic reactions.

- (i).Dissolving ammonium nitrate in water.
- (ii).Dissolving sodium thiosulphate in water.

2. ***Exothermic reactions.*** These are reactions during which heat is liberated to the surroundings. The temperature of the surroundings increase. The products are at a lower energy levels than the reactants. And the enthalpy change is negative (ΔH).

Examples of exothermic reactions

- (i).Adding concentrated sulphuric acid to little water in a test tube.
- (ii).Reacting dilute sulphuric acid with magnesium ribbon.
- (iii).Reacting sodium hydroxide and dilute hydrochloric acid.
- (iv).Burning magnesium in air.

Enthalpy

The energy or heat content of a system is called the enthalpy. It is denoted by the symbol **H**. The enthalpy change of the system is indicated as ΔH (Δ =Change of) and refers to heat evolved or absorbed.

Heat energy is measured in kilojoules (KJ). The heat change in a chemical reaction (ΔH) is expressed as a negative value if heat is evolved.

$\Delta H = -92 \text{ KJ}$ means the heat change for the reaction is exothermic.

$\Delta H = +113 \text{ KJ}$ means the heat change for that reaction is endothermic.

Heat of reaction

The heat of reaction for a chemical reaction is the amount of heat evolved or absorbed when molar quantities of substances react together completely.

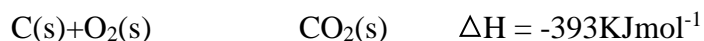
Types of heats of reactions

1. Heat of combustion

This is the amount of heat liberated when one mole of a substance is completely burnt in oxygen.

Example;

1. Carbon undergoes complete combustion according to the equation.



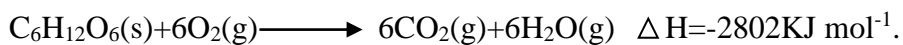
Calculate the amount of heat evolved when 20g of carbon is completely burnt.

12g of carbon liberate 393KJ.

1g of carbon liberates $\frac{393}{12}$ KJ

20g of carbon liberate $\left(\frac{393}{12} \times 20\right) \text{ KJ}$
 $= 655 \text{ KJ}.$

2. Glucose burns in air according to the equation.



Calculate the amount of heat produced when 45g of glucose is completely burnt in air.

Molar mass of $\text{C}_6\text{H}_{12}\text{O}_6 = 12 \times 6 + 1 \times 12 + 16 \times 6$

$$= 180 \text{ g}$$

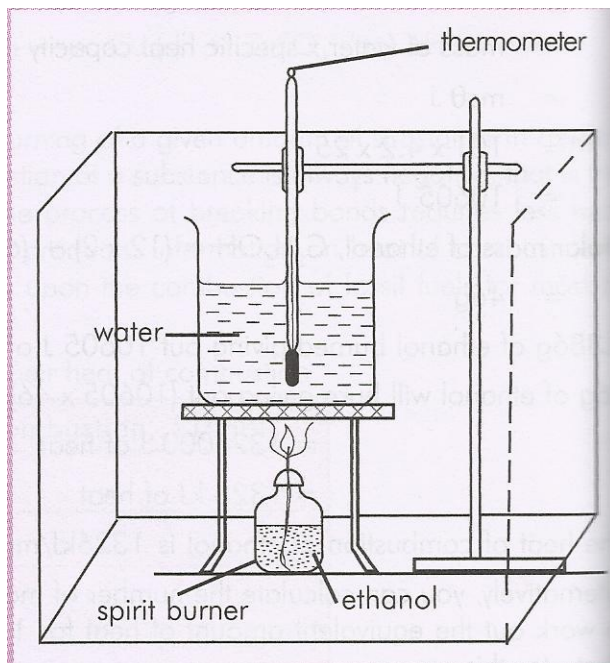
180g of $\text{C}_6\text{H}_{12}\text{O}_6$ liberate 2802KJ

1g of $\text{C}_6\text{H}_{12}\text{O}_6$ liberates $\frac{2802}{180} \text{ KJ}$

45g of $\text{C}_6\text{H}_{12}\text{O}_6$ liberate $\left(\frac{2802}{180} \times 45\right) \text{ KJ}$

$$= 700.5 \text{ KJ}$$

Determination of heat of combustion of ethanol



- A known mass of water (150cm^3) is put in a metal can. Initial temperature of water is noted, $t_1^\circ\text{C}$.

- The spirit lamp is filled with ethanol and weighed. The mass of the lamp and ethanol before burning is noted, $m_1 \text{ g}$.
- The lamp is lit and placed under the metal can. To shield it from draughts; the whole apparatus is placed in a large tin to prevent heat loss.
- The metal tin is heated until the temperature of the water. The final temperature of the water is noted, $t_2^\circ\text{C}$.
- The lamp is put out immediately and reweighed to find the new mass of lamp and ethanol after burning, $m_2 \text{ g}$.

Treatment of results:

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

Mass of water = 150 g

Temperature rise = $(t_2 - t_1)$

Specific heat capacity of water = $4.2 \text{ J K}^{-1} \text{ g}^{-1}$

Heat gained by water = $\text{mass} \times \text{specific heat capacity} \times \text{temperature rise}$

$$= mc\theta$$

$$= 150 \times 4.2 \times (t_2 - t_1) \text{ J}$$

Molar mass of ethanol, $\text{C}_2\text{H}_5\text{OH} = 12 \times 2 + 1 \times 5 + 16 + 1$

$$= 46 \text{ g}$$

$(m_1 - m_2) \text{ g}$ of $\text{C}_2\text{H}_5\text{OH}$ liberate $630(t_2 - t_1)$

1 g of $\text{C}_2\text{H}_5\text{OH}$ liberates $\frac{630(t_2 - t_1)}{(m_1 - m_2)} \text{ J}$

46 g of $\text{C}_2\text{H}_5\text{OH}$ liberates $\frac{630(t_2 - t_1)}{m_1 - m_2} \times 46 \text{ J}$

Note: The heat of combustion in this experiment would be lower than the theoretical value because;

- Heat is lost to the environment during the experiment.
- Some heat produced warms the metal can and air but does not pass into the water.

Examples;

1. When 0.382g of ethanol was burnt, the heat liberated raised the temperature of 100g of water from 16.5°C to 43.5°C. Calculate the heat of combustion of ethanol (SHC of water = $\text{Jg}^{-1}\text{K}^{-1}$, C = 12, H = 1, O = 16).

Heat gained by water = $MC\theta$

$$= 100 \times 4.2 \times (43.5 - 16.5) \text{ J}$$

$$= 11340 \text{ J}$$

$$= \frac{11340}{1000}$$

$$= 11.34 \text{ KJ}$$

Molar mass of $\text{C}_2\text{H}_5\text{OH} = 12 \times 2 + 1 \times 5 + 16 + 1$

$$= 46 \text{ g}$$

0.382g of $\text{C}_2\text{H}_5\text{OH}$ liberate 11.34KJ

1g of $\text{C}_2\text{H}_5\text{OH}$ liberates $\left(\frac{11.34}{0.382}\right) \text{ J}$

46g of $\text{C}_2\text{H}_5\text{OH}$ liberate $\left(\frac{11.34}{0.382} \times 46\right) \text{ KJ}$

$$= 1365.5 \text{ KJ.}$$

2. When 0.16g of methanol was burnt, the heat generated raised the temperature of 200g of water from 21°C to 25°C. Calculate the heat of combustion of methanol (S.h.c of water = $4.2 \text{ Jg}^{-1}\text{K}^{-1}$, C = 12, O = 16, H = 1).

Heat gained by water = $MC\theta$

$$= 200 \times 4.2 \times (25 - 21) \text{ J}$$

$$= 3360\text{J}$$

$$= \frac{3360}{1000}$$

$$= 3.36\text{KJmol}^{-1}$$

Molar mass of methanol, $\text{CH}_3\text{OH} = 12 + 1 \times 3 + 16 + 1$

$$= 32\text{g}$$

0.16g of CH_3OH liberate 3.36KJ.

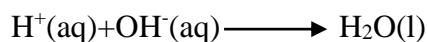
1g of CH_3OH liberates $\frac{3.36}{0.16}$

32g of CH_3OH liberate $\left(\frac{3.36}{0.16} \times 32\right) \text{KJmol}^{-1}$

$$= 672\text{KJmol}^{-1}$$

2.Heat of neutralization

This is the amount of heat evolved when one mole of aqueous hydrogen ions react with one mole of aqueous hydroxide ions to form one of water.



Determination of heat of neutralization between sodium hydroxide and hydrochloric acid.

- A fixed volume of 2M hydrochloric acid(100cm^3) is measured into a plastic beaker. The initial temperature of the acid is noted.
- A fixed volume of 2Msodium hydroxide(100cm^3) is measured into another plastic beaker and its initial temperature noted.
- Hydrochloric acid is poured into a beaker containing sodium hydroxide, stirring gently with a thermometer. The highest temperature of the mixture is noted.

Treatment of results

Initial temperature of hydrochloric acid = 21.0°C .

Initial temperature of sodium hydroxide= 22.0°C

Average temperature of solutions= 21.5°C

Final temperature of the mixture = 34.5°C

Specific heat capacity of solution = 4.2Jg⁻¹k⁻¹

Heat evolved= MCθ

$$= 200 \times 4.2 \times (34.5 - 21.5) J$$

$$= 10920 J$$

$$= \frac{10920}{1000}$$

$$= 10.92 KJ.$$

Moles of HCl or NaOH= $\left(\frac{2}{1000} \times 100\right) \text{ moles}$

$$= 0.2 \text{ moles.}$$

0.2 moles of HCl liberate 10.92 KJ.

1 mole of HCl liberates $\frac{10.92}{0.2}$

$$= 54.6 KJ mol^{-1}$$

Examples;

1.50 cm³ of 2M hydrochloric acid reacts with 50 cm³ of 2M sodium hydroxide, both at 22°C in a plastic beaker. The mixture was stirred and its maximum temperature was 35°C (specific heat capacity of solution = 4.2 Jg⁻¹k⁻¹, density of solution = 1 gcm⁻³). Calculate the enthalpy of neutralization of sodium hydroxide.

Heat gained by the solution = mcθ

$$= 100 \times 4.2 \times (35 - 22) J$$

$$= 5460\text{J}$$

$$= \frac{5460}{100}$$

$$= 5.46\text{KJ}$$

$$\text{Moles of NaOH} = \left(\frac{50}{1000} \times 2 \right) \text{ moles}$$

$$= 0.1\text{moles}$$

0.1moles of NaOH liberate 5.46KJ

$$1\text{mole of NaOH liberates } \frac{5.46}{0.1} \text{KJ}$$

$$= 54.6\text{KJmol}^{-1}$$

2. When 50.0cm^3 of 1.0M sulphuric acid was added to 50.0cm^3 of 2M sodium hydroxide, the temperature of the resultant mixture rose by 13.6°C (specific heat capacity of solution = $4.2\text{Jg}^{-1}\text{K}^{-1}$, density of solution = 1gcm^{-3}). Calculate the enthalpy of neutralization of sodium hydroxide.

Heat gained by the solution = $mc\theta$

$$= 100 \times 4.2 \times 13.6\text{J}$$

$$= 5712\text{J}$$

$$= \frac{5712}{1000} \text{J}$$

$$= 5.712\text{KJ}$$

$$\text{Moles of NaOH} = \left(\frac{50.0}{1000} \times 2 \right)$$

$$= 0.1\text{moles}$$

0.1moles of NaOH liberate 5.712KJ

$$1\text{mole of NaOH liberates } \frac{5.712}{0.1}$$

$$= 57.12 \text{ KJmol}^{-1}$$

Heat of solution

Enthalpy of solution is the amount of heat evolved or absorbed when one mole of a substance is completely dissolved in water so that further dissolution causes no further temperature changes.

Determination of heat of solution of Ammonium nitrate.

- A known volume of water (100cm³) is placed in a beaker and its temperature noted.
- A known mass of ammonium nitrate (4.0g) is added to the water in the beaker.
- Stir the mixture well with a stirrer and record the lowest temperature of the solution reached.

Typical results

Initial temperature of water = 18.5⁰C

Final temperature of water = 16.0⁰C

Temperature fall = 2.5⁰C

Mass of ammonium nitrate = 4.0g

Mass of water = 100g

Specific capacity of water = 4.2 Jg⁻¹k⁻¹

Heat absorbed = mcθ

$$= 100 \times 4.2 \times 2.5$$

$$= 1050 \text{ J}$$

$$= \frac{1050}{1000} \text{ KJ}$$

$$= 1.05 \text{ KJ}$$

Molar mass of ammonium nitrate, $\text{NH}_4\text{NO}_3 = 14 + 1 \times 4 + 14 + 16 \times 3$
 $= 80\text{g}$

4.0g of NH_4NO_3 liberate 1.05KJ

1g of NH_4NO_3 liberates $\left(\frac{1.05}{4.0}\right) \text{KJ}$

80g of NH_4NO_3 liberate $\left(\frac{1.05}{4.0} \times 80\right) \text{KJmol}^{-1}$
 $= +21.0 \text{KJmol}^{-1}$.

Example;

0.5g of ammonium nitrate was dissolved in 50cm^3 of water. The temperature dropped from 24°C to 18°C . Calculate the enthalpy of solution of ammonium nitrate (specific heat capacity of water = $4.2\text{Jg}^{-1}\text{K}^{-1}$, density of water = 1gcm^{-3} , N = 14, H = 1, O = 16).

Heat absorbed = $mc\theta$

$$= 50 \times 4.2 \times (24 - 18)$$

$$= 1260\text{J}$$

$$= \frac{1260}{1000} \text{KJ}$$

$$= 1.26\text{KJ}$$

Molar mass of $\text{NH}_4\text{NO}_3 = 14 + 1 \times 4 + 14 + 16 \times 3$

$$= 80\text{g}$$

0.5g of NH_4NO_3 liberate 1.26KJ

1g of NH_4NO_3 liberate $\frac{1.26}{0.5} \text{KJ}$

80g of NH_4NO_3 liberate $\left(\frac{1.26}{0.5} \times 80\right) \text{KJ}$

$$= 201.6\text{KJmol}^{-1}$$

Heat of Displacement

This is the amount of heat absorbed or evolved when one mole of a substance is displaced from its compound.

Determination of heat of displacement of copper.

- Measure 50cm³ of 0.2M copper (II) sulphate solution into a plastic bottle fitted with a bung carrying a thermometer.
- Record the initial temperature of the solution.
- Weigh 1g of iron filings, add the filings to the solution and shake well.
- Record the highest temperature attained by the solution.

Results and calculations.

Initial temperature of CuSO₄ = 21.0⁰C

Final temperature = 27.0⁰C

Rise in temperature = 6.0⁰C

Mass of 50cm³ of CuSO₄ solution = 50g

Specific heat capacity of solution = 4.2Jg⁻¹k⁻¹

Heat change = mcθ

$$= 50 \times 4.2 \times 6.0\text{J}$$

$$= 1260\text{J}$$

$$= \frac{1260}{1000}\text{KJ}$$

$$= 1.26\text{KJ}$$

$$\text{Moles of CuSO}_4 = \frac{50}{1000} \times 0.2$$

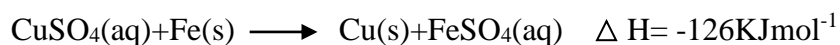
$$= 0.01 \text{ moles.}$$

0.01 moles of CuSO_4 liberate 1.26 KJ

$$1 \text{ mole of } \text{CuSO}_4 \text{ liberates } \frac{1.26}{0.01} \text{ KJ}$$

$$= -126 \text{ KJ mol}^{-1}$$

The thermo chemical equation for the reaction is



Example:

(a). When an excess of iron filings is added to 25 cm^3 of 0.4M copper(II) solution, there was a temperature rise of 15°C . Calculate the molar enthalpy of displacement of copper (specific heat capacity of solution = $4.2 \text{ J g}^{-1} \text{ K}^{-1}$, density of solution = 1 g cm^{-3}).

$$\text{Heat change} = mc\theta$$

$$= 25 \times 4.2 \times 15$$

$$= 1575 \text{ J}$$

$$= \frac{1575}{1000} \text{ KJ}$$

$$= 1.575 \text{ KJ}$$

$$\text{Moles of } \text{CuSO}_4 = \left(\frac{25}{1000} \times 0.4 \right) \text{ moles}$$

$$= 0.01 \text{ moles.}$$

0.01 moles of CuSO_4 liberate 1.575 KJ

$$1 \text{ mole of } \text{CuSO}_4 \text{ liberates } \frac{1.575}{0.01} \text{ KJ}$$

$$= +157.5 \text{ KJ mol}^{-1}$$

(a) Write the thermo chemical equation for the reaction



2. To 100 cm^3 of 0.2M copper(II) sulphate solution in a plastic beaker was added excess zinc dust. There was a temperature change from 20°C to 30.4°C .

(a) Why was a plastic beaker used?

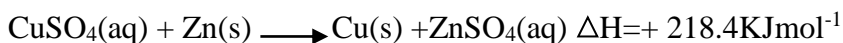
The plastic beaker is a poor conductor of heat and minimizes heat loss to the surrounding.

(b) State what was observed.

The blue copper(II) sulphate solution turns to colourless and a brown solid is deposited.

(b) Calculate the molar heat of the reaction(specific heat capacity of solution= $4.2\text{Jg}^{-1}\text{K}^{-1}$, density of solution= 1gcm^{-3}). Answer(= 218.4KJmol^{-1})

(c) Write a thermo chemical equation for the reaction.



Exercise

1.(a) Describe an experiment that can be carried out to determine the heat of combustion of propanol.

(b) When 0.458g of propanol was burnt, the heat evolved raised the temperature of 100g of water from 18°C to 44°C . Calculate the heat of combustion of propanol.(C=12,H=1, O=16).

(specific heat capacity of solution= $4.2\text{J/ g}^{\circ}\text{C}$, density of solution= 1g/ cm^3).

(c) Name two products of incomplete combustion of propanol.

2. 50cm^3 of 2M sulphuric acid and 50cm^3 of 2M potassium hydroxide, both at 20°C , were mixed in a plastic beaker. The mixture was stirred and it's maximum temperature was 42.0°C (specific heat capacity of solution= $4.2\text{J/ g}^{\circ}\text{C}$, density of solution= 1g/ cm^3).

(i) Write an ionic equation for the reaction.

(ii) Calculate the heat of the reaction.

3.(a) Write an equation to show how ethanol can be prepared from glucose.

(b) When methane burns in oxygen, heat is produced. Write equation for the combustion of methane in excess oxygen.

- (c) The heat of combustion of methane is -890 kJ mol^{-1} . Calculate the volume of methane gas at s.t.p that when burned in excess oxygen would raise the temperature of 178g of water by 10°C (specific heat capacity of water = $4.2 \text{ J g}^{-1} \text{ }^{\circ}\text{C}^{-1}$).

REACTION RATES AND REVERSIBLE REACTIONS.

During a chemical reaction, there are three main events that usually occur.

- (i) The reactants (substances at the start) get used up and new substances (products) form in their place.
- (ii) Heat energy is either evolved (exothermic reaction) or taken in (endothermic reaction).
- (iii) For a fixed amount of reactant, a fixed amount of product is eventually formed.

Chemical reactions vary from very fast to very slow. Burning of fuel (petrol) in a car engine takes place at an explosive rate while rusting of iron takes several weeks.

Reaction Rates

The rate of a chemical reaction refers to the amount of products produced or the amount of reactants used up in a certain time. It is the measure of how fast the reaction takes place.

$$\text{Rate of reaction} = \frac{\text{change in amount of reactant to product}}{\text{change in time}}$$

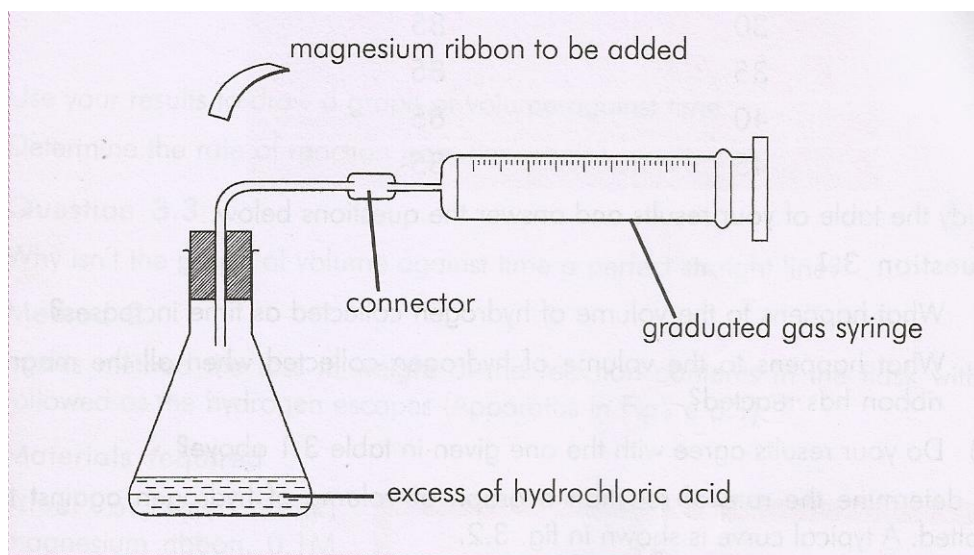
$$\text{Units} = \text{Moles per litre / s}$$

For molecules of gases to react, they must possess a minimum amount of energy when colliding (collision theory). This minimum amount of energy that must be possessed by colliding molecules for a reaction to occur is called **activation energy**.

The smaller the activation energy, the faster the rate of reaction and the greater the activation, the slower the rate of reaction.

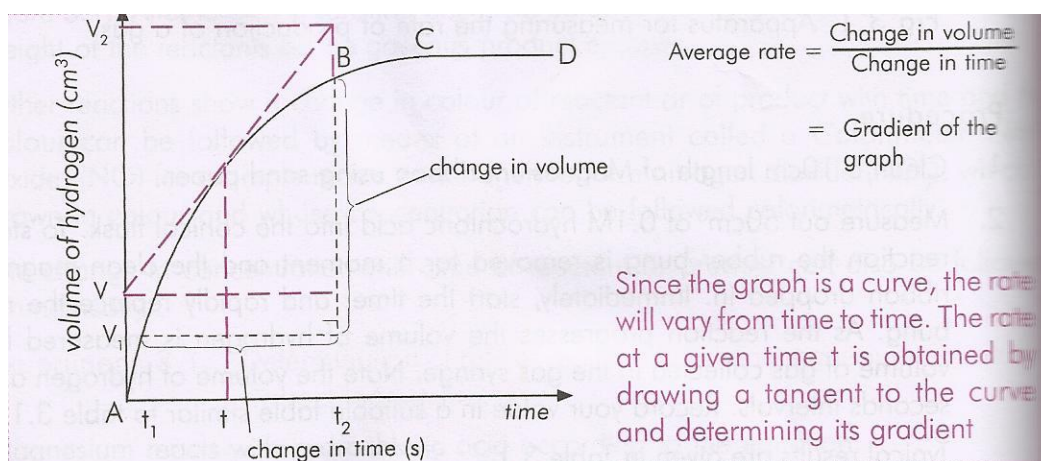
Determination of Rate of reaction

Determination of rate of reaction between magnesium ribbon and dilute hydrochloric acid.

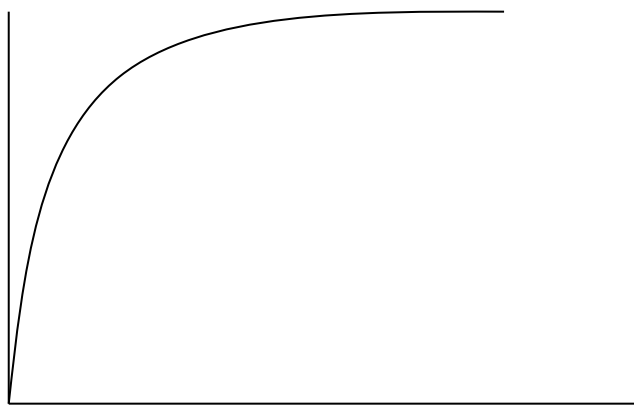


- ❖ Clean a piece of magnesium ribbon using sand paper.
- ❖ Measure a known volume of hydrochloric acid into a conical flask. Remove the rubber bung for a moment and drop in the clean magnesium ribbon. Immediately start the stop clock and rapidly replace the rubber bung.
- ❖ The volume of hydrogen in the syringe is recorded at regular time intervals until the reaction is complete
- ❖ $\text{Mg(s)} + 2\text{HCl(aq)} \longrightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{(g)}$

Rate of reaction is obtained by determining the gradient of the tangent to the curve at a given time interval.



Graph of volume of hydrogen evolved against time

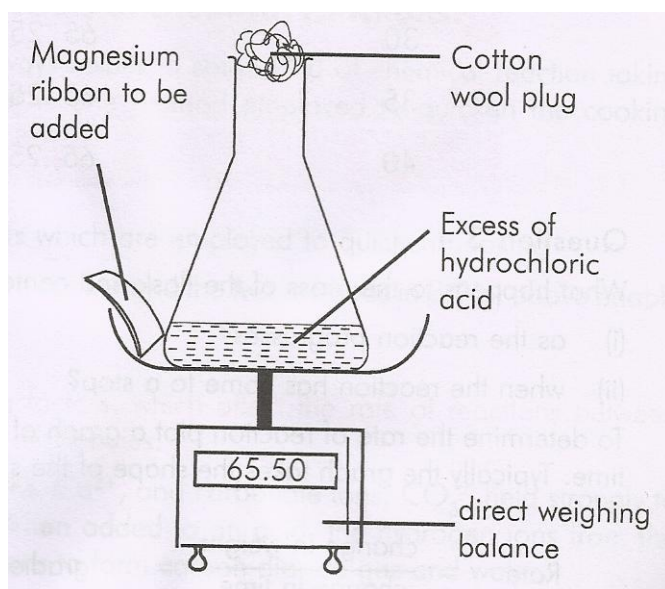


To determine the rate of the reaction at a given time, t , the tangent to the curve is drawn at that time. The gradient of the tangent is the rate of reaction at that time

The units are cm^3/s

Method 2

- ❖ Measure a known volume of hydrochloric acid into the flask and plug the flask with cotton wool to prevent any loss of acid spray.
- ❖ Place the flask containing the acid onto a direct weighing balance. Remove the plug and note the weight.
- ❖ Drop the ribbon into the acid and immediately plug the flask and start the stop clock.
- ❖ Record the weight of the flask and the contents at regular time interval.



To determine the rate of reaction, plot a graph of weight of flask and contents against time

$$\text{Rate} = \frac{\text{Change in weight}}{\text{Change in time}}$$

FACTORS THAT AFFECT THE RATE OF REACTION.

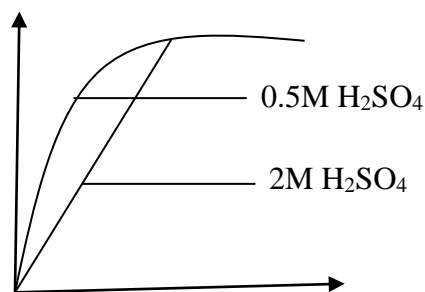
The rate of a chemical reaction is affected by

1. Concentration
2. Temperature
3. Surface area
4. Catalyst
5. Pressure
6. Light.

Concentration.

Increase in concentration of one or more of the reactants increase the rate of the reaction.

Increase in concentration leads to increase in number of reacting particles which in turn leads to increase in the number collisions per unit time. The reaction proceeds faster and takes a shorter time to reach completion.



The effect of concentration can be studied using sodium thiosulphate and dilute hydrochloric acid.

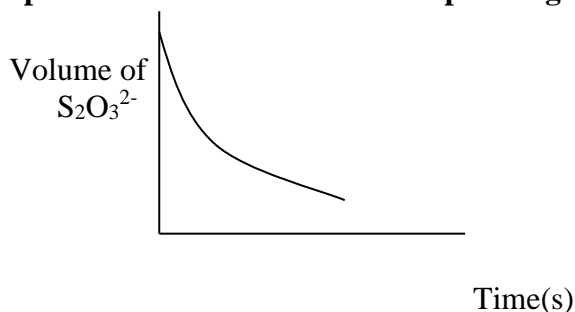
- ❖ Measure a known volume of sodium thiosulphate into a conical flask. Place the conical flask on a piece of white paper marked with a cross.
- ❖ Measure a known volume of hydrochloric acid in another measuring cylinder.
- ❖ Quickly add the acid to the thiosulphate in a conical flask and at the same time start the stop clock, swirl carefully the mixture then leave to stand on the paper.
- ❖ Look down through the mixture and note the time, t , for the cross to disappear.

- ❖ Repeat the procedure using the different concentrations of sodium thiosulphate.

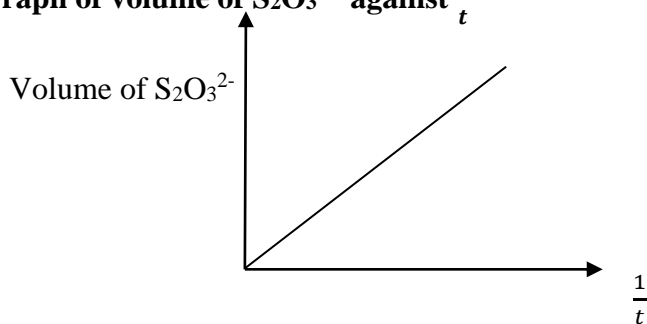
Plot graphs of;

- Concentration of sodium thiosulphate against time.
- Concentration of sodium sulphate against $\frac{1}{t}$

Graph of volume of sodium thiosulphate against time

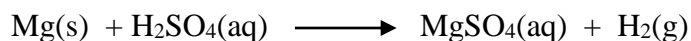


Graph of volume of $\text{S}_2\text{O}_3^{2-}$ against $\frac{1}{t}$



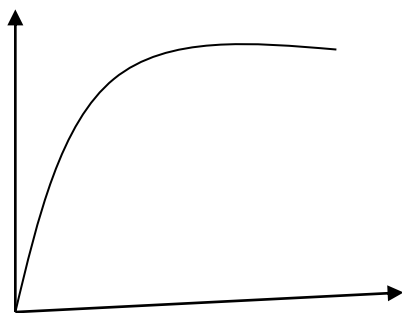
1 Magnesium was reacted with excess dilute sulphuric acid.

- Write an equation for the reaction



- on the same axis, sketch a graph of volume of hydrogen evolved against time, when equal volumes of the following are reacted with the same mass of magnesium.

- 0.5 sulphuric acid
- 2M sulphuric acid



(c) On the same axis, sketch a graph of volume of hydrogen evolved against time when the following are reacted with the same concentration and volume of sulphuric acid.

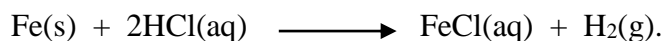
(i). 10g of magnesium

(ii). 50g of magnesium.

2. When a certain volume of 0.1M HCl was reacted at room temperature with excess iron filings. 120cm³ of the gas were produced.

(a) Draw a labelled diagram to show how the rate of the reaction was determined.

(b) Write equation for the reaction that took place



(c) Calculate the

(i). Volume of 0.1M hydrochloric acid required to produce 120cm³ of a gas (1mole of gas occupies 2400cm³ at room temp, Fe = 56)

24000cm³ contain 1 mole

$$120\text{cm}^3 \text{ contain } \frac{1}{24000} \times 120$$

$$= 0.005 \text{ moles of hydrogen}$$

$$\text{Mole ratio HCl : H}_2 = 2:1$$

$$\text{Moles of HCl} = 2 \times 0.005$$

$$= 0.01 \text{ moles.}$$

$$0.1 \text{ moles are contained in } \frac{1000}{0.1} \times 0.01$$

$$= 100\text{cm}^3 \text{HCl}$$

(ii) Mass of iron filings that reacted

$$\text{Mole ratio of Fe : HCl} = 1:2$$

$$\text{Moles of Fe} = \frac{1}{2} \times 0.01$$

$$= 0.005\text{moles}$$

1 mole of Fe contain 56g

$$0.005 \text{ moles of Fe contain } \frac{56}{1} \times 0.005$$

$$= 0.28\text{g of Fe}$$

2 Temperature.

Increase in temperature increases the rate of reaction.

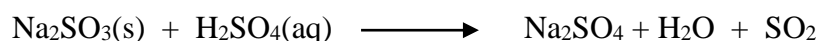
As the temperature increases, the particles of the reacting substance move faster colliding with each other often, increasing the rate of reaction.

Effect of temperature on the rate of reaction.

- Measure 50cm^3 of sodium thiosulphate made by dissolving 20g of thiosulphate in 500cm^3 in a beaker and note its temperature.
- Place the beaker on a piece of paper marked with a cross.
- Add 10cm^3 of hydrochloric acid(2M) to the thiosulphate solution and swirl the mixture. At the same time start the stop clock.
- Look through the mixture and note the time it takes for the cross to disappear.
- Repeat the experiment with sodium sulphate solution heated to different temperatures eg 20°C , 30°C , 40°C and 50°C .
- Each case note the time for the cross to disappear.

1. Sodium sulphite solution was added to dilute sulphuric acid.

(a) Write an equation for the reaction that took place.



(b) On the same axis, Sketch a graph of volume of sulphur dioxide against time when the reaction was carried out at.

(i) 24°C

(ii) 40°C

Surface area

The rate of reaction increases with increase in surface area.

A powder reacts in a much shorter time than the chip i.e. the smaller the solid particles, the faster the rate of reaction. This is because the fine powder has a greater surface area and exposes more particles for the reaction with the liquid.

1 Calcium carbonate lumps were mixed with dilute nitric acid in a conical flask. The mass of the flask and its contents was weighed with time.

(a). Write equation for the reaction.

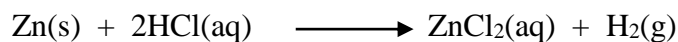


(b)i. Sketch a graph that would be obtained when powdered calcium carbonate is used instead of calcium carbonate lumps. Label the graph B.

(ii) Sketch a graph of mass of flask and its contents against time. Label the graph A

2. A certain mass of Zinc powder was reacted with hydrochloric acid at room temperature.

(a)i Write an equation for the reaction



(ii). Draw a graph to show how the volume of the gaseous product varies with time.

(b) What would be the effect of:

(i) Adding copper (II) sulphate solution to the reaction mixture at room temperature.

❖ The rate of reaction increase.

(ii). Using the same mass of zinc instead of zinc powder.

❖ The rate of reaction decreases.

Give a reason for your answer in b(ii)

❖ The rate of reaction decreases because the surface area available for contact is small.

Catalyst.

A catalyst is a substance which alters the rate of a chemical reaction but remains chemically unchanged at the end of the reaction.

Characteristics of acatalyst

- A catalyst is unchanged chemically at the end of the reaction.
- A small amount of catalyst is able to bring about significant change in a chemical reaction.
- A catalyst does not affect the position of chemical equilibrium of reversible reactions.
- A catalyst does not initiate a reaction, it only accelerates a reaction that is already taking place.
- A catalyst is specific.

The rate of reaction increases when the catalyst is used. A catalyst lowers activation energy so that more molecules can react.

Example.

- 1 In an experiment to determine the rate of reaction between zinc and sulphuric acid, dilute sulphuric acid was reacted with zinc granules to which copper(II) sulphate solution was added and the gas evolved at various times were measured. The results are shown below.

Time(minutes)	0	5	10	15	20	25	30
Volume of gas(cm ³)	0	10	20	25.5	29.5	32	32

(a)What is the role of copper (II)sulphate solution.

To speed up the rate of the reaction between zinc and dilute sulphuric acid.

(b)(i) Plot a graph of volume of hydrogen evolved against time.

(ii) Describe how you would determine the rate of reaction at 12 minutes.

The rate of reaction is determined by determining the gradient(slope) of the graph at a point corresponding to 12 minutes.

2.(a) Name the catalysts used in the following reactions.

(i) the contact process

Vanadium (V) oxide

(ii)the Haber process

Finely divided iron catalyst

(iii)The decomposition of hydrogen peroxide

Manganese(IV)oxide

(a)Potassium chlorate decomposes according to the equation.



(i)Name the catalyst that may be used in this reaction.

Manganese(IV)oxide

(ii)On the same axis, sketch a graph of volume of oxygen evolved against time when;

- the catalyst is not used.
- 0.5g of the catalyst are used.

Pressure

Pressure affects reactions which occur in gas phase only. High pressure brings gas molecules closer so they collide more frequently leading to an increase in the rate of the reaction.

Light

Some reactions are speeded up considerably when exposed to bright sunlight. These molecules absorb energy in form of light rather than heat, such reactions are called photochemical reactions.

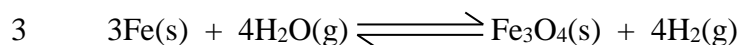
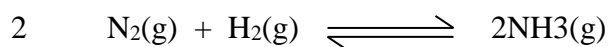
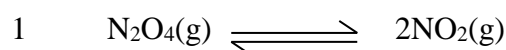
Processes which are affected by light include;

- Photosynthesis, in which plants are unable to manufacture starch when placed in the dark.
- Decomposition of silver bromide into silver and bromine, a basis of photography.
Photographic film consists of tiny silver bromide crystals. When light falls on the film, it causes the decomposition of the bromide and the extent of decomposition depends upon the brightness of the light.

REVERSIBLE REACTIONS

This is a reaction which occurs in both forward and backward directions at the same time.

Examples



Industrial application of reversible reaction

Haber process



This reaction is exothermic and fewer molecules are formed.

Temperature

Since the reaction is exothermic, an optimum temperature of 450°C leads to increase in the rate of formation of ammonia.

Pressure.

Since the reaction occurs with a decrease in volume, high pressure increases rate of formation of ammonia

Concentration

If the ammonia formed is removed the rate of reaction will increase leading to formation of more ammonia.

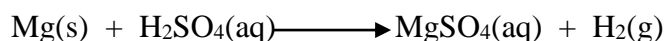
Catalyst.

Finely divided iron catalyst is used in Haber process. It increases the rate of formation of ammonia.

Exercise

1 12.0g of clean magnesium ribbon were added to 50cm³ of 1.0M sulphuric acid. The volume of the gas evolved was measured at fixed time interval until the reaction stopped.

(a) Write equation for the reaction that took place



b(i). Sketch a graph of volume of gas against time. Label the graph x

(ii) On the same axis sketch the graph that would be obtained if 12.0g of Magnesium powder were used instead of magnesium ribbon. Label the graph Y

(c)i Give a reason for the difference in the graphs X and Y

Magnesium powder offers a large surface area over which the reaction occurs than magnesium ribbon and takes a short time to reach completion than magnesium ribbon.

(ii). Name one other factor that can cause similar results as in b(i) above

❖ Increase in temperature

- ❖ Addition of a catalyst of CuSO_4 .
- ❖ Increase in concentration of sulphuric acid.

Exercise

1.(a) What is meant by the rate of reaction?

(b) Explain briefly how rate of reaction is affected by

(i) concentration of reactants.

(ii) temperature

(c) The table below shows the volume of hydrogen evolved at various time intervals when magnesium powder was reacted with 2M hydrochloric acid.

Time(s)	0	2	4	6	8	10	12	14
Volume(cm^3)	0	17.0	23.8	29.0	33.0	36.5	37.5	37.5

(i) Plot a graph of volume of hydrogen versus time.

(ii) Determine the rate of reaction at 3 seconds.

(iii) Determine the volume of hydrogen evolved after a time interval of 5 seconds.

(iv) State how the rate of reaction at 3 seconds would be affected if magnesium powder was replaced by magnesium ribbon, and give a reason for your answer.

2.(a) (i) State what is meant by rate of reaction.

(ii) With the aid of a labelled diagram, describe an experiment that can be carried out to determine the rate of production of carbon dioxide gas from calcium carbonate and hydrochloric acid.

(c) The table below shows the variation in mass of calcium carbonate with time when reacted with dilute hydrochloric acid.

(d)

Mass of CaCO ₃ (g)	84	64	49	27	11	9	8
Time(mins)	0	1	2	4	7	8	9

- (i) Plot a graph to show how the mass of calcium carbonate varies with time.
- (ii) Determine the mass after 5 minutes.

(c) From your graph determine,

- (i) The rate of reaction at 3 minutes.
- (ii) The time taken for half of the calcium carbonate to react.

(d) State at least two ways in which the rate of production of carbon dioxide can be increased.

3. (a) 31.25g of zinc carbonate powder were reacted with 1M hydrochloric acid solution under room temperature.

(i) Write equation for the reaction that took place.

(ii) Sketch a graph to show the variation of volume of the gaseous product with time and label this graph Q.

(iii) On the same axes, sketch a graph to show how the volume of the gaseous product would vary if 2M hydrochloric acid was used instead of the 1M hydrochloric acid. Label this graph R.

(b) Calculate the volume of the gaseous product formed at room temperature (Zn=65; C=12; O=16, 1mole of gas occupies 24dm³ at room temperature).

4. (a) 12.4g of large pieces of zinc carbonate were reacted with 50cm³ of 2M hydrochloric acid at room temperature. The decrease in mass was measured at regular time intervals.

(i) Write an equation for the reaction.

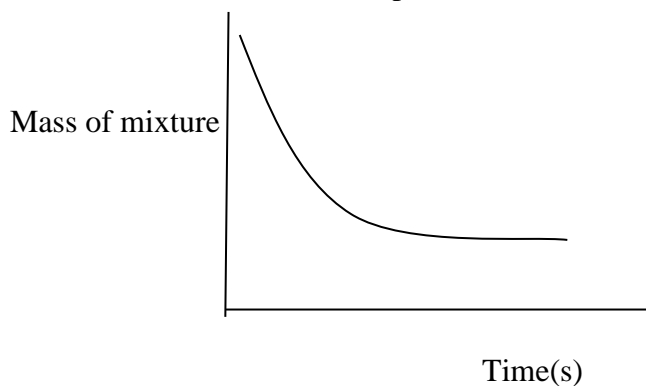
(ii) Sketch a graph to show variations of decrease in mass with time.

(b) State what would be observed if the same mass of zinc powder was used instead of the large pieces. Give a reason for your answer.

(c) State what would be observed if the same mass of large pieces of zinc carbonate was used at 40°C. Give a reason for your answer.

5. (a) State the factors that can affect the rate of a chemical reaction.

(b) A mixture of a known mass of magnesium and a certain volume of 2M hydrochloric acid was put in a conical flask and the mass of the mixture was recorded at various time intervals. The result of the experiment is shown in the graph below.



On the same axes, draw a graph that would be obtained when the same mass of magnesium was reacted with the same volume of 1M hydrochloric acid.

(c) 5.0g of calcium carbonate was reacted with 20.0cm³ of 2M hydrochloric acid.

- Write equation for the reaction between hydrochloric acid and calcium carbonate.
- Calculate the mass of calcium carbonate that was left (C=12; Ca=40; O=16).

6. (a) (i) What is rate of reaction?

(ii) How does particle size affect the rate of reaction? Explain your answer.

(b) The table below shows the time taken for sulphur to form when various concentrations of sodium thiosulphate were used

Concentration of S ₂ O ₃ ²⁻ (M)	0.2	0.6	0.8	1.2	1.6
Time for sulphur to form(s)	60	20	15	10	7.5
$\frac{1}{t}$ (sec ⁻¹)	0.017	0.05	0.07	0.10	0.13

Plot the graph of $\frac{1}{t}$ (sec^{-1}) vertical against concentration of thiosulphate.

(c) (i) Explain the relationship between rate of reaction and $\frac{1}{t}$

(ii) Deduce from the graph, how the rate of reaction varies with concentration of thiosulphate.

(e) Name one reagent that can be used to test for sulphur dioxide and state what would be observed if the reagent named was used.

7. In an experiment to determine the rate of reaction between zinc and sulphuric acid, dilute sulphuric acid was reacted with zinc granules to which some copper (II) sulphate solution was added. The volumes of hydrogen gas evolved at various times were measured. The results are shown in the table below.

8.

Time(minutes)	0	5	10	15	20	25	30
Volume of gas(cm^3)	0	10	20	25.5	29.5	32	32

(a) (i) What is the role of copper (II) sulphate solution?

(ii) Write an ionic equation for the reaction above.

(iii) Explain what would happen to the reaction if zinc granules were replaced with zinc powder.

(b) (i) Plot a graph of volume of hydrogen evolved against time.

(ii) Describe how you would determine the rate of reaction at 12minutes.

(iv) Compare the rate of reaction at 12 minutes with that at 20 minutes. Give a reason for your answer.

8. (a) A certain mass of zinc powder was reacted with hydrochloric acid at room temperature.

(i) Write an equation for the reaction.

- (ii) Draw a graph to show how the volume of the gaseous product varied with time.
- (b) What would be the effect of
- (i) Adding copper (II) sulphate solution to the reaction mixture at room temperature.
- (ii) Using the same mass of zinc granules instead of zinc powder.
9. 8g of zinc powder was added to 50cm³ of 1M hydrochloric acid in a conical flask.
- (a) Write equation for the reaction which took place.
- (b) (i) Describe how the rate of reaction can be determined. Draw a diagram to illustrate your answer.
- (ii) Sketch a graph to show the rate of reaction. Label this graph X.
- (c) in another experiment, 8g of zinc powder was added to 50cm³ of 0.5M hydrochloric acid.
- (i) Sketch a graph for the rate of the reaction using the same axes in b(ii).Label this graph Y.
- (ii) Explain the shapes of the two graphs.

END

