CHEMISTRY

10 - 12

What is chemistry?

Chemistry is the study of matter and the properties of the forms of matter.

Branches of chemistry

1. Organic chemistry

It is the study of compounds of carbon and its components

2. Biochemistry

It is the study of chemical reactions or processses that occur in living things eg respiration and photosynthesis

3. Geochemistry

It is the study of chemical composition and chemical reactions associated with the earth and other planents

4. Inorganic chemistry

It is the study of compounds which do not contain carbon. It is the branch of chemistry relating to inorganic compounds.

The particulate nature of matter

Matter

Matter is anything that occupies space and has mass.

Basic units of matter

There are three basic units of matter. These are:

- Atoms
- Ions
- Molecules

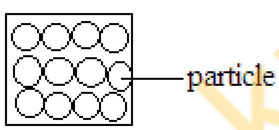
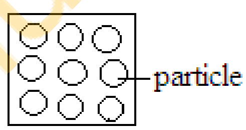
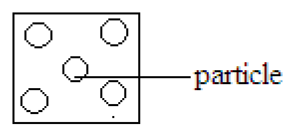
State of matter

State of matter is the form in which matter exists. Matter exists in three forms. These are:

- Solids
- Liquids
- Gases

Examples of solids	Examples of liquids	Examples of gases
Stone	Water	Oxygen
Glass block	Cooking oil	Hydrogen
Wooden block	Paraffin	Carbon dioxide
Copper block	Petrol	Carbon monoxide

Characteristic properties of the three states of matter

	Solids	Liquids	Gases
Shape	Have fixed shape	Have no fixed shape. They take the shape of the container in which they are placed.	Have no fixed shape
Volume	Have fixed volume	Have fixed volume	Have no fixed volume. Particles spread to fill the space available.
Compressibility	Can not be compressed	Can not be compressed	Can be compressed.
Arrangement of particles	Particles are closely packed and arranged in a regular pattern. The particles are held together by strong electrostatic forces of attraction called cohesive forces. 	Particles are slightly further apart than in solids. Paticles are held together by weak electrostatic forces of attraction. 	Particles are much further apart from each other. The forces which hold the particles together are negligible. 
Movement of particles	Particles move by vibrating at fixed positions	Particles move by vibrating rapidly over short distances. Particles move from one position to the other.	Particles move at random at a very high speed.

Solids and liquids can not be compressed because their particles are close together. However, gases can be compressed because the gas particles are far apart from each other and can be forced to move closer by exerting pressure.

Changes in state

Changes in state are physical changes that occur when the particles of a substance absorb or lose energy.

A substance can change from one state to another when it is either heated or cooled.

[A] Heating

Heating involves the addition or supply of heat to a substance.

As a substance is heated, it absorbs energy and changes from a solid to a liquid and finaly to a gas. The kinetic energy possessed by its particles increases and they move vigorously.

Effects of heating substances

1. Melting

Melting is change of state from solid to liquid.

For example, ice changes to water when heated

$$\underset{[\text{ice}]}{\text{Solid}} \xrightarrow{\text{Melting}} \underset{[\text{water}]}{\text{Liquid}}$$

Melting takes place when the particles of a solid absorb energy to overcome the forces holding them in fixed positions and move. They rearrange themselves to form a liquid.

The temperature at which a substance changes from solid to liquid is called melting point.

Uses of melting

(a) It is used in welding

(b) It is used in cutting and shaping of metals in the industry

2. Evaporation

Evaporation is the change of state from liquid to gas.

For example, water changes to steam (water vapour) when heated.

$$\underset{[\text{water}]}{\text{Liquid}} \xrightarrow{\text{Evaporation}} \underset{[\text{steam}]}{\text{Gas}}$$

Uses of evaporation

(a) It is used in drying clothes

(b) It is used in obtaining crystals from solutions

Factors that affect rate of evaporation

- Surface area
- Wind current
- Humidity
- Temperature

Note

Evaporation and boiling are both physical processes that change a liquid into a gas.

The liquid absorbs heat energy during these physical changes in state.

Differences between evaporation and boiling

Evaporation	Boiling
Occurs at any temperature below boiling	Occurs at boiling point
Occurs only at the surface of the liquid	Occurs throughout the liquid
No bubbles are observed	Bubbles are observed
Occurs slowly	Occurs rapidly

3. Sublimation

Sublimation is the direct change of state from solid to gas by heating or gas to solid by cooling without passing through the liquid state.

$$\text{Solid} \xrightarrow{\text{Sublimation by heating}} \text{Gas}$$

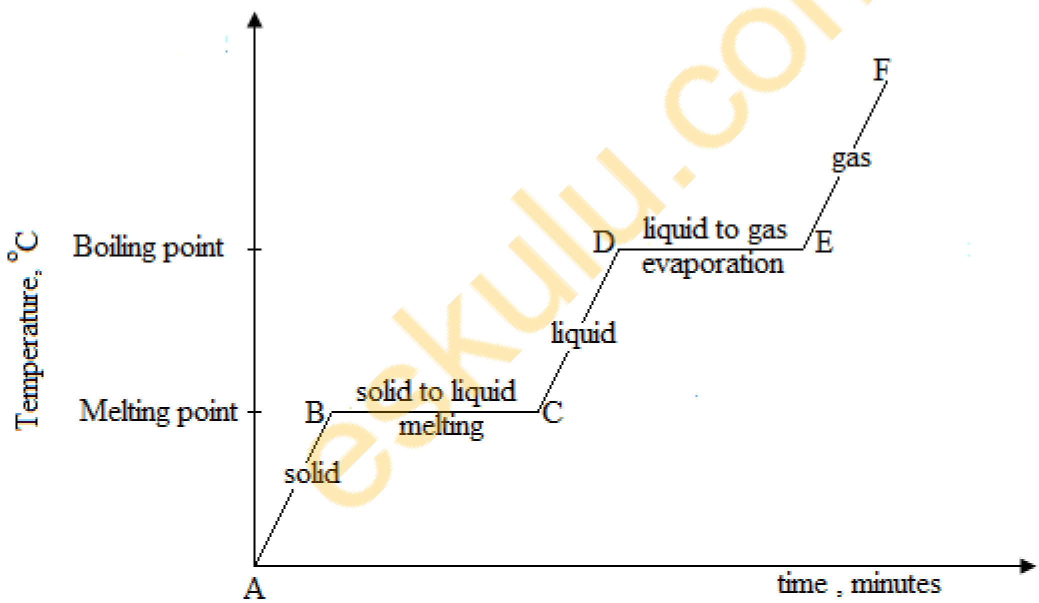
$$\text{Gas} \xrightarrow{\text{Sublimation by cooling}} \text{Solid}$$

Examples of substances that can sublime

- Iodine
- Ammonium sulphate
- Ammonium chloride
- Carbon dioxide (upon cooling to form ice)

The heating curve

The heating curve is a graph showing changes in temperature with time for a substance being heated



Section AB: The substance remains in solid state. The heat energy provided is absorbed by the solid particles and they vibrate harder about their fixed positions.

Section BC: A mixture of solid and liquid will be observed at this temperature which is called the melting point of the substance. At point C, the solid has turn completely into a liquid.

Section CD: Thesubstance remains in liquid state. The liquid particles continue to absorb heat energy and their kinetic energy increases, causing the temperature of the liquid to rise.

Section DE: A mixture of liquid and gas will be observed at this temperature which is called boiling point of the substance. At point E, the liquid has turn completely into a gas.

Section EF: The gas particles will absorb energy and move further apart as they become more energetic. The temperature of the gas will rise.

Summary

Slope sections of the heating curve: As a substance is heated, it absorbs heat energy and its temperature rises, then it changes from solid to liquid and finally to gas.

Flat sections of the heating curve: The flat section shows the melting point and boiling point. Here the temperature remains constant over a period of time as energy being absorbed is used to change the state of a substance

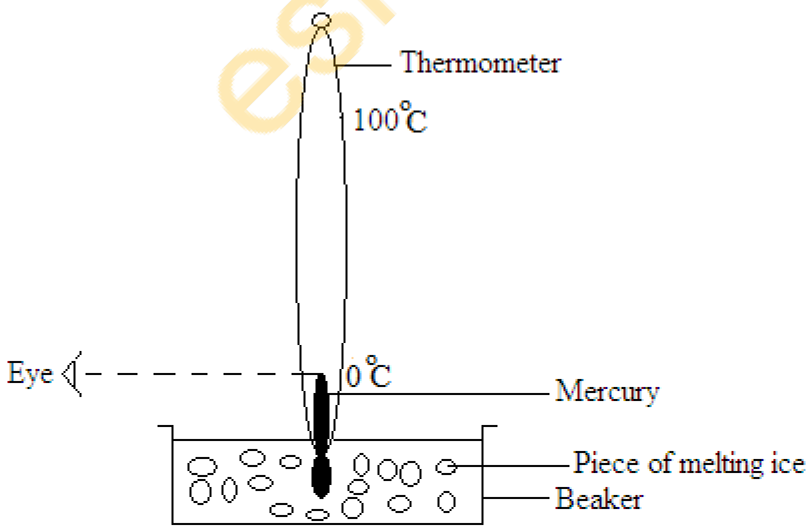
Note

A pure substance has a fixed temperature. It has an exact boiling point and melting point.

Impurities raise the boiling point and lower the melting point.

Example

1. The diagram below shows an experiment on changes of state

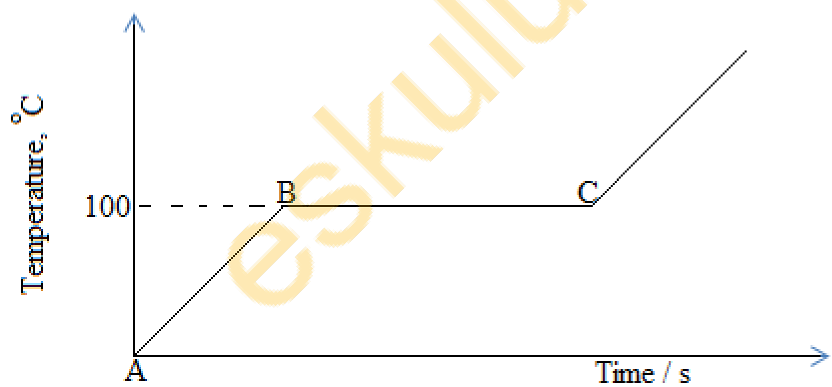


(a) What is the reading on the thermometer?
(b) What must be done to the melting ice for it to completely change to the next state of matter?
(c) Describe the change of state that the ice will undergo in question (b) above
(d) What term is used to describe the temperature at ice changes its state?
(e) Some substances can change from solid state into gaseous state without becoming a liquid. What term is used to describe such a reaction?

Solution

(a) 0^oC (Temperature for melting ice)
(b) It must be heated/warmed/put in sunlight
(c) Melting
(d) Melting point
(e) Sublimation

2. Consider the graph below



(a) What are the two possible states of matter at A and B?
(b) Name the change of state at A and B

Solution

(a) At A: Solid and liquid

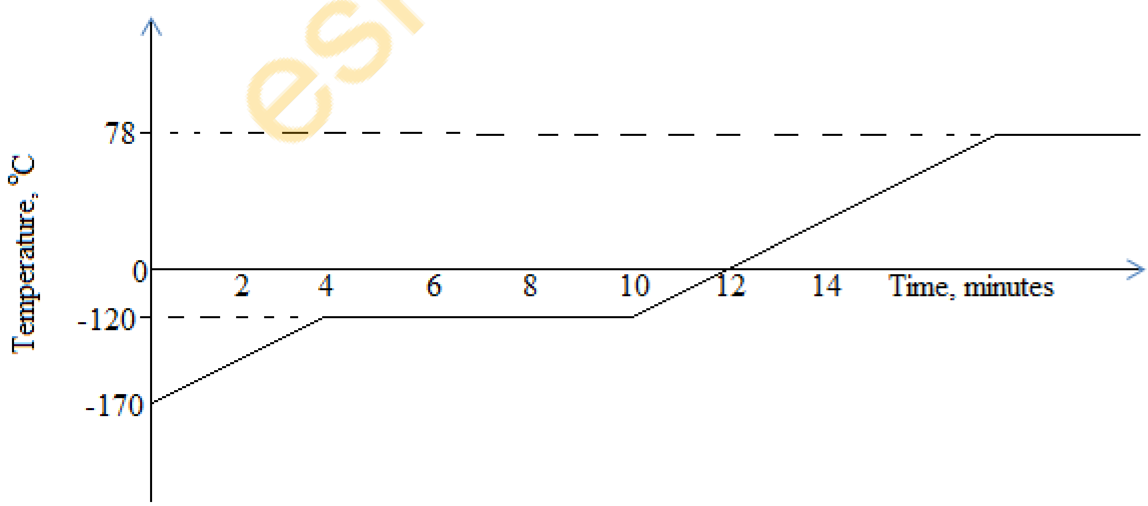
At B: Liquid and gas

(b) At A: Melting

At B: Evaporation

Exercise

1. The graph below shows the temperature of a sample of ethanol varied with time.



(a) What is the melting point of ethanol?

(b) What is the boiling point of ethanol?
(c) For how long did the sample of ethanol melt?
(d) How are you able to tell from the graph that the sample of ethanol is pure?

[B] Cooling

Cooling involves the removal of heat from a substance

Effects of cooling a substance

1. Condensation

Condensation is the change of state from gas to liquid

For example, steam changes to water when cooled

$$\underset{\text{[steam]}}{\text{Gas}} \xrightarrow{\text{Condensation}} \underset{\text{[water]}}{\text{Liquid}}$$

2. Freezing

Alternative term: Solidification

Freezing is the change of state from liquid to solid

For example, water changes to ice when cooled

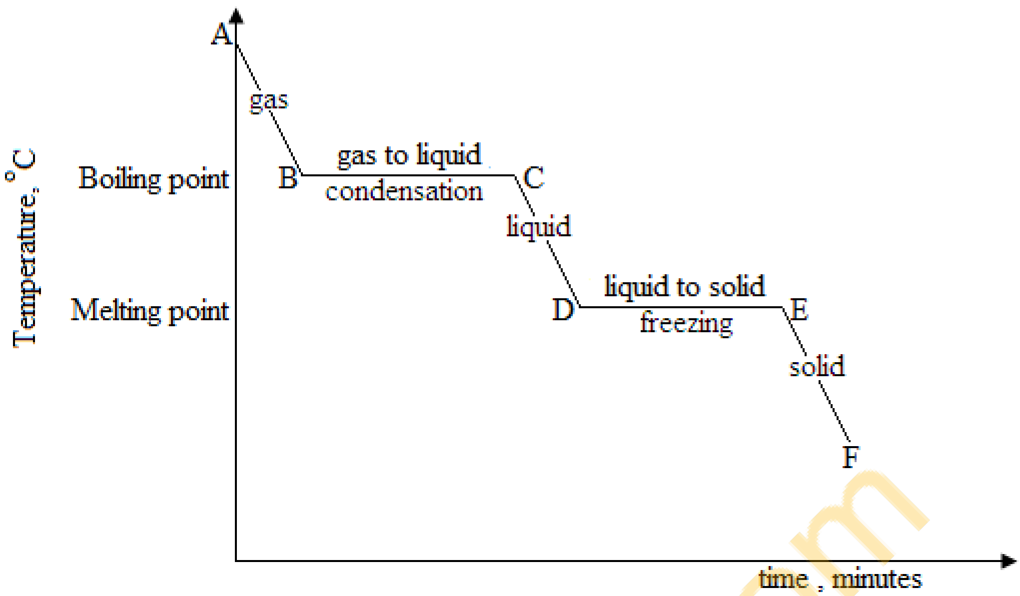
$$\underset{\text{[water]}}{\text{Liquid}} \xrightarrow{\text{Freezing}} \underset{\text{[ice]}}{\text{Solid}}$$

The temperature at which a liquid changes into a solid is called freezing point.

Freezing point is also called the melting point for a pure substance.

The cooling curve

The cooling curve is a graph showing changes in temperature with time for a substance being cooled.



Section CD: The substance remains a liquid. As cooling takes place, the particles lose kinetic energy and moves more slowly.

Section DE: A mixture of a liquid and solid is observed as the liquid particles rearrange themselves to form a solid structure.

Section EF: The temperature of the solid drops as cooling continues.

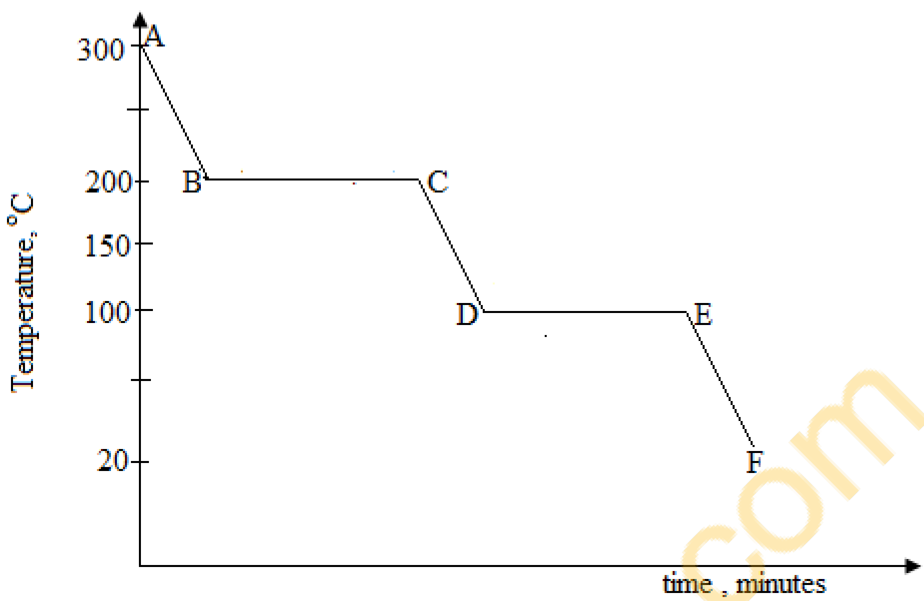
Summary

Slope sections of the cooling curve: As a substance is cooled, it loses heat energy and its temperature falls, then it changes from gas to liquid and finally to solid.

Flat sections of the cooling curve: The flat section shows the melting point and boiling point. Here the temperature remains constant over a period of time as energy being lost is used to change the state of a substance.

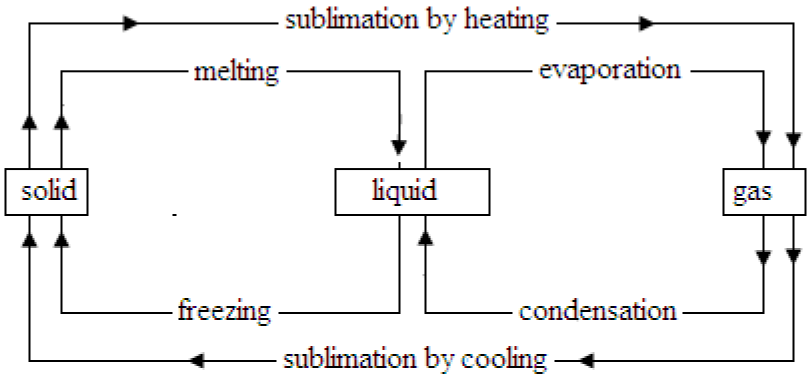
Exercise

1. The graph below shows a cooling curve of a substance as its temperature falls from $300^{o}C$ to $20^{o}C$.



(a) At $250^{o}C$, is the substance a solid, liquid or gas?
(b) What was the boiling point of the substance?
(c) What was the melting point of the substance?
(d) Why does the temperature stay constant over the section BC and DE despite the fact that the substance is losing energy to the surrounding?

Summary of the changes in state



Example

1. Choose from the following terms to answer this question

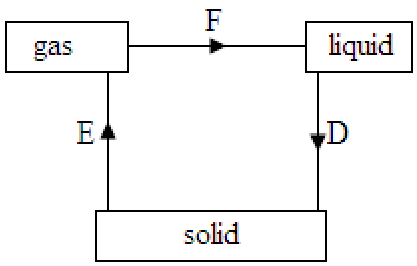
   Sublimation, diffusion, matter, liquefaction, ion
   Which of the terms above describes:
   (a) The physical material of the universe
   (b) A basic unit of matter
   (c) Carbon dioxide gas changing to dry ice
   (d) The spreading movement of particles
   (e) A gas changing to a liquid

Solution

(a) Matter
(b) Ion
(c) Sublimation
(d) Diffusion
(e) Liquefaction

Exercise

1. Study the diagram below and answer the questions that follow

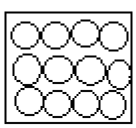


(a) Name the processes D, E and F
(b) Describe what you would do to convert a liquid into a solid
(c) State one of the basic units of matter

2. Matter is classified as solid, liquid or gas. State two physical properties of each of the following:

(a) Solid
(b) Liquid
(c) Gas

3. The diagram represents the arrangement of the particles ia a solid



(a) Descibe the movement of particles in the solid
(b) How does this movement alter as the temperature of the solid is increased
(c) How does this movement alter as the solid melts
(d) Samples of the gases carbon dioxide, $M_r = 44$, and hydrogen, $M_r = 2$, are at same temperature.

(I) Compare the speeds of the molecules in these two gases

(II) Equal masses of steam and water contain the same number of molecules. Explain why the volume of the steam is much greater than that of water

Physical changes and chemical changes

Physical changes

A physical change is change in which no new substance is formed.

Chemical changes

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

Differences between physical and chemical changes

Physical change	Chemical change
No new substance is formed e.g. melting a solid	A new substance is formed e.g. burning substances
Usually the change is easily reversible e.g. boiling a liquid	Usually the change is not easily reversible e.g. precipitation of a solid by the reactions of two solutions
Usually no energy is given out or taken in e.g. heating a wire by electricity	Usually energy is given out or taken in e.g. decomposition of substances by electricity
The mass of the substance remains the same e.g. magnetizing iron	The mass of the new substance is different from that of the starting substance e.g. rusting of iron

The kinetic theory of matter

The theory states that: Matter is made up of tiny particles which are in a continuous random motion.

Experimental evidence of the kinetic theory of matter

1. Brownian motion

Brownian motion is the term used to for the continuous random of particles, particularly of gases and liquids

This phenomenon was first observed by Robert Brown in 1827 who, while studying pollen grains under water, he observed that the pollen grains were moving about in a random way. This same phenomenon can be observed by studying smoke particles in air.

Experiment

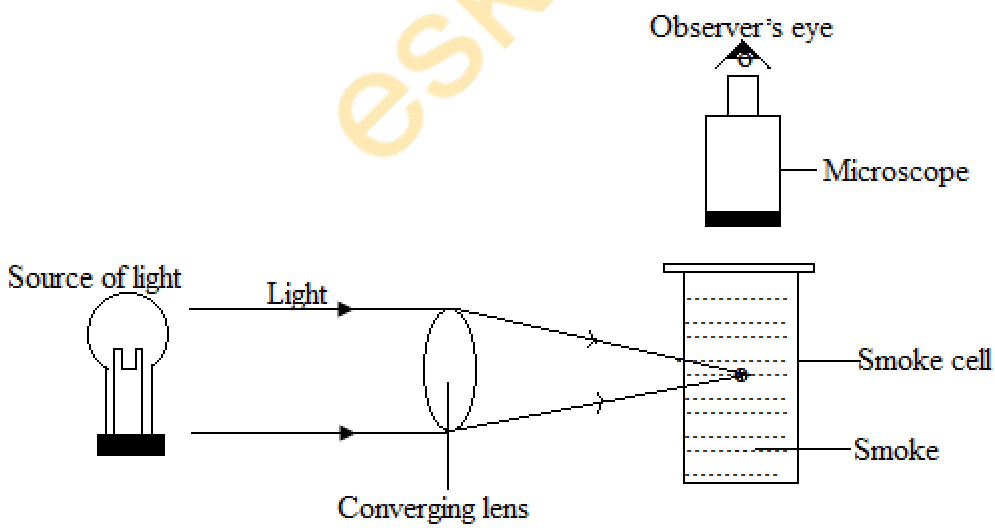
Aim: To study the random motion of smoke particles or to observe Brownian motion in a smoke cell.

Apparatus

- Glass cell
- Source of light
- Microscope
- Converging lens
- Source of smoke

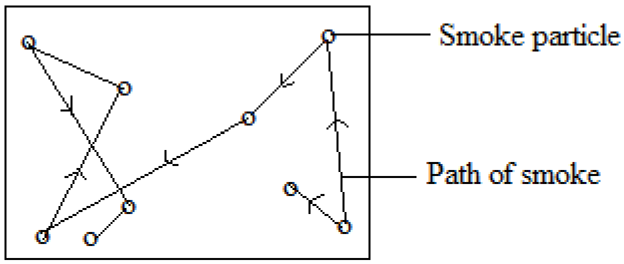
Method

A small glass cell (smoke cell) in which smoke has been trapped is viewed through a microscope. A microscope is used because the smoke particles are too tiny to be seen using the naked eye. A converging lens is used to focus light from the lamp into the smoke cell. The experimental arrangement is shown below



Observation

When light strikes the smoke particles, they appear as bright points of light under the microscope moving randomly in a zig – zag path. The smoke particles appear as spots of light because they reflect some of the light from the source of light towards the microscope.



Explanation

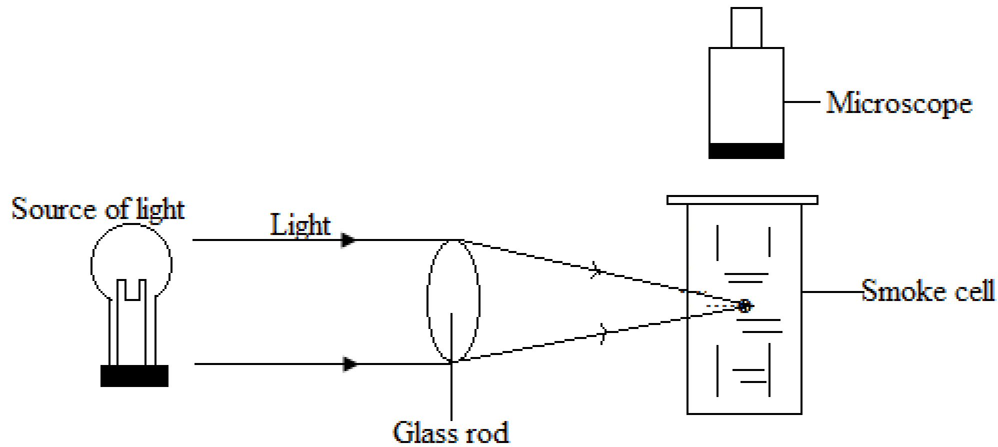
The zig – zag movement is due to the collision of the smoke particles with invisible air molecules that move about randomly in the smoke cell. This is called Brownian motion.

Conclusion

The air molecules are in a continuous random motion colliding with the smoke particles and the walls of the smoke cell

Exercise

1. The figure below shows one of the forms of an apparatus used to observe Brownian motion of smoke particles in air. Mr Naosa D. K looking through the microscope sees tiny bright specks which he describes as 'dancing about'.

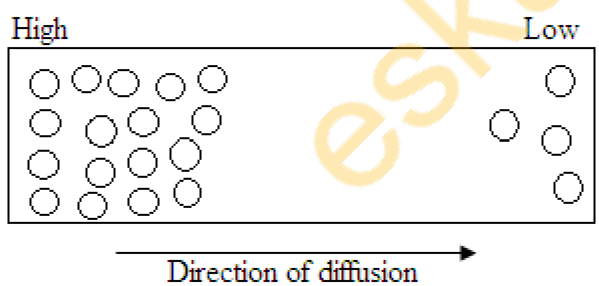


(a) What are bright specks?
(b) Why are the specks 'dancing about'?
(c) State the conclusion that can be drawn from the Brownian motion?

Diffusion

Definition: Diffusion is the movement of particles from the region of high concentration to the region of low concentration.

Definition is very strong evidence for the mobility of particles



Rate of diffusion

Rate of diffusion is the amount of gas or liquid diffusion in a unit of time.

Factors that affect the rate of diffusion

Temperature: Rate of diffusion is faster is if the temperature is high and slower if the temperature is low.

Concentration: Rate of diffusion is faster if there is a large difference in the concentration of particles between two points.

Size of particles: Rate of diffusion is faster if size of particles is small and slower if the size of particles is large.

Note: The states of matter in which diffusion takes place are liquids and gases.

Diffusion in liquids

Experiment

Aim: To demonstrate diffusion in liquids

Apparatus

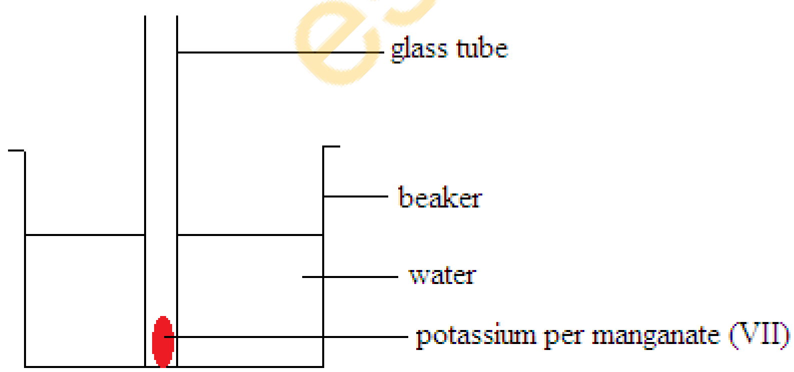
Potassium permanganate (VII)

Water

Glass tube

Method

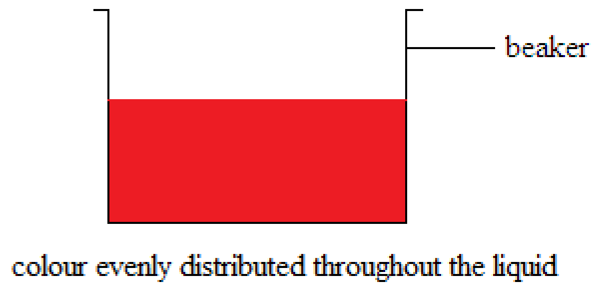
Using a glass tube, place a crystal of potassium per manganate (VII) in the middle of the beaker filled with water. The glass tube is used to ensure that the mixing with water does not occur before the crystal is in place.



Leave the beaker undisturbed and observe carefully.

Observation

The crystal of potassium permanganate (VII) dissolves and the purple colour slowly spread to the bottom of the beaker and eventually the colour distributes itself throughout the liquid.



Conclusion

Diffusion has taken place and water turns purple because the particles of potassium permanganate (VII) have diffused to all parts of the water.

Diffusion in gases

Diffusion in gases is faster than in liquids. The gas molecules move randomly at a very high speed.

Experiment

Aim: To demonstrate diffusion of oxygen gas and nitrogen dioxide gas

Apparatus

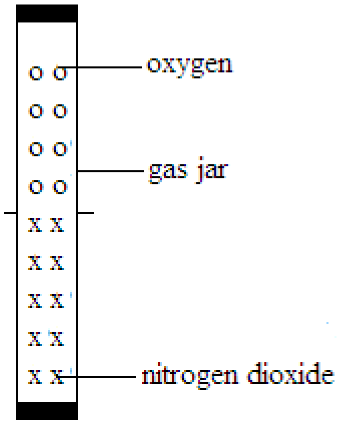
Oxygen gas

Nitrogen dioxide gas

Two gas jars

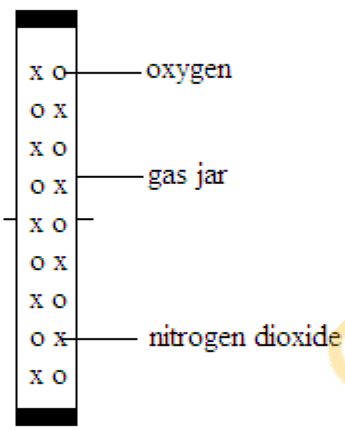
Method

Invert a gas jar containing oxygen and place it on top of a gas jar containing nitrogen dioxide gas.



Observation

The brown fumes of nitrogen dioxide gas diffuses slowly into a gas jar containing oxygen gas and spread out and mix evenly.



Conclusion

Diffusion has taken place since the nitrogen dioxide molecules have moved randomly and mix evenly with the oxygen molecules.

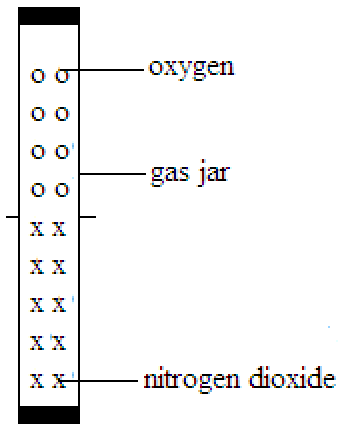
Other examples of diffusion

When a stopper is taken out of the perfume bottle, the smell is noticed because the particles move from a region of high concentration to region of low concentration.

Exercise

1. Matter is made up of tiny particles as it can be evidenced from the process by diffusion

   (a) Explain what is meant by diffusion
   (b) In which state(s) of matter does diffusion occur?
   (c) A gas jar of oxygen gas was inverted and placed on top of a gas jar containing nitrogen dioxide as shown below.

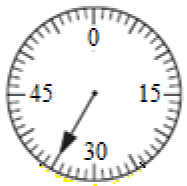


Draw a similar diagram to show the arrangement of the molecules of the two gases after being in contact for 30 minutes.

   (d) When the stopper is taken out of a bottle of perfume, the smell can soon be noticed. Explain from your knowledge of particles why this happens.

Various apparatus used in chemistry

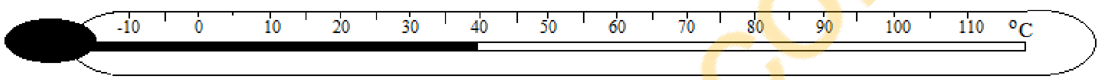
1. Stop watch



Use: It is used for measuring time.

SI unit for time: Second, s.

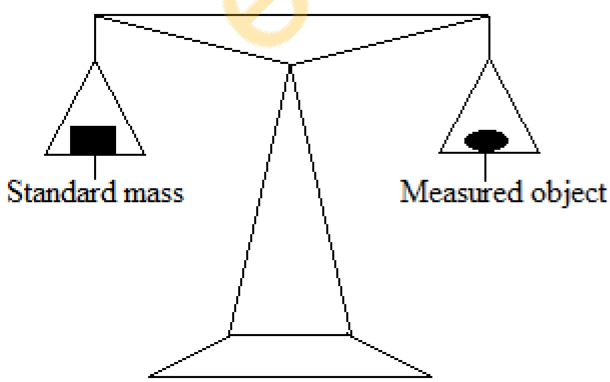
2. Laboratory thermometer



Use: It is used for measuring temperature.

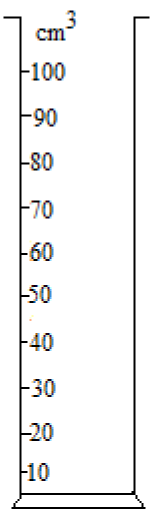
SI unit for temperature: Kelvin, K.

3. Beam balance



Use: It is used to measure mass
SI unit for mass: Kilogram, Kg

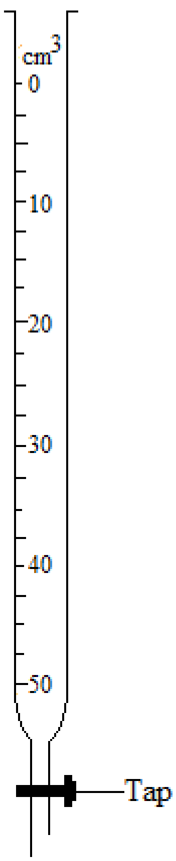
4. Measuring cylinder



Use: It is used for making approximate measurements of volumes of liquids

Measuring cylinders are in different sizes

5. Burette



Use: It is used for accurate measurements of different volumes of liquids

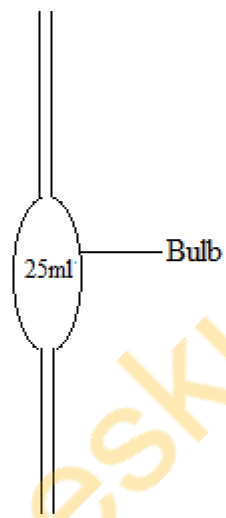
It can measure small amounts of volume. For example, a burette can be used to measure **exactly**22.5cm^{3}of the liquid. It has an accuracy of 0.1cm^3.

It has a scale which starts from 0cm^{3}at the top up to 50cm^{3}at the bottom

The scale is more sensitive than the measuring cylinder

A burette has a long narrow shape which ensures a 'long movement' for a small volume of liquid delivered out of the jet.

6. Pipette



Use: It is used for delivering accurately a fixed volume of liquid

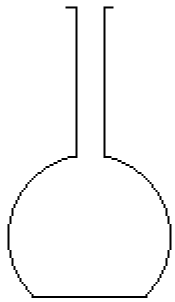
A pipette has to be filled carefully by sucking up the liquid or by using a special adaptor.

The volume of the liquid delivered is marked on the bulb and it is usually25.0cm^{3}or 25.0ml. It is accurate to one decimal place. Most of the liquid is held in the bulb.

7. Flasks

There are three main types of flasks; each type is usually used for holding liquids.

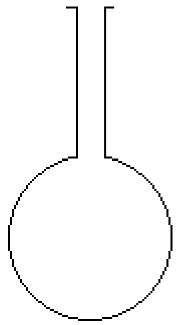
(a) Flat bottomed flask



The flat bottomed flask can stand on the table.

Use: It is used in carrying out reactions involving a solid and a liquid.

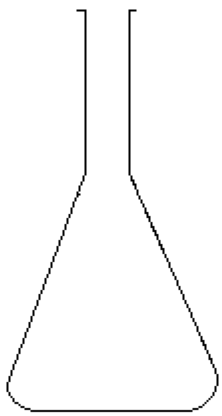
(b) Round bottomed flask



Its shape enables uniform heating of the liquid it contains

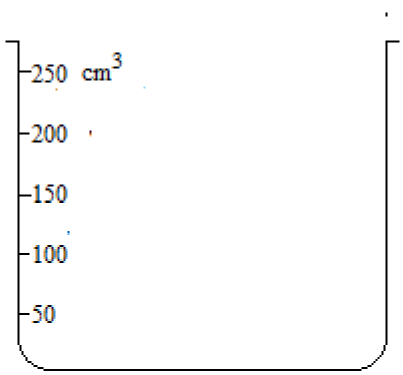
Use: It is used for heating liquids for longer periods

(c) Conical flask



Use: It is used for mixing liquids while shaking.

8. Beakers

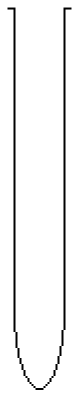


Beakers are in different sizes

Use: They are used on tripod stand and gauzes for heating liquids

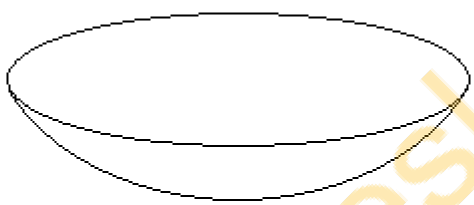
They are used for mixing liquids while using volumes which are too big for the test tube.

9. Test tube



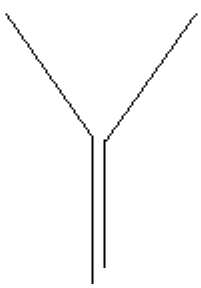
Use: It is used for heating and mixing liquids or solids

10. Evaporating dish



Use: It is used for drying substances or keeping them free from moisture.

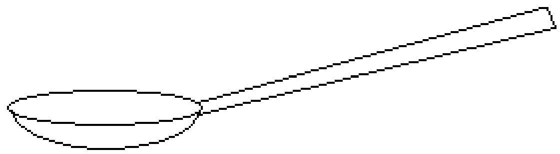
11. Funnel



Use: It is used in pouring liquids.

It is used to guide liquids and other substances into containers.

12. Spatula



It is a flexible metal, plastic or rubber utensil

Use: It is used to scoop, lift, spread, or mix substances

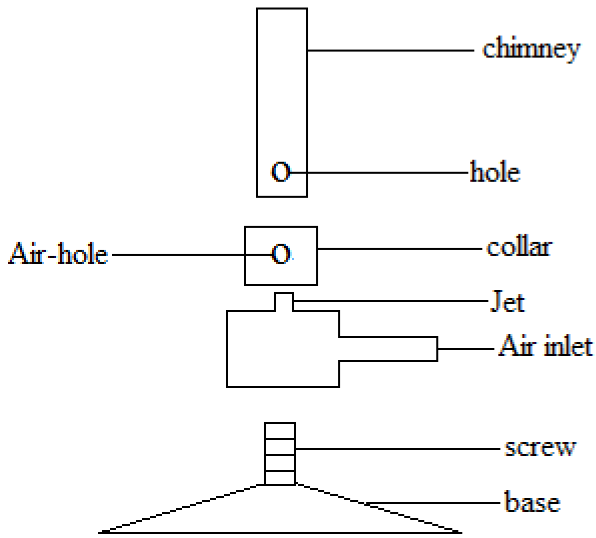
13. The Bunsen burner

The Bunsen burner is the most common tool for heating.

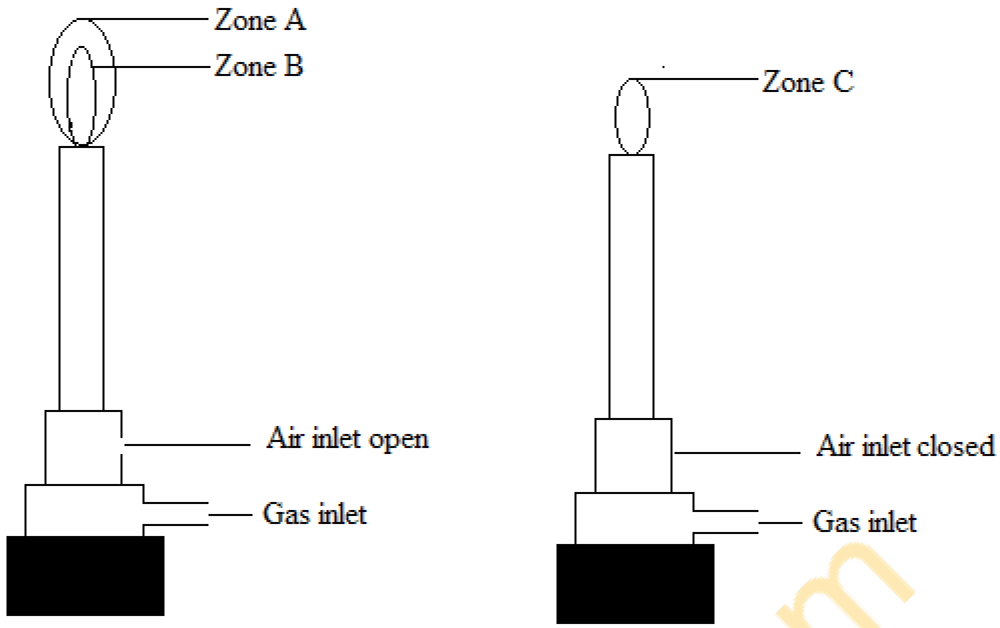
It is connected to the cylinder or gas tap and lighted.

Use: It is used as a source of heat.

Component parts of a Bunsen burner



Flames produced by the Bunsen burner



Zone A

It is a blue flame

It produces the greatest amount of heat because methane gas is completely burnt. It is the hottest part of the flame and it used for heating.

$$\text{Methane}_{(g)} + \text{Oxygen}_{(g)} \rightarrow \text{Carbon dioxide}_{(g)} + \text{Water}_{(l)}$$

Zone B

It is a blue green flame

It contains unburnt hydrocarbons

It appears blue green because of the incomplete combustion of methane.

$$\text{Methane}_{(g)} + \text{Oxygen}_{(g)} \rightarrow \text{Carbon monoxide}_{(g)} + \text{Water}_{(l)}$$

It has the lowest temperature and it is called non luminous.

Zone C

It is a yellow flame

It is luminous because it gives out light

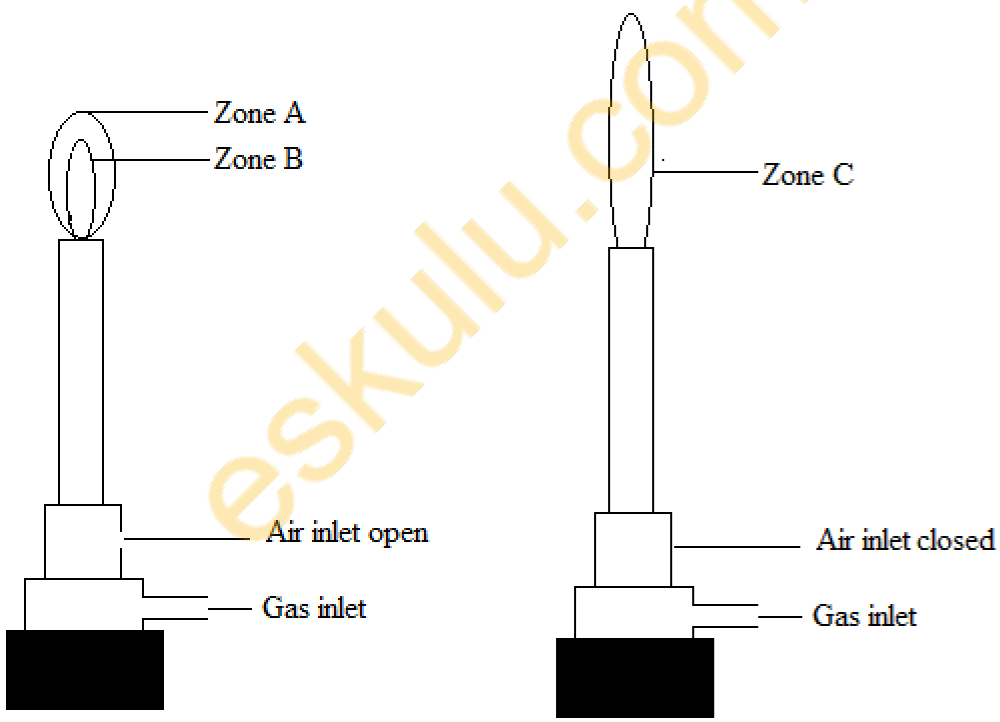
The yellow flame is called a ‘dirty flame’ because it coats things held inside it with a black deposit. This black substance is carbon.

$\text{Methane}_{(g)} + \text{Oxygen}_{(g)} \rightarrow \text{Carbon}_{(s)} + \text{Water}_{(l)}$

The cause of the colour in zone C is carbon charring due to incomplete combustion of methane

Exercise

1. The diagram below represent the flames on a burner using methane as the fuel



The equation for the complete combustion of methane in zone A is:

$\text{Methane}_{(g)} + \text{Oxygen}_{(g)} \rightarrow \text{Carbon dioxide}_{(g)} + \text{Water}_{(l)}$

(a) What colour would you expect the flame to be in

(I) Zone B
(II) Zone C

(b) Explain what causes the colour in zone C
(c) Give an equation for the combustion of methane in zone C.
2. The table below shows some apparatus used in the laboratory.

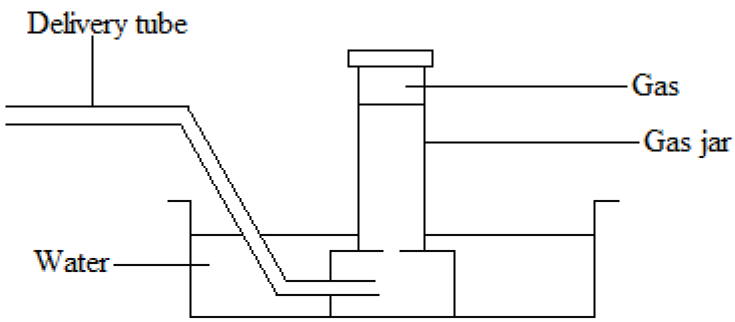
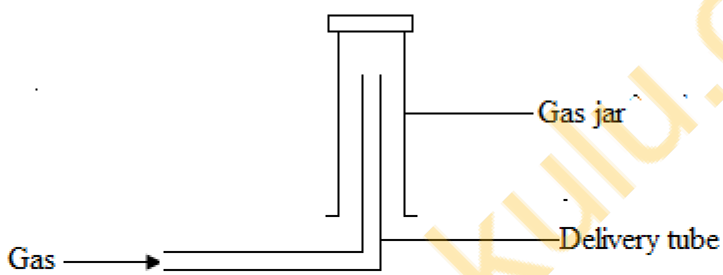
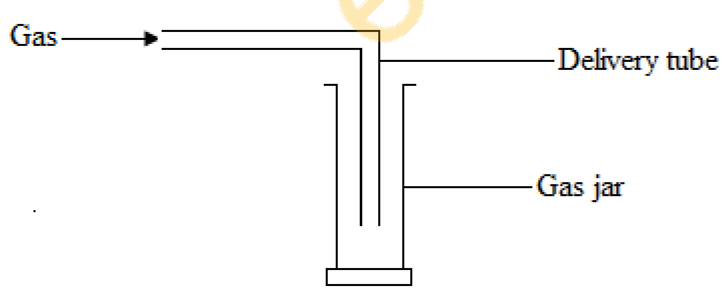
Separating funnel	Beaker	Gas jar
Burette	Spatula	Laboratory thermometer
Desiccator	Bunsen burner	Bee-hive shelf
Evaporating dish	Pipette	Tripod stand

State the apparatus used…

(a) as a source of heat
(b) for measuring a fixed volume of liquid
(c) for drying substances or keeping them free from moisture
(d) for measuring temperature
(e) for separating immiscible liquids

Collection of gases

Density of the gas and the solubility of the gas in water are the two factors used to determine the method used to collect a gas:

Method of collection	Type of gases to be collected
Displacement of water 	Gases that are insoluble in water e.g. hydrogen and oxygen
Displacement of air – upward delivery 	Gases that are less dense than air e.g. hydrogen and ammonia
Displacement of air – down ward delivery 	Gases that are denser than air e.g. hydrogen chloride and carbon dioxide.

Criteria of purity

Definition: This refers to all those physical properties which are characteristic of a pure substance.

The purity of a substance can be tested by determining:

(a) the boiling point
(b) the freezing point
(c) the melting point
(d) the density

Impure substances show variations in physical properties

Importance of purity of substances

Purity of substances like food, drugs and water is important because very small amounts of impurities may cause serious illness or death.

The companies processing food and manufacturing drugs check regularly to ensure that their products are pure.

Separation techniques

Separation techniques are methods used to obtain pure substances from their mixtures.

Substances usually exist as mixtures and special ways are used to separate them.

Terms used in separation techniques

Mixture: It is a substance which consists of two or elements which are physically combined but not chemically combined.

Solution: It is a mixture made up of a solvent and a solute.

Solvent: It is a liquid in which a solute dissolves.

Solute: It is a substance that dissolves in a solvent. A solute can either be a solid, liquid or gas.

Miscible liquids: They are liquids that can mix completely.

Immiscible liquids: They are liquids that do not mix completely.

Residue: It is a solid that is trapped on the filter paper during filtration.

Filtrate: It is a clear liquid collected after filtration.

Methods of purification

1. Filtration

This is a method of separating an insoluble solid from a liquid using a filter.

Experiment

Aim: To separate a mixture of an insoluble solid a and soluble solid e.g. a mixture of sand and salt

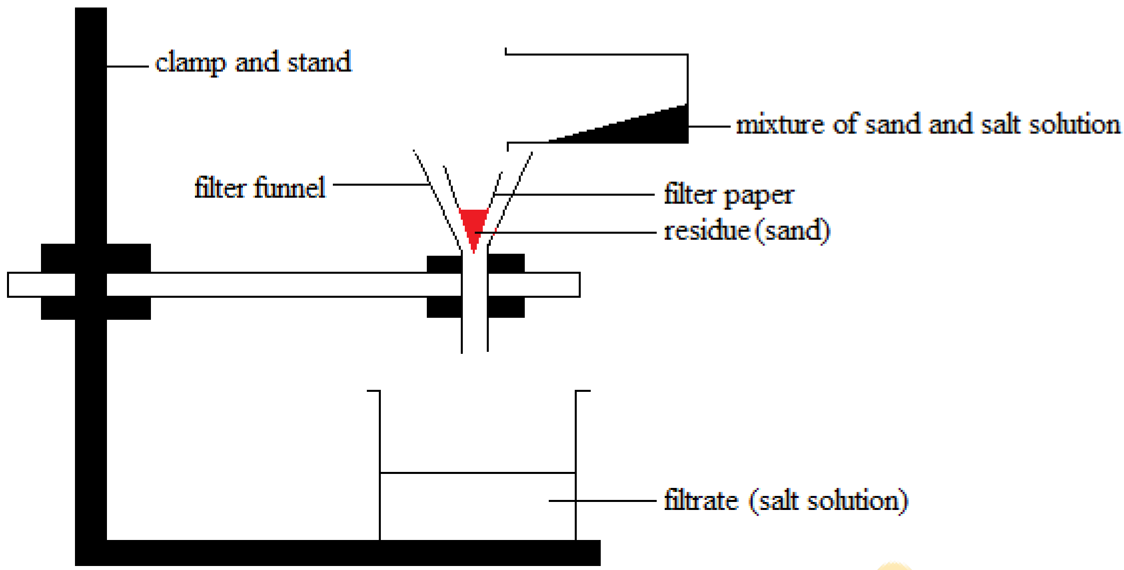
Apparatus

- Filter paper
- Filter funnel
- Conical flask
- Beaker
- Mixture of sand and salt

Method

Place a mixture of sand and salt in a beaker. Add water and stir. The salt which is soluble dissolves in water to form a salt solution

Pour the mixture of sand and salt solution into a filter funnel containing a filter paper as shown below



Put a little amount of the filtrate into the evaporating dish. Heat the filtrate until all the water is driven off.

Observation

The salt solution passes through the filter paper and is collected in the conical flask as a filtrate while sand remains on the filter paper as a residue. When the filtrate (salt solution) is heated in the evaporating dish, salt which is a solute will remain in the evaporating dish while water which is a solvent will go away as steam.

Conclusion

A mixture of two solids, one soluble and the other insoluble can be separated by dissolving, filtration and evaporation.

Application of filtration

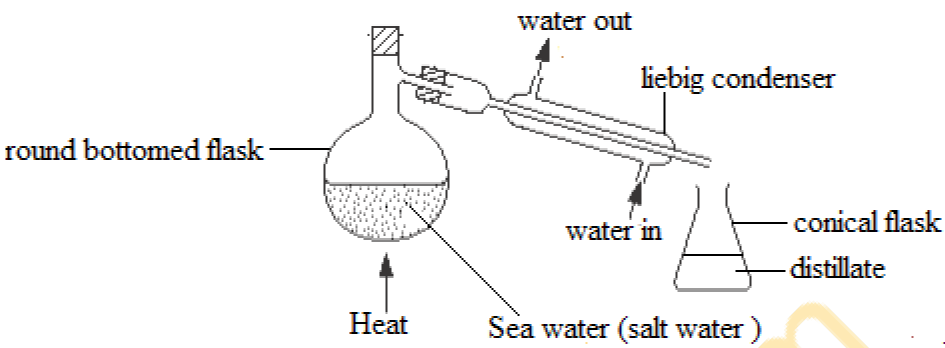
(a) It is used in the purification of drinking water at the water works.
(b) It is used in car engines to remove impurities from oil, petrol etc.

2. Distillation

It is a process of vapourizing a liquid and then condensing the vapour.

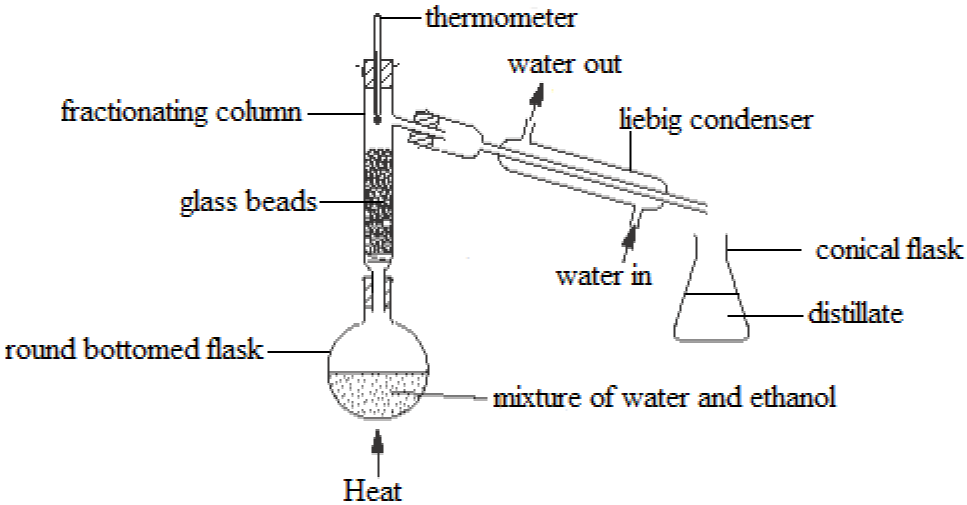
(a) Simple distillation

It is a process used to separate a pure liquid from a solution containing dissolved solids. For example, simple distillation can be used to separate pure water from sea water



(b) Fractional distillation

It is process of separating a mixture of two or more miscible liquids which have different boiling points. For example, fractional distillation can be used to separate ethanol from a mixture of ethanol and water. As a rough guide, the boiling points of the liquids to be separated should be at least 20^oC apart.Ethanol and water have different boiling points. Ethanol boils at 78^oC and water boils at 100^oC. When temperature reaches 78^oC, ethanol, which is more volatile than water, boils off first. As the vapour passes through the glass beads in the fractionating column, water vapour condenses and returns back to the flask. Only ethanol vapour reaches the fractionating column and enters the liebig condenser where it condenses. The liquid ethanol finally collects in the conical flask.



Properties upon which the mixture (liquids) depends to allow separation

1. The liquids must be miscible
2. The liquids must have difference boiling points

Note

- The fractionating column is long tube filled with glass beads. The glass beads provide a large surface area for condensation and allow efficient separation of the components in the mixture.
- The liebig condenser is kept in the slanting position to avoid the distillate formed by condensation from running back into the fractionating column; and also to ensure that cold water completely surrounds the inner glass tube where the vapour passes in order to provide maximum cooling and avoid loss of vapour.
- Cold water enters the condenser from the bottom to ensure that the region is the coldest so that all the vapour entering the condenser will turn into a liquid.
- The condenser cools the vapour causing it to condense into liquid
- The thermometer is placed at the top of the fractionating column so that it registers the temperature of the vapour.

Distillate

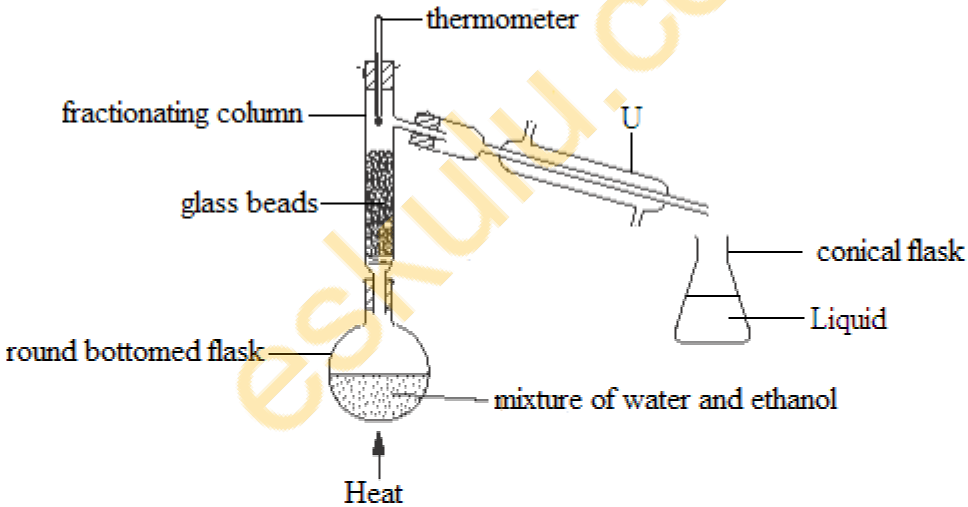
It is a pure and condensed liquid obtained by distillation.

Industrial application of fractional distillation

(a) Separation of various components of crude oil (petroleum)
(b) Used in the manufacture of spirits such as whisky, rum, gin etc.
(c) Separation of liquid air into nitrogen and oxygen (nitrogen boils at $-196^{o}C$ while oxygen boils at $-183^{o}C$)

Example

1. The diagram below shows the experiment on separation of a mixture



(a) What is the name of the separation process shown in the diagram?
(b) State two properties of the mixture that allows the separation to take place
(c) What is the function of the apparatus labeled U?
(d) Name the first part of the mixture that will be collected in the conical flask
(e) Suggest one industrial use of the separation technique shown in the diagram

Solution

(a) Fractional distillation
(b) Liquids must be miscible / mix completely (but not react)

Liquids must have different boiling points

(c) Cools the vapour causing it to condense into liquid
(d) Ethanol
(e) Separation of crude oil (at oil refinery plant)

Exercise

1. Two miscible liquids with boiling points $78^{o}C$ and $100^{o}C$ were mixed accidentally.

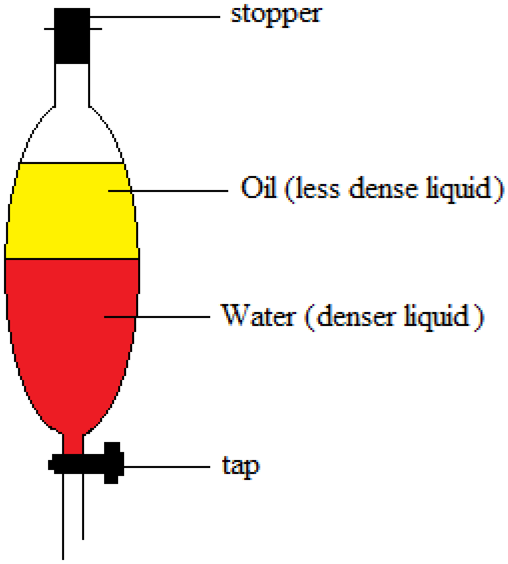
   (a) Name the process which can be used to separate the mixture
   (b) Draw a labeled diagram showing the arrangement of the apparatus used to separate the mixture.

3. Separating funnel

The separating funnel is used to separate two or more immiscible liquids e.g. a mixture of water and oil

The method depends on the differences in densities of liquids to be separated.

The less dense liquid floats on the surface while the denser liquid sinks to the bottom.



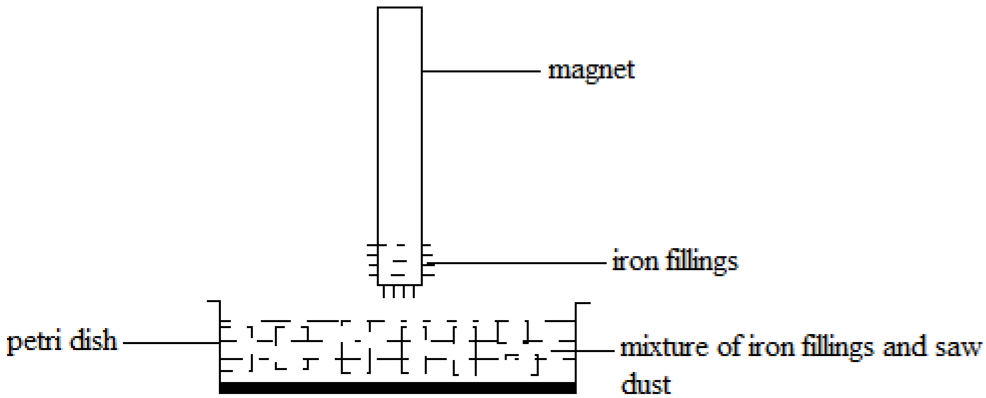
Exercise

1. Water and oil are immiscible. How would you separate a mixture of the two?

4. Magnetism

It is used to separate a mixture when one component is magnetic and other one is non-magnetic.

For example magnetism can be used to separate a mixture of saw dust and iron fillings. The magnet attracts iron fillings, only leaving particles of saw dust in the petri dish.



5. Decantation

Decantation is the process of separating by carefully pouring a solution from a container in order to leave the precipitate (solid particles) at the bottom. Solid impurities are allowed to settle down and the liquid is poured out leaving the solid at the bottom of the container. For example it can be used to separate a mixture of sand and water

6.Crystallization

This is a process of separating pure solid from an impure solution.

For example crystallization can used to obtain copper (II) sulphate crystals from an impure copper (II) sulphate solution.

Experiment

Aim: To obtain copper (II) sulphate crystals from an impure copper (II) sulphate solution.

Apparatus

- Copper (II) sulphate solution
- Beaker
- Stirring rod
- Filter paper
- Evaporating dish
- Tripod stand

- Bunsen burner

Method

Dissolve an impure solid in a solvent

Filter off the solution into an evaporating dish

Heat the solution to evaporate most of the solvent

Cool the concentrated solution. Some solids appear as pure crystals

Pour off the solution to obtain the crystals. Dry the crystals on the filter paper

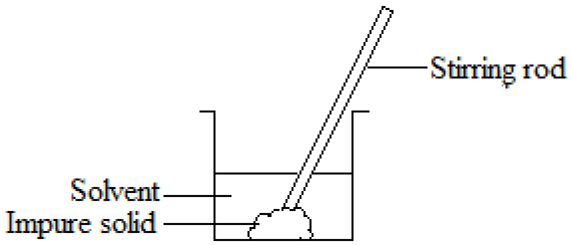
Observation

Some solids appear as pure crystals

Conclusion

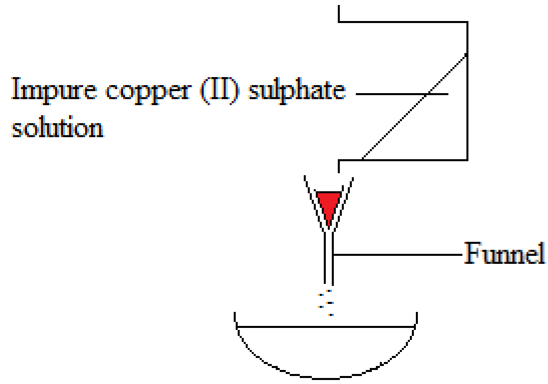
Copper (II) sulphate crystals from an impure copper (II) sulphate solution can be obtained by dissolving, filtration and evaporation, cooling and drying.

1. Dissolve



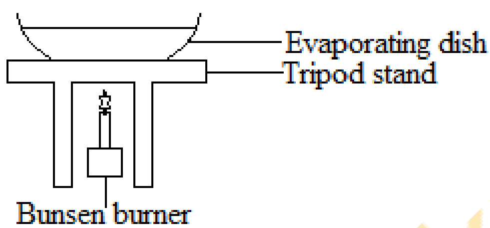
The impure solid is dissolved in a solvent

2. Filter



The solution is filtered into an evaporating dish

3. Evaporate

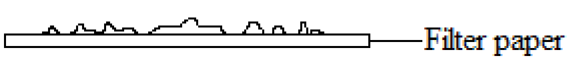


The solution is heated to evaporate most of the solvent

4. Cool

The hot solution is allowed to cool. The solid appears as pure crystals.

5. Dry



The cold solution is poured off to obtain the crystals. The crystals may be dried by pressing them between sheets of filter paper

Differences between crystallization and evaporation

Crystallization must be differentiated from evaporation to dryness

(a) In crystallization, the solvent is only partially evaporated, leaving a small amount of solution in which the crystals form. Impurities may be left behind in the solution when the crystals are filtered off.
(b) In evaporation to dryness, all the solvent is removed. The crystals formed may be impure.

7. Chromatography

Chromatography is a technique for separating mixtures of solutes using a solvent and a separating medium.

In the case of paper chromatography, the separating medium is paper and the solvent is ethanol.

Substances in a mixture are separated according to their solubilities in the same solvent.

The more soluble component will tend to remain in the solution and travel further up the chromatogram while the less soluble component will separate out on to paper.

Experiment

Aim: To separate out components of black ink by paper chromatography.

Procedure

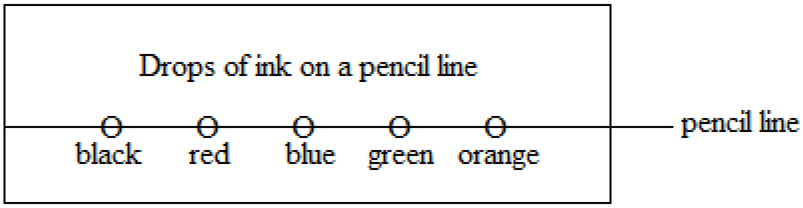
Use a pencil to draw the start line because ink contains dyes which contaminates the solvents

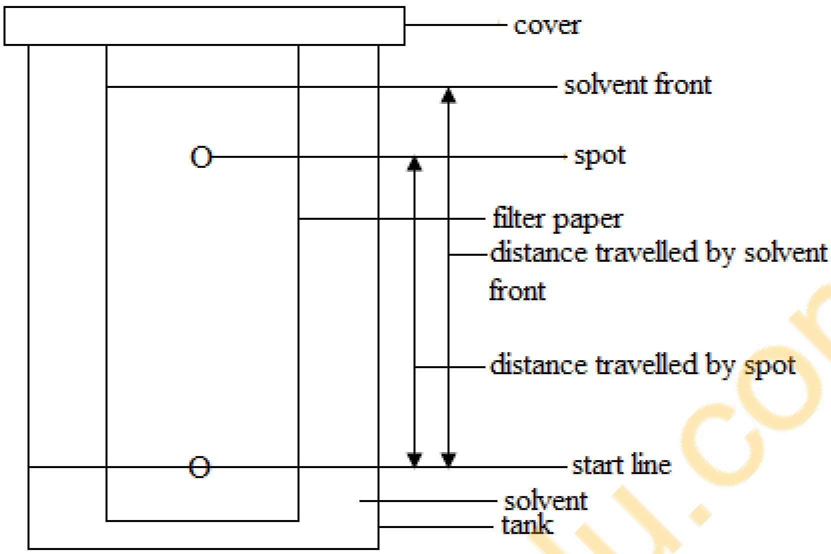
Use the black ink sample to make a small dot on the start line, together with other coloured ink to use as reference. The dots should be small to prevent the spreading of dyes sideways and thereby getting mixed up with other spots next to them.

Fold the paper into a cylinder and place it into a beaker containing the solvent, ensuring that the start line is above the solvent level.

Cover the beaker while the chromatogram develops.

Remove the chromatogram from the beaker just as the solvent reaches the top of the paper.

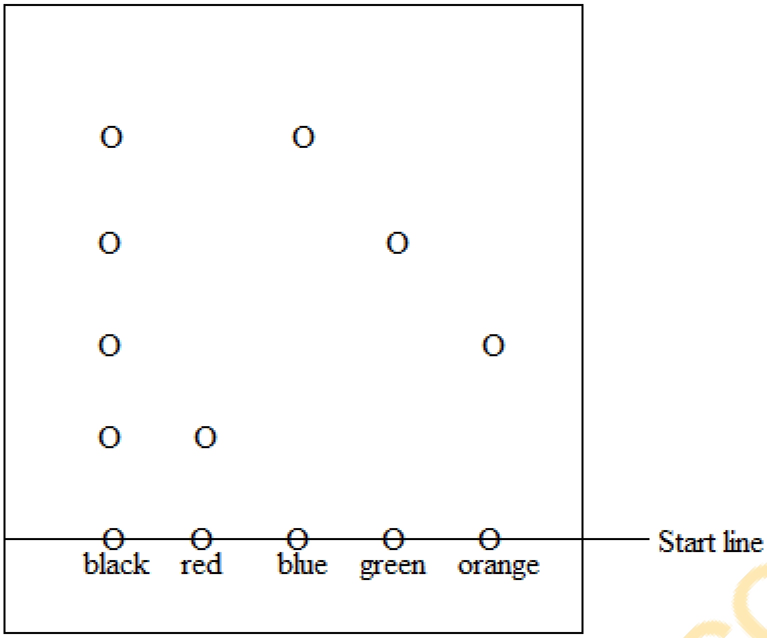




Note

- If the start line is below the solvent level, the sample dots will dissolve into the solvent instead of moving up the paper.
- In descending paper chromatography, a longer sheet of paper can be used because the solvent moves faster and longer distance
- Descending paper chromatography has an advantage because as the solvent flows down the paper by capillary action, it is aided by gravity.
- The tank should be kept closed to keep the air saturated with solvent vapour to reduce evaporation of the moving solvent.
- One advantage of using a locating agent in the experiment is that it reacts with the spots and gives the latter colours for easy location

Interpretation of results



Dots that have travelled the same distance from the start line in the same solvent belong to same substance.

Conclusion

Black ink contains four coloured components; red, blue, green and orange.

Retardation factor

Alternative term: Reference factor

Symbol: R_f

Definition: Retardation factor, R_f, is the ratio of the distance travelled by the spot, compared with the distance travelled by the solvent front, both measured from the start line

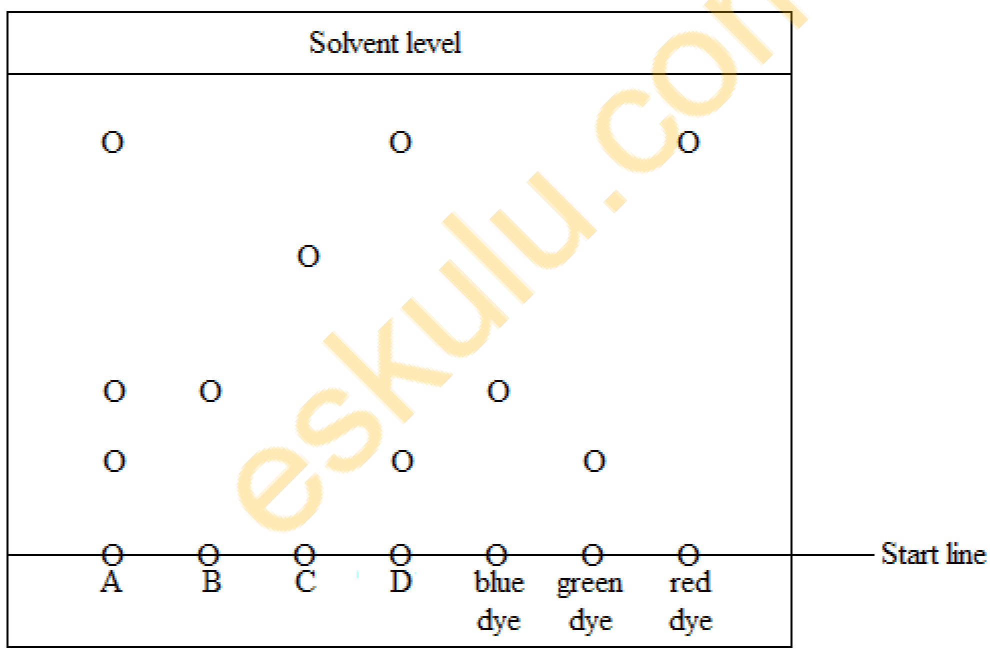
Formula: $R_f = \frac{\text{Distance travelled by spot from start line}}{\text{Distance travelled by the solvent front from start line}}$

Application of chromatography

1. Separating pigments from plants
2. Identifying flavouring components in food stuffs
3. Separating amino acids from proteins
4. Separating antibiotic drugs from their growing media

Example

1. The diagram below shows a chromatogram obtained using solutions of three single dyes (blue, green and red) and four other solutions (A, B, C and D).



(a) Which of the solutions A, B, C and D contain the following:

(I) One dye only
(II) Three of the dyes
(III) Green and red only

(IV) A dye rather than blue, green and red

(b) In preparing the chromatogram, the following instructions were given. Suggest a reason for each instruction;

(I) The start line should be drawn with pencil rather than ink
(II) At the end of the experiment, the solvent front should be near the top of the paper
(III) The spots of solutions and dyes on the starting line should be small.

Solution

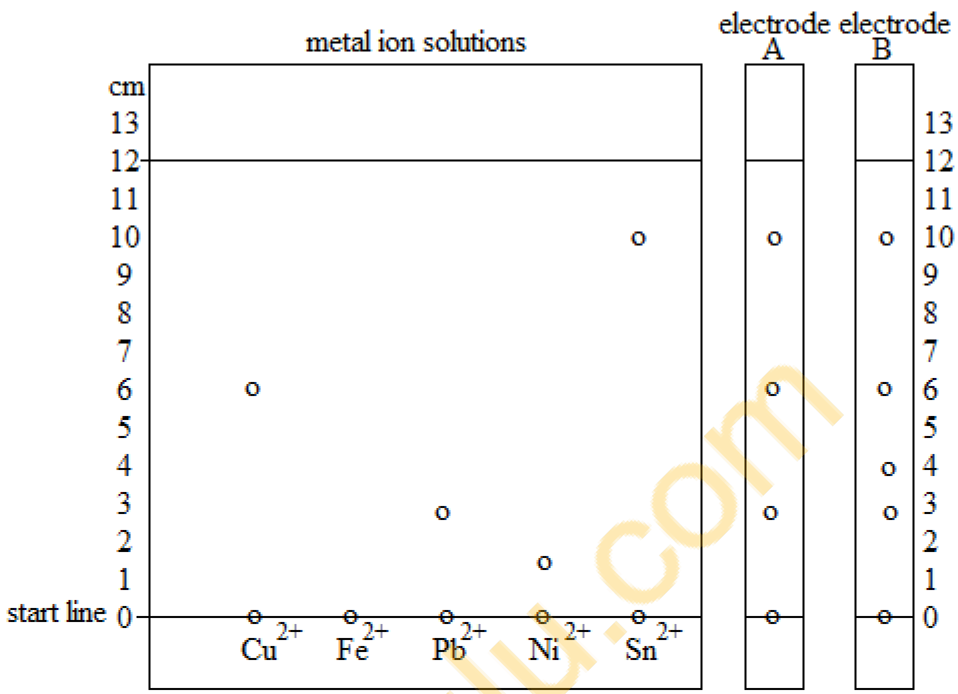
(a) (I) Solution B

(II) Solution A
(III)Solution D
(IV)Solution C

(b) (I) Ink contains dyes, it thus contaminates the solvent and gives a wrong result

(II) This is to ensure that the separation of the dyes is complete
(III) To prevent spreading of the dyes sideways thereby getting mixed up with other .
. spots next to them.

2. In order to compare the impurities present in the electrodes from two lead –acid batteries, samples from the electrodes were dissolved in a suitable acid and the resulting solutions were chromatographed, together with solutions containing five known metal ions. After treating with a locating agent, the chromatograms were shown below



(a) What is the advantage of using a locating agent in the experiment?
(b) Were the two electrodes from the same manufacturer? Give a reason for your answer
(c) In a chromatogram, the R_f value of an ion is defined as:

$$R_f \text{value} = \frac{\text{Distance travelled by spot of the iron}}{\text{Distance travelled by the solvent}}$$

In this chromatogram, R_f value for Pb^{2+} is $\frac{3}{12} = 0.25$

(I) What is the R_f value for Cu^{2+}?
(II) Find the distance travelled by the spot of Fe^{2+} which has an R_f value of 0.75

Solution
(a) It reacts with the spots and gives the latter colours for easy location

(b) No, there chromatograms show that they have different ions

(c) (I) R_f value for Cu^{2+}= $\frac{\text{Distance travelled by spot of the } Cu^{2+}}{\text{Distance travelled by the solvent}}$

$$= \frac{6}{12}$$

= 0.5

(II) R_f value = $\frac{\text{Distance travelled by spot of the iron}}{\text{Distance travelled by the solvent}}$

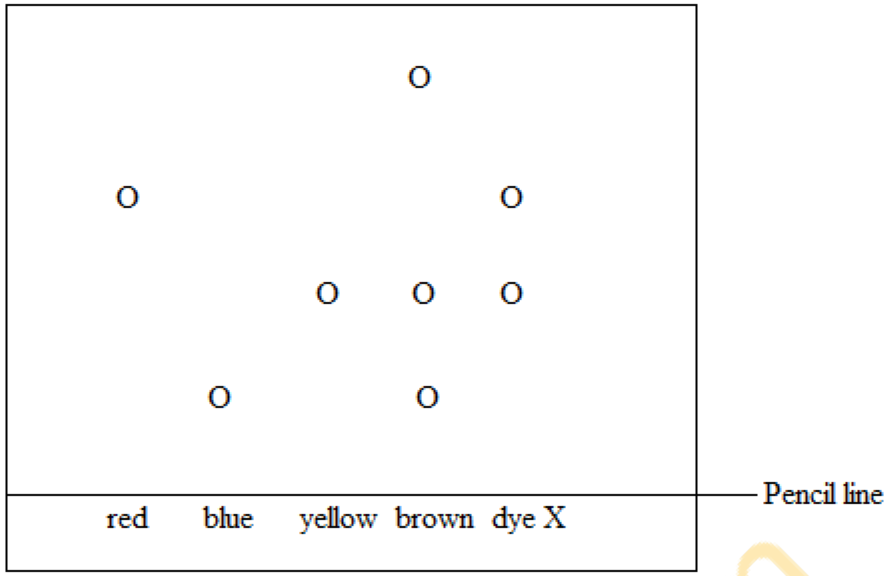
Distance travelled by spot of iron = R_f x distance travelled by solvent

= 0.75 x 12

= 9cm

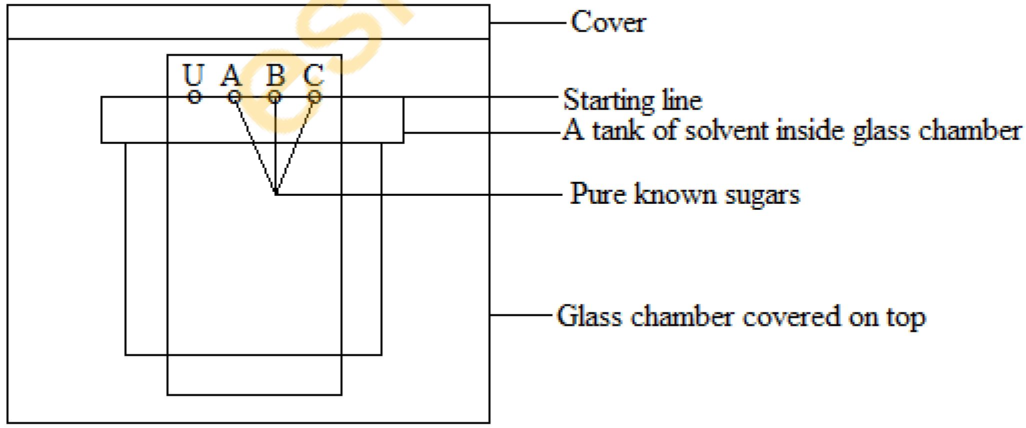
Exercise

1. Naosa Jane wants to find out which coloured dyes have been mixed together to make dye X. She separates a sample of dye X and samples of coloured dyes using paper chromatography. Her results are shown below.

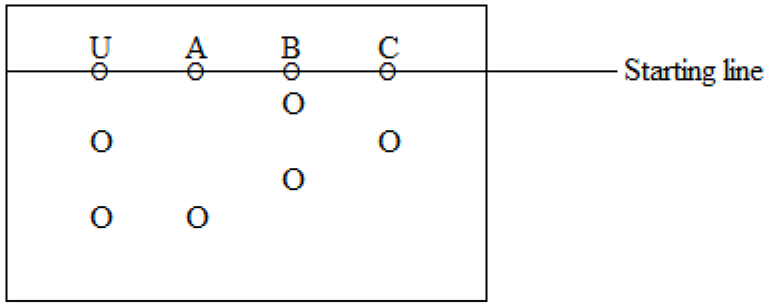


(a) Explain why the line is drawn in pencil and not in ink?
(b) Which colours are present in dye X?
(c) Which coloured dye contains a substance not present in any of the other coloured dyes?

2. The diagram below shows the apparatus that could be used to separate and identify components of an unknown mixture of sugars (U). Study the diagram and answer the questions that follow.



After development, the paper looks like this.



(a) What disadvantage has this technique over the one with the solvent moving upwards
(b) Identify the sugars in the unknown sample using the letters in the diagram
(c) By what process does the solvent move in the experiment above?
(d) Why should the tank be kept closed in the experiment?

Summary of separation techniques

Separation technique	**Substances to be separated**	**Example**
Filtration	Insoluble solid and liquid	Muddy water
Crystallization / Evaporation	Solute (soluble solid) from its solution	Salt solution
Distillation	Solvent from its solution	Salt solution
Fractional distillation	Miscible liquids with different boiling points	Ethanol and water Crude oil Liquid air
Decantation / Sedimentation	Insoluble suspension settles to form sediment	Mealie – meal and water
Separating funnel	Immiscible liquids	Oil and water
Floatation	Less dense solid and liquid	Charcoal dust and water
Magnetism	Magnetic materials	Iron fillings and sulphur powder
Paper chromatography	Dissolved substances	Dyes and pigments of ink

Example

1. Complete the following statement

   (a) A _____________is the substance left on a filter paper
   (b) The ____________is the liquid that passes through a filter paper
   (c) ______________is the process of separating a liquid from solid sediment by pouring
   (d) _______________is used to separate vinegar from dressing
   (e) The process of _______a liquid and then _______the vapour is known as_______

2. Name the chemical technique, which could usefully be used to separate water from a solution of sodium chloride in water.
3. Sugar cannot be separated from sugar solution by filtering. Explain why.

Solution

1. (a) Residue (b) Filtrate (c) Decantation (d) Centrifugation
   (e) Evaporation, condensation, distillation
2. Distillation
3. Because it is spread all through the solvent in tiny particles

Elements, mixtures and compounds

1. Elements

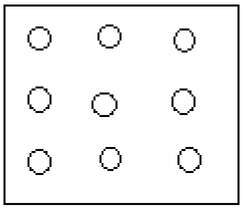
Definition: An element is a pure substance which cannot be split into two or more other simpler substances by chemical means.

Examples of elements

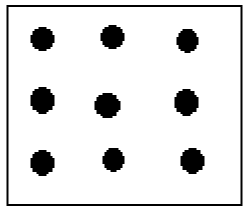
- Iron
- Sulphur
- Magnesium

- Oxygen

Diagrammatical representation of elements



or



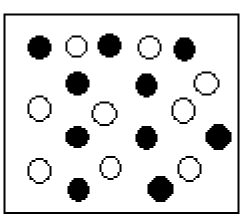
2. Mixture

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

Examples of mixtures

- Air: Air is a mixture of oxygen, carbon dioxide, nitrogen and other gases.
- Sugar solution: Sugar solution is a mixture of sugar and water.
- Brass: Brass is a mixture of zinc and copper.

Diagrammatical representation of a mixture



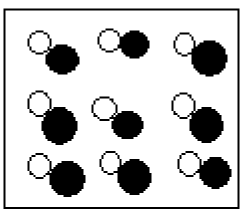
3. Compound

Definition: A compound is a substance which consists of two or more elements chemically combined.

.**Examples of compounds**

- Water: Water consists of the elements hydrogen and oxygen.
- Common salt: Common salt consists of the elements sodium and chlorine.
- Carbon dioxide: Carbon dioxide consists of the elements carbon and oxygen.
- Sugar: Sugar consists of the elements carbon, hydrogen and oxygen.
- Iron sulphide: Iron sulphide consists of the elements iron and sulphur.

Diagrammatical representation of a compound



Differences between mixtures and compounds

Mixture	Compound
The substances in a mixture can be separated by physical means. It is easy to separate a mixture into its components because each component keeps its own properties.	The elements in a compound cannot be separated by physical means. The particles of elements are combined chemically in a fixed ratio. It is difficult to separate a compound into its constituent elements
Energy is not usually given out or absorbed when mixing occurs	Energy (heat, light or sound) is usually given out when a compound is formed
The properties of a mixture (density, colour) are an average of those of the substances in it	The properties of a compound are quite different from those of the elements in it. This is because a chemical change has taken place.
The composition of a mixture is variable. The substances can be present in any proportions by mass. In a mixture, the particles of each substance remain separate and the number of each can vary	The composition a compound is fixed. The elements are combined in definite proportions by mass

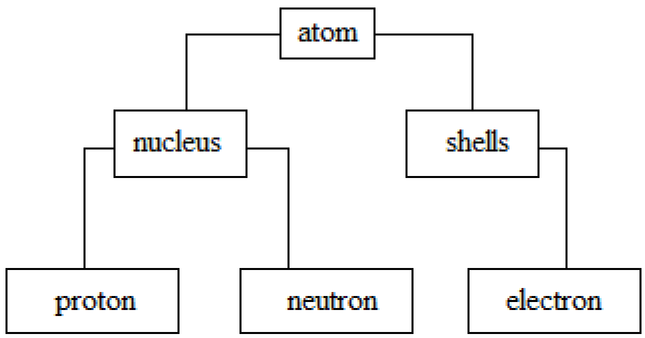
Exercise

1. Consider the following list of common substances

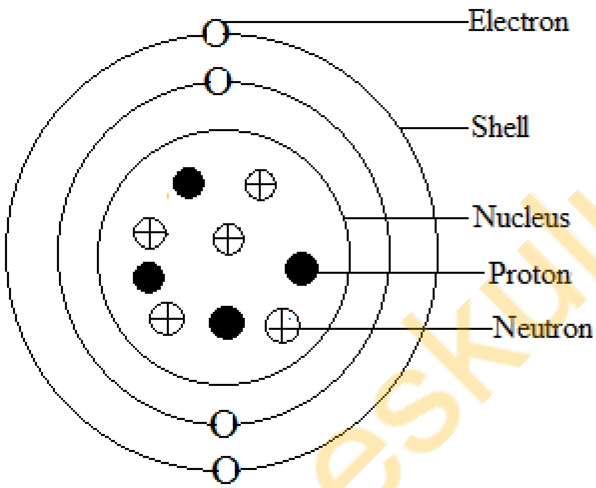
   Air, salt, brass, sugar, glass, water, limestone
   Choose from the list a substance which:
   (a) is a mixture containing both elements and compounds
   (b) is a mixture of compounds
   (c) is a mixture of elements
   (d) is a compound containing only two elements
   (e) is a compound which contains carbon

Structure of an atom

Definition: An atom is the smallest particle of an element that can take part in a chemical reaction.



Atoms are electrically neutral. They have no overall charge. This is because the number of protons is equal to the number of electrons



Fundamental particles of an atom

1. Electron

It is a negatively charged particle.

It has a charge of -1

It is found in shells of around the nucleus of an atom. A shell is a concentric ring around the nucleus

It has a mass of $\frac{1}{1840}$ atomic mass units (a.m.u)

2. Proton

It is a positively charged particle

It has a charge of +1

It is found in the nucleus of an atom

It has a mass of 1atomic mass units (a.m.u)

3. Neutron

It is a neutral particle

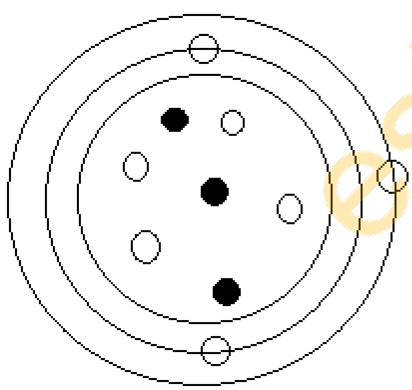
It has no charge

It is found in the nucleus of an atom

It has a mass of 1atomic mass units (a.m.u)

Example

1. The diagram below represents an atom of lithium.



What do the symbols represents?

Solution

⊖	electrons
●	protons
○	neutrons

Exercise

1. Complete the table below

Particle	Relative charge	Position in the atom	Relative mass
	−1	Shells around the nucleus	
Proton		Nucleus	
			1

Proton number

Alternative term: Atomic number

Symbol: Z

Definition: It is the number of protons in the nucleus of an atom

Mass number

Alternative term: Nucleon number

Symbol: A

Definition: It the sum of protons and neutrons in the nucleus of an atom

Formula: A = Z + N

A = mass number

Z = proton (atomic number)

N = number of neutrons

Example

1. An atom of sodium has 11 protons and 12 neutrons. Calculate its mass number.

Data	Solution
A =? Z = 11 N = 12	A = Z + N A = 11 + 12 A = 23

2. An atom has the notation $^{39}_{19}Y$. State the number of protons (p), electrons (e) and neutrons (n) in one atom of Y.

Solution

p = 19

e = 19

n = 39 – 19

= 20

Exercise

1. An atom of aluminium has a mass number of 27 and has 13 protons. Work out the number of neutrons.
2. An atom has the notation $^{32}_{15}X$. State the number of protons, electrons and neutrons in one atom of X.

Atomic numbers and mass numbers for the first 20 elements

Element	Symbol	Atomic number	Mass number
Hydrogen	H	1	1
Helium	He	2	4

Lithium	Li	3	7
Beryllium	Be	4	9
Boron	B	5	11
Carbon	C	6	12
Nitrogen	N	7	14
Oxygen	O	8	16
Fluorine	F	9	19
Neon	Ne	10	20
Sodium	Na	11	23
Magnesium	Mg	12	24
Aluminium	Al	13	27
Silicon	Si	14	28
Phosphorous	P	15	31
Sulphur	S	16	32
Chlorine	Cl	17	35.5
Argon	Ar	18	40
Potassium	K	19	39
Calcium	Ca	20	40

Nuclides

A nuclide consists of a symbol with the atomic number in front of the symbol and just below it and the mass number also in front and just above the symbol.

Nuclide notation: ${}^{A}_{Z}X$

A = mass number

X = symbol of an atom

Z = proton (atomic number)

Isotopes

Definition: Isotopes are atoms of the same element having the same number of protons but different number of neutrons (and mass numbers)

Carbon has three isotopes.

Carbon −12, $^{12}_{6}C$	p = 6 n = 12 – 6 = 6
Carbon − 13, $^{13}_{6}C$	p = 6 n = 13 – 6 = 7
Carbon − 14, $^{14}_{6}C$	p = 6 n = 14 – 6 = 8

Hydrogen has three isotopes

Hydrogen − 1, $^{1}_{1}H$ (Ordinary hydrogen, H)	p = 1 n = 1 – 1 = 0
Hydrogen − 2, $^{2}_{1}H$ (Deuterium, D)	p = 1 n = 2 – 1 = 1
Hydrogen − 3, $^{3}_{1}H$ (Tritium, T)	p = 1 n = 3 – 1 = 2

Chlorine has two isotopes

Chlorine − 35, $^{35}_{17}Cl$	p = 17 n = 35 – 17 = 18
Chlorine − 37, $^{37}_{17}Cl$	p = 17 n = 37 – 17

	= 20

Some uses of isotopes

1. Hydrogen isotopes are used in the study of nuclear energy
2. Radioisotopes e.g. radio cobalt, ${}^{60}_{27}$Coare used in the treatment of cancer.
3. Isotopes e.g. Uranium −235 isotopes, are used in the manufacture of nuclear weapons
4. Carbon −14 isotope is used in radio carbon dating machine to determine the age of fossils.

Mass spectrometer

The mass spectrometer is used for the separation of isotopes of an element and determination of their respective masses.

Relative atomic mass

Symbol: A_r

Definition: Definition: Relative atomic mass of an element is the mass of one atom of an element compared to $\frac{1}{12}$ the mass of carbon − 12 isotope.

Formula:$A_r = \frac{A_r \text{ of isotope A x percentage abundance}}{100\%} + \frac{A_r \text{ of isotope B x percentage abundance}}{100\%}$

Example

1. Chlorine has two main isotopes, ${}^{35}_{17}$Cl and ${}^{37}_{17}$Cl with abundances of 75% and 25% respectively. Calculate the relative atomic mass for chlorine.

$$A_{r(\text{Cl})} = \frac{A_r \text{ of } {}^{35}_{17}\text{Cl x percentage abundance}}{100\%} + \frac{A_r \text{ of } {}^{37}_{17}\text{Cl x percentage abundance}}{100\%}$$
$$= \frac{35 \text{ x } 75\%}{100\%} + \frac{37 \text{ x } 25\%}{100\%}$$

$= \frac{2625}{100} + \frac{925}{100}$

$= 26.25 + 9.25$

$= 35.5$

Note

The relative atomic mass of chlorine is 35.5 (not a whole number) because it is the average of two isotopes i.e. chlorine − 35 and chlorine − 37.

Exercise

1. Chlorine has two isotopes, $^{35}_{17}Cl$ and $^{37}_{17}Cl$

   (a) Define the term isotopes
   (b) State the number of neutrons in each of the following isotopes of chlorine

      (I) Chlorine – 35
      (II) Chlorine – 37

   (c) Explain why the relative atomic mass of chlorine on the periodic table is not a whole number?
   (d) What is the difference in the nuclei of the isotope with nucleon number 7 and its isotope whose nucleon number is 6?
   (e) Define relative atomic mass

2. Natural Neon occurs as a mixture of three isotopes, 90.92% $^{20}_{10}Ne$, 0.257% $^{21}_{10}Ne$ and $^{22}_{10}Ne$.

   (a) What is the relative percentage abundance of $^{22}_{10}Ne$?
   (b) Calculate the relative atomic mass of neon giving your answer to the nearest whole number

3. The table below shows the three isotopes of the element hydrogen and their nucleon (mass) numbers.

(a) Complete the table to show the number of particles in the three nuclides.

Symbol	Name	Nucleon (mass) number	Protons	Neutrons
H	Hydrogen	1		
D	Deuterium	2		
T	Tritium	3		

(b) The boiling point of D_2O is 101.6^oC but that of H_2O is100.0^oC. Suggest a reason for this difference

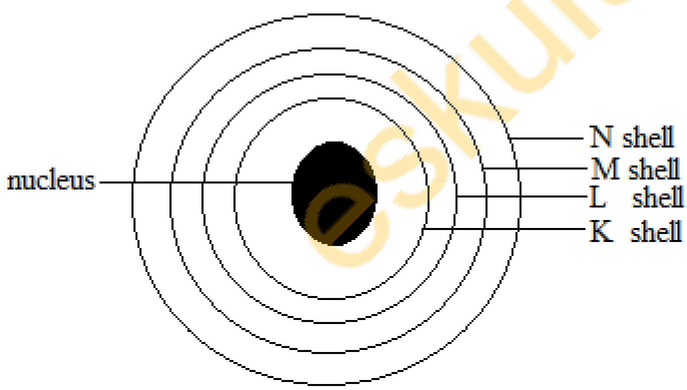
Electron shells

Alternative term: Energy levels.

Definition: Electron shells are concentric rings around the nucleus

The nucleus is also called the core of an atom.

The shells, in order of their increasing distance from the nucleus, are assigned by letters K, L, M, N etc.



Electron configuration

Alternative term: Electronic structure

Definition: Electron configuration is the arrangement of electrons in shells of atoms

The maximum number of electrons each shell can accommodate is given by the formula: $2n^2$, where n is the number of shells.

K shell, n = 1	$= 2n^2$ $= 2 \times 1^2$ $= 2 \times 1 \times 1$ = 2 electrons
L shell, n = 2	$= 2n^2$ $= 2 \times 2^2$ $= 2 \times 2 \times 2$ = 8 electrons
M shell, n = 3	$= 2n^2$ $= 2 \times 3^2$ $= 2 \times 3 \times 3$ = 18 electrons
N shell, n = 4	$= 2n^2$ $= 2 \times 4^2$ $= 2 \times 4 \times 4$ = 32 electrons

Metals have 1 or 2 or 3 electrons in the outer most shells

Non-metals have 4 or 5 or 6 or 7 electrons in the outer most shell

The outer most shell is the last shell

Ways of showing the electron configuration

[A] By writing

The symbol of an element is written first followed by the number of electrons in each shell separated by a dot.

Example

1. Show by writing the electron configuration in each of the following nuclides:

(a) $^{23}_{11}Na$ (b) $^{12}_{6}C$ (c) $^{20}_{10}Ne$ (d) $^{7}_{3}Li$ (e) $^{35.5}_{17}Cl$ (f) $^{39}_{19}K$

Solution

(a) Na 2. 8. 1
(b) C 2.4
(c) Ne 2.8
(d) Li 2.1
(e) Cl 2.8.7
(f) K 2.8.8.1

Exercise

1. Show by writing the electron configuration for each of the following nuclides:

(a) ${}^{1}_{1}H$ (b) ${}^{27}_{13}Al$ (c) ${}^{20}_{18}A_r$ (d) ${}^{32}_{16}S$

[B] By drawing

First show the electron configuration by writing, then draw the shells in form of rings or circles and indicate the number of electrons in each shell as crosses (x) or dots (·)

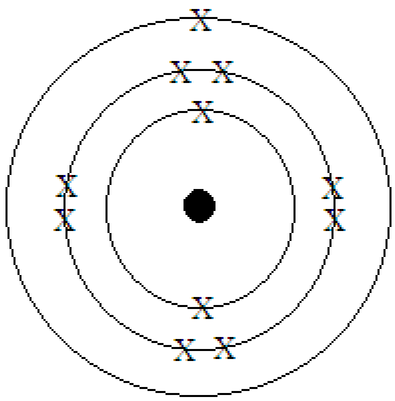
Examples

1. Show by drawing the electron configuration for each of the following nuclides:

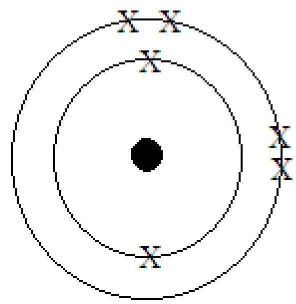
(a) ${}^{23}_{11}Na$ (b) ${}^{12}_{6}C$ (c) ${}^{20}_{10}Ne$ (d) ${}^{7}_{3}Li$

Solution

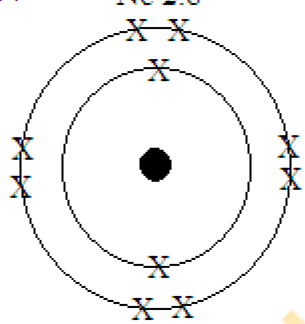
(a) Na 2.8.1



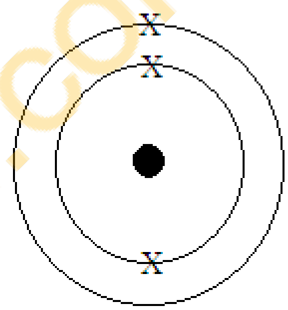
(b) C 2.4



(c) Ne 2.8



(d) Li 2.1



Electron configuration for the first 20 elements

Element	**Symbol**	**Atomic number**	**Electrons in shells** K .L.M. N
Hydrogen	H	1	1
Helium	He	2	2
Lithium	Li	3	2 . 1
Beryllium	Be	4	2 . 2
Boron	B	5	2 . 3
Carbon	C	6	2 . 4

Nitrogen	N	7	2 . 5
Oxygen	O	8	2 . 6
Fluorine	F	9	2 . 7
Neon	Ne	10	2 . 8
Sodium	Na	11	2 .8 . 1
Magnesium	Mg	12	2 .8 . 2
Aluminium	Al	13	2 .8 . 3
Silicon	Si	14	2 .8 . 4
Phosphorous	P	15	2 .8 . 5
Sulphur	S	16	2 .8 . 6
Chlorine	Cl	17	2 .8 . 7
Argon	Ar	18	2 .8 . 8
Potassium	K	19	2 .8 .8 . 1
Calcium	Ca	20	2 .8 .8 . 2

Ions

Definition: An ion is a charged particle

Types of ions

There are two types of ions; cations and anions

A. Cations

A cation is a positively charged particle

A cation has more protons than electrons

Formation of cations

Metals lose their outer most shell electrons to form cations

Cations can be divided into three categories

1. Monovalent cations

These are formed when a metal loses a single electron

Na 2. 8. 1 (neutral sodium atom)

$^{23}_{11}Na$	Protons	11
	Electrons	11
	Neutrons	12

Sodium atom loses a single electron to form sodium ion

$Na - e \rightarrow Na^{+}$

Na^{+}2 . 8 (sodium ion)

$^{23}_{11}Na^{+}$	Protons	11
	Electrons	10
	Neutrons	12

2. Divalent cations

These are formed when a metal loses two electrons

Mg 2. 8. 2 (neutral magnesium atom)

$^{24}_{12}Mg$	Protons	12
	Electrons	12
	Neutrons	12

Magnesium atom loses two electrons to form magnesium ion

$Mg - 2e \rightarrow Mg^{2+}$

Mg^{2+} 2. 8 (magnesium ion)

$^{24}_{12}Mg^{2+}$	Protons	12
	Electrons	10
	Neutrons	12

3. Trivalent cations

These are formed when a metal loses three electrons

Al 2. 8. 3 (neutral aluminium atom)

$^{27}_{13}Al$	Protons	13
	Electrons	13
	Neutrons	14

Aluminium loses three electrons to form aluminium ion

$Al - 3e \rightarrow Al^{3+}$

Al^{3+}2 . 8 (aluminium ion)

$^{27}_{13}Al^{3+}$	Protons	13
	Electrons	10
	Neutrons	14

B. Anions

An anion is a negatively charged particle

An anion has more electrons than protons

Formation of anions

Non-metals gain electrons to form anions

Anions can be divided into three categories

1. Monovalent anions

These are formed when a non-metal gains a single electron

F 2.7 (neutral fluorine atom)

${}^{19}_{9}F$	Protons	9
	Electrons	9
	Neutrons	10

Fluorine gains a single electron to form fluoride ion

$F + e \rightarrow F^{-}$

F^{-} 2. 8 (fluoride ion)

${}^{19}_{9}F^{-}$	Protons	9
	Electrons	9
	Neutrons	10

2. Divalent anions

These are formed when a non-metal gains two electrons

O 2 . 6 (neutral oxygen atom)

${}^{16}_{8}O$	Protons	8
	Electrons	8
	Neutrons	8

Oxygen atom gains two electrons to form oxide ion

$O + 2e \rightarrow O^{2-}$

O^{2-}2 . 8 (oxide ion)

${}^{16}_{8}O^{2-}$	Protons	8
	Electrons	10
	Neutrons	8

3. Trivalent anions

These are formed when a non-metal gain three electrons

N 2 . 5 (neutral nitrogen atom)

$^{14}_{7}N$	Protons	7
	Electrons	7
	Neutrons	7

Nitrogen atom gains three electrons to form nitride ion

$N + 3e \longrightarrow N^{3-}$

N^{3-} 2. 8 (nitride ion)

$^{14}_{7}N^{3-}$	Protons	7
	Electrons	10
	Neutrons	7

Example

1. Work out the number of protons, electrons and neutrons in each of the following:

   (a) $^{39}_{19}K$
   (b) $^{31}_{15}P^{3-}$
   (c) $^{40}_{20}Ca^{2+}$

Solution

	Protons	**Electrons**	**Neutrons**
$^{39}_{19}K$	19	19	20
$^{31}_{15}P^{3-}$	15	18	16

$^{40}_{20}Ca^{2+}$	20	18	20

Exercise

1. A lithium atom has a mass number of 7 and has 3 protons

   (a) Write the formula for the ion formed by lithium.
   (b) Explain how the lithium ion is formed.
   (c) Selecting from the following ions:

   Ca^{2+}, Cl^{-}, Li^{+}, Na^{+}, O^{2-}
   Choose the ion that;
   (I) has the same arrangement of electrons as Al^{3+}
   (II) has only two electrons
   (III) has the same arrangement of electrons as neon

2. Complete the table below

Ion	Number of protons	Number of electrons
F^{-}	9	
	11	10
Mg^{2+}		10

3. The table below shows the structure of several particles

Particle	**Electrons**	**Protons**	**Neutrons**
A	12	12	12
B	12	12	14
C	10	12	12
D	10	8	8

E	9	9	10

(a) Which three particles are neutral atoms?
(b) Which particle is a negative ion? What is the charge on this ion?
(c) Which particle is a positive ion? What is the charge on this ion?
(d) Which two particles are isotopes

Bonding of atoms

Definition: Bonding is chemical combination of two or more atoms.

Compounds and molecules result from chemical bonding

Only outer most shell electrons take part in bonding.

Atoms are held together by the forces of attraction or bonds.

A bond is a force of attraction between atoms

Why do atoms form bonds?

Atoms form bonds in order to be stable

Atoms react with one another in order to acquire full outer most shells like those of noble gases

Structure of noble gases

Noble gases are atoms that have eight electrons in the outer most shells except helium which has only two electrons.

Helium which has only two electrons obeys a duplex rule.

Other noble gases which have eight electrons in their outer most shells obey an octet rule.

Examples of noble gases

Noble gas	**Symbol**	**Electron configuration**
Helium	He	2
Neon	Ne	2. 8
Argon	Ar	2. 8. 8

Types of bonding

There are three types bonding

- Ionic bonding
- Covalent bonding
- Metallic bonding

Ionic bonding

Alternative term: Electrovalent bonding.

Ionic bonding involves the transfer of electrons from a metal to a non-metal

A metal loses electrons while a non-metal gains electrons.

Electrovalency

Electrovalency is the number of electrons lost or gained by an atom.

Ionic bond

Alternative term: Electrovalent bond

An ionic bond is the force of attraction between oppositely charged ions.

Ionic compounds

Ionic compounds consist of cations and anions.

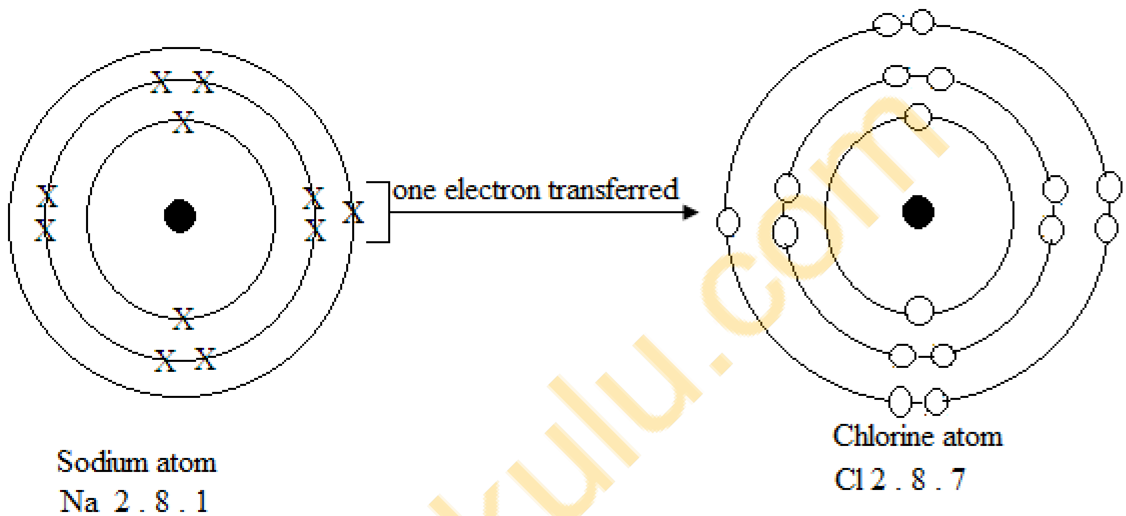
Examples of ionic compounds

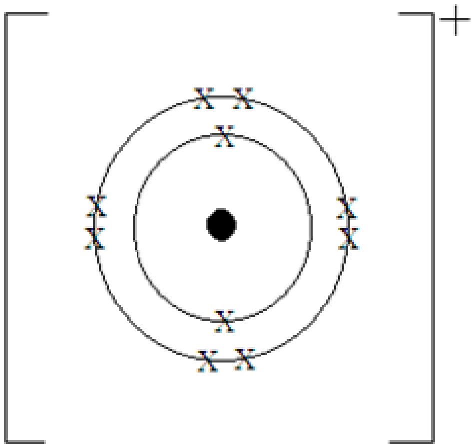
- Sodium chloride, $NaCl$
- Magnesium oxide, MgO
- Calcium chloride, $CaCl_2$

Formation of an ionic bond

Formation of sodium chloride

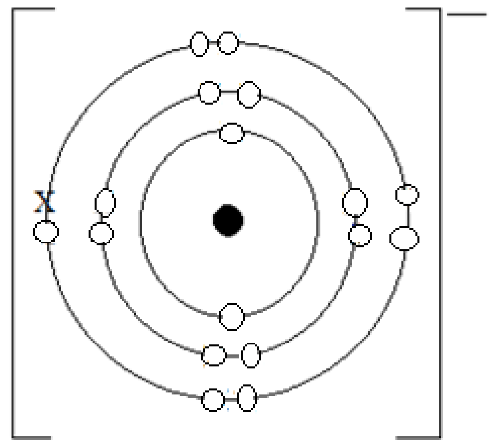
Electron configuration for sodium atom: Na 2.8.1	Electron configuration for chlorine atom: Cl 2.8.7
Sodium atom loses a single outer most shell electron to form sodium ion with a positive charge. $Na - e \rightarrow Na^{+}$ Na^{+} 2.8	Chlorine atom gains a single electron lost by sodium to form chloride ion with a negative charge. $Cl + e \rightarrow Cl^{-}$ Cl^{-} 2.8.8
The sodium and chloride ions attract each other to form a neutral compound called sodium chloride and the ionic bond is formed between the oppositely charged ions. $Na^{+} + Cl^{-} \rightarrow NaCl$	





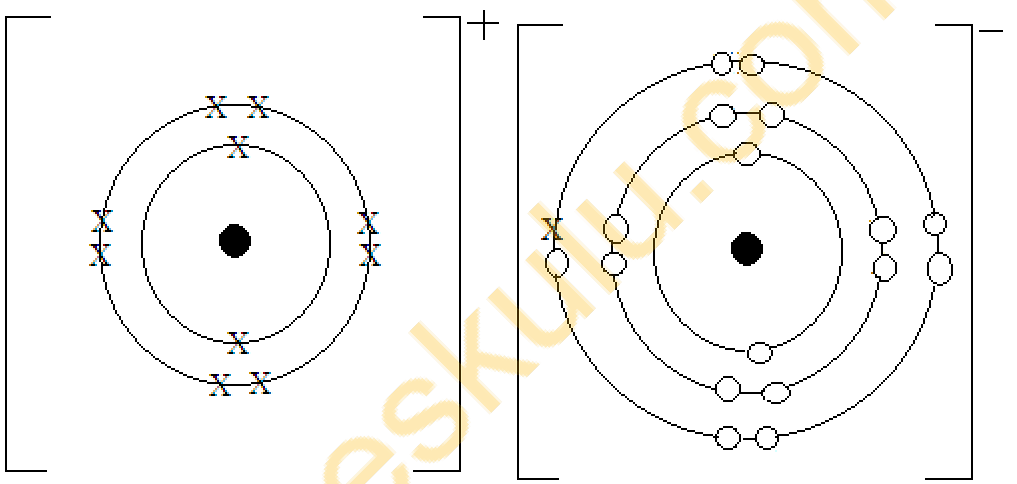
Sodium ion

Na^+ 2 . 8



Chloride ion

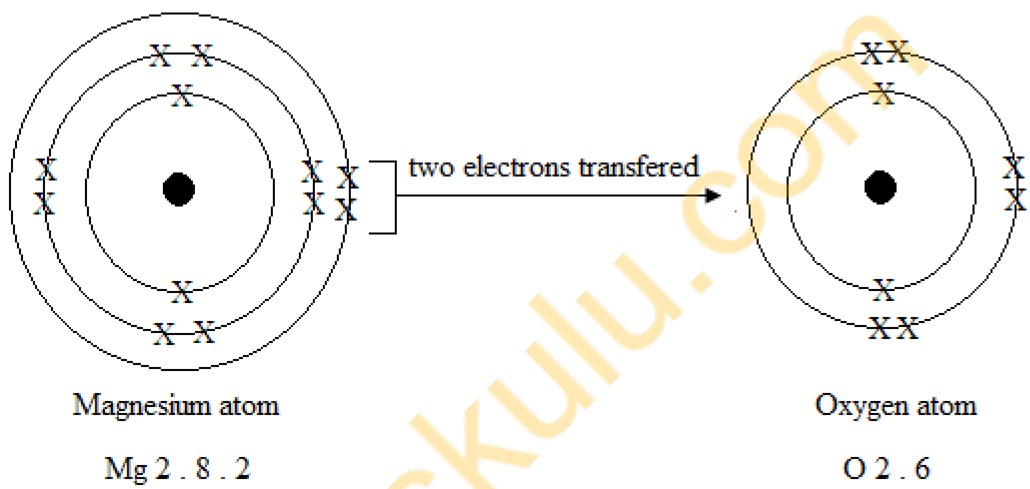
Cl^- 2 . 8 . 8

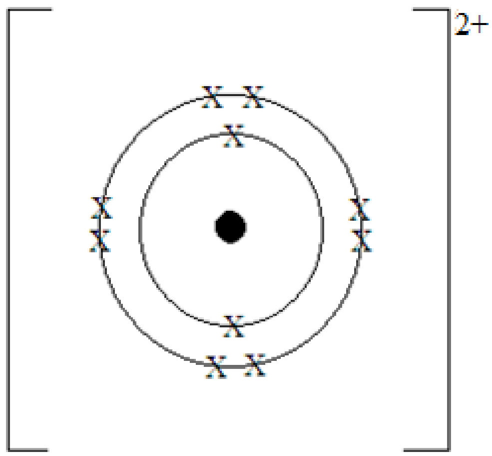


Sodium chloride, NaCl

Formation of magnesium oxide

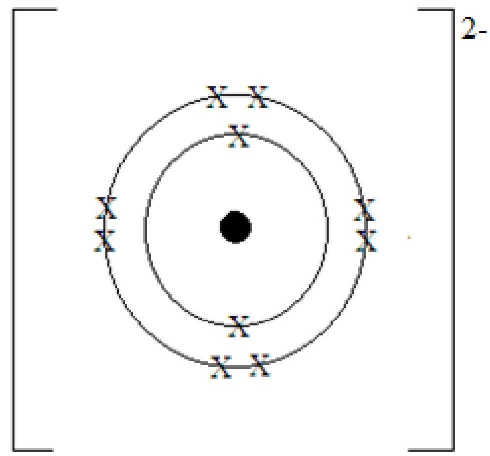
Electron configuration for magnesium atom: Mg 2.8.2	Electron configuration for oxygen atom: O 2.6
Magnesium atom loses its two outer most shell electrons to form magnesium ion with a charge of +2. $Mg - 2e \rightarrow Mg^{2+}$ Mg^{2+} 2.8	Oxygen atom gains two electron lost by magnesium to form oxide ion with a charge of −2. $O + 2e \rightarrow O^{2-}$ O^{2-} 2.8
The magnesium and oxide ions attract each other to form a neutral compound called magnesium oxide and the ionic bond is formed between the oppositely charged ions. $Mg^{2+} + O^{2-} \rightarrow MgO$	





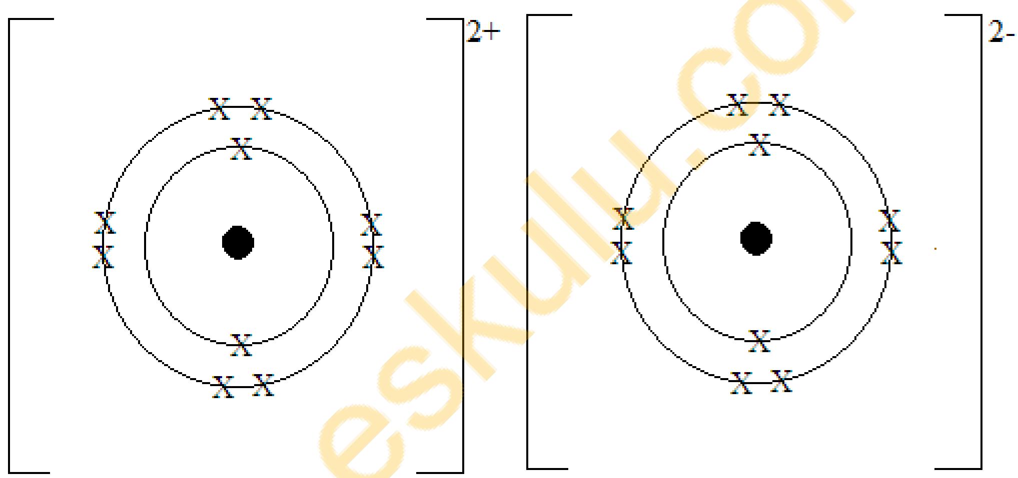
Magnesium ion, Mg^{2+}

Mg^{2+} 2 . 8



Oxide ion, O^{2-}

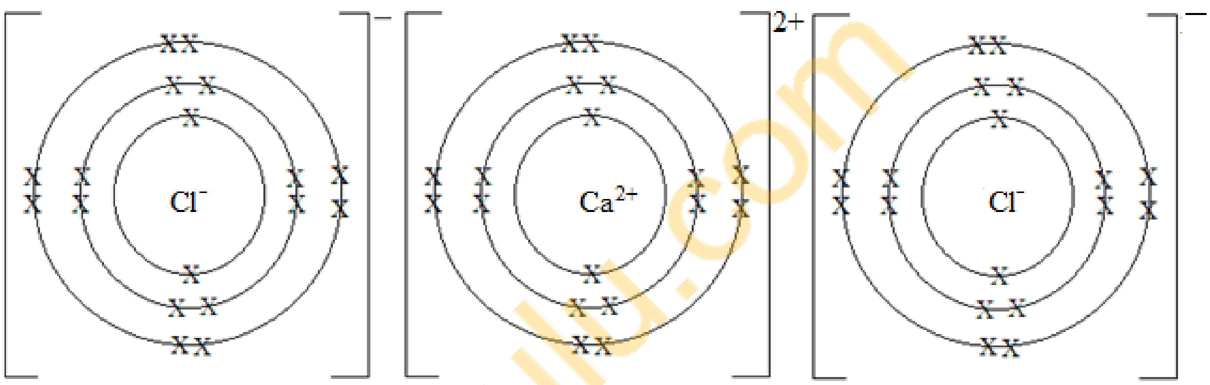
O^{2-} 2 . 8



Magnesium oxide, MgO

Formation of calcium chloride

<table>
<tr><td>Electron configuration for calcium atom: Ca 2.8.8.2</td><td>Electron configuration for chlorine atom: Cl 2.8.7</td></tr>
<tr><td>Calcium atom loses its two outer most shell electrons to form calcium ion with a charge of +2.

$Ca - 2e \rightarrow Ca^{2+}$
Ca^{2+} 2.8.8</td><td>The two electrons lost by calcium atom are gained by two atoms chlorine atom. Each chlorine atom gains only one electron to form a chloride ion.
$Cl + e \rightarrow Cl^{-}$
Cl^{-} 2.8.8
$2Cl + 2e \rightarrow 2Cl^{-}$</td></tr>
<tr><td colspan="2">The calcium and chloride ions attract each other to form a neutral compound called calcium chloride and the ionic bond is formed between the oppositely charged ions.
$$Ca^{2+} + 2Cl^{-} \rightarrow CaCl_2$$


Calcium chloride, $CaCl_2$</td></tr>
</table>

Characteristics of ionic compounds

1. They are made up of positively and negatively charged ions.
2. They have high melting and boiling points because of strong electrostatic forces of attraction between ions.
3. They are soluble in water but insoluble in organic solvents such as ethanol and petrol.
4. In aqueous solution or molten state, they conduct electricity because the ions are free to move.
5. They are non-volatile and generally solids at room temperature.

Example

1. Sodium chloride is an ionic solid.

(a) Show by writing the electronic structure of both a sodium ion and a chloride ion.

(b) Sodium chloride has a melting point of about 800 °C.

   (I) Explain why sodium chloride has a high melting point.

   (II) Magnesium oxide, MgO, has a similar structure to sodium chloride. Suggest why the melting point of magnesium oxide is higher than that of sodium chloride.

(c) Explain why solid sodium chloride will not conduct electricity but molten sodium chloride will.

Solution

(a) $Na^{+}2.8$

Cl^{-} 2.8.8

(b) (I) Because of the strong attraction between oppositely charged ions

(II) Because of the higher charges on the ions hence stronger attraction

(c) Ions cannot move in the solid but can move in the melt

Exercise

1. A metal X (atomic number 11) reacts with chlorine to form a white solid chloride Y.

   (a) Write down a balanced chemical equation for the reaction

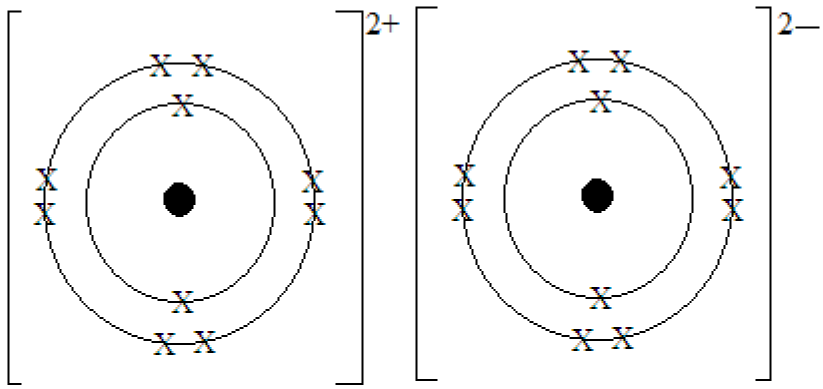
   (b) Show by drawing, the arrangement of electrons in X:

      (I) before the reaction

      (II) after the reaction

   (c) State three characteristics of Y.

2. The figure below shows the structure of a compound



(a) Name the compound shown
(b) Write down the chemical formula of the compound shown
(c) What type of bonding is present in the compound?
(d) State any two properties you would expect the compound to have

3. Draw a dot and cross diagram to show the bonding lithium fluoride, LiF

Covalent bonding

Alternative term: Molecular bonding

Covalent bonding involves the sharing of the outer most electrons between non-metal atoms.

Covalency

Covalency is the number of electrons an atom shares with another atom.

Covalent bond

Alternative term: Molecular bond

A covalent bond consists of a shared pair of electrons and it is formed between non-metals which share one or pairs of electrons.

Atoms are held by the attraction between their positive nuclei and the shared electrons.

Covalent compounds

Alternative term: Molecular compounds

Covalent compounds are usually molecules. A molecule the smallest particle of an element or compound which exists independently, that is in a free state.

They are formed when non-metals combine by sharing electrons.

As a result of sharing electrons, each non-metal acquires a completely filled outer most shell.

Examples of covalent compounds

- Hydrogen molecule, H_2
- Water, H_2O
- Carbon dioxide, CO_2
- Ammonia, NH_3

Formation of hydrogen molecule

Formula: H_2

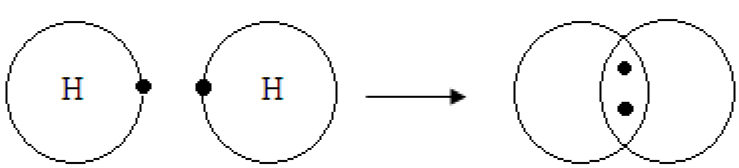
H – H

Each hydrogen atom has one electron.

When the two hydrogen atoms combine, they share electrons.

Note: only outer most shell electrons are shown.

$$H \quad + \quad H \quad \rightarrow \quad H_2$$

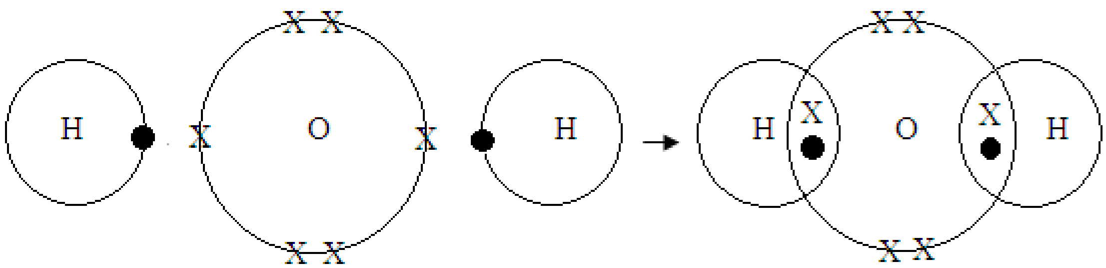


Formation of water molecule

Formula: H_2O

H – O – H

H + O + H → H_2O

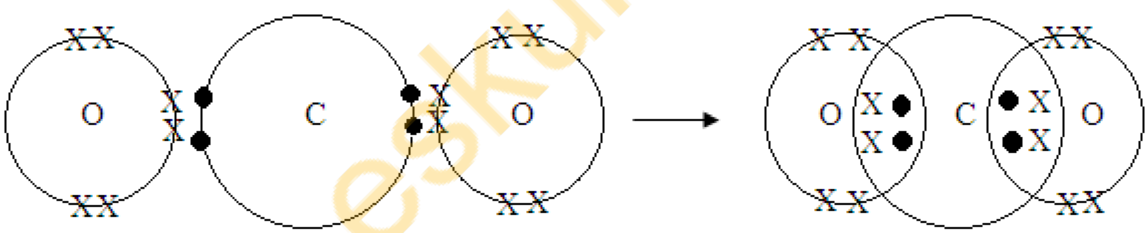


Formation of carbon dioxide

Formula: CO_2

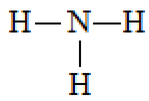
O = C = O or O:: C:: O

The two pairs of electrons are shared between a carbon atom and each oxygen atom and a double covalent bond is formed.

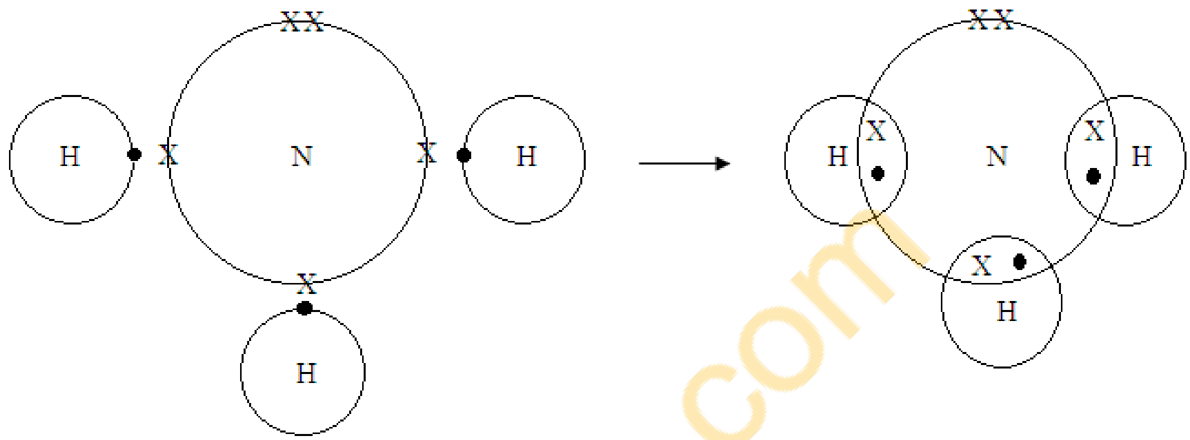


Formation of ammonia

Formula: NH_3



Ammonia is formed from bonding between one nitrogen atom and three hydrogen atoms.

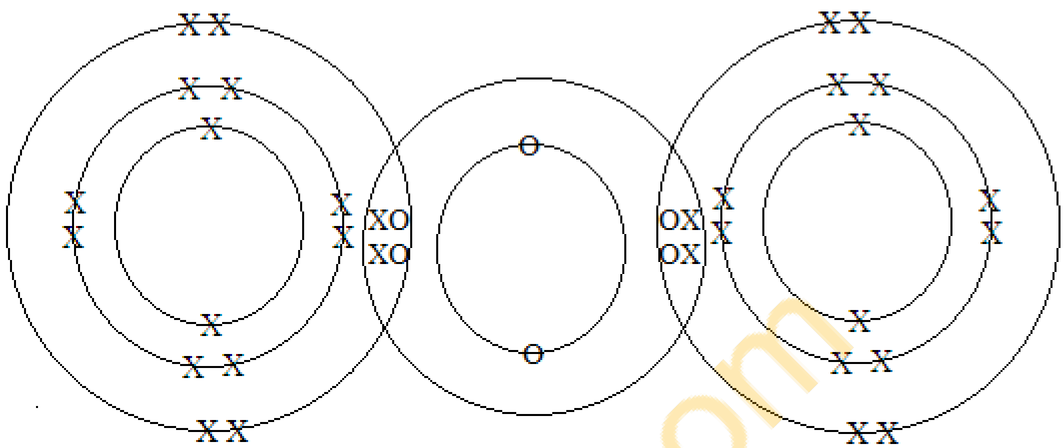


Characteristics of covalent compounds

1. They are made up of molecules
2. They have low melting and boiling points because of the weak forces of attraction which hold the molecules.
3. They are insoluble in water but soluble in organic solvents such as ethanol and petrol.
4. They do not conduct electricity in solid or solution form because they are made up of molecules.
5. They are generally volatile

Example

1. The diagram below shows the dot and cross structure of compound R showing all the shells and electrons



(a) Use the periodic table to deduce the chemical and structural formulae of compound R
(b) Compare the bonding in compound R to that in calcium chloride
(c) Explain the difference in electrical conductivity between compared R and calcium chloride in liquid form

Solution

(a) Chemical formula: CS_2

Structural formula: S = C = S

(b) The bonding in R involves the sharing of electrons between non – metals and in calcium chloride the bonding involves the transfer of electrons between a metal and a non- metal.
(c) Compound R will not conduct electricity in liquid form because it does not contain ions, so it is a non – electrolyte while calcium chloride will conduct electricity because it is made up of positively and negatively charged ions.

Exercise

2. With the aid of diagrams, show how covalent bonds are formed in the following molecules:

   (a) Oxygen, O_2
   (b) Ethane, CH_4
   (c) Hydrogen chloride, HCl
   (d) Chlorine, Cl_2

Metallic bonding

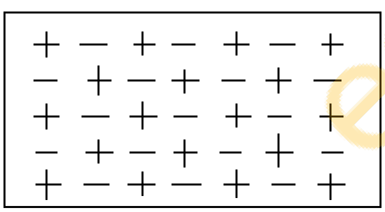
Metallic bonding is the attraction between the positively charged metal ions and the free electrons in a metallic lattice.

It involves the sea of electrons around positively charged particles inside a metal structure.

The electrons are free to move anywhere in the metallic lattice. The electrons are said to be delocalized.

Sodium atoms, for example, lose a single electron from the outer most shell. When a large number of sodium atoms lose these electrons, the result is many free electrons.

Sodium metal lattice



key

+ sodium ions
— electrons

Valency

Alternative term: Combining power

Definition: Valency is the number of electrons lost or gained or shared by an atom of the element to attain a stable structure.

Valencies of some elements

Valency	Metals		Non metals	
	Name	**Symbol**	**Name**	**Symbol**
1	Potassium Silver Sodium Copper (I)	K Ag Na Cu	Chlorine Hydrogen	Cl H
2	Barium Calcium Copper (II) Iron (II) Lead (II) Magnesium Mercury Zinc	Ba Ca Cu Fe Pb Mg Hg Zn	Oxygen Sulphur	O S
3	Aluminium Iron (III)	Al Fe	Nitrogen Phosphorus (III)	N P
4	Lead (IV)	Pb		
5			Phosphorus (V)	P

Radical

A radical is a group of atoms which is present in several compounds but incapable of independent existence.

Radicals and their Valencies

Radical	Formula	Valency
Ammonium	NH_4	
Chlorate	ClO_3	
Chloride	Cl	
Hydrogen carbonate	HCO_3	1
Hydrogen sulphate	HSO_4	
Hydroxide	OH	
Nitrate	NO_3	
Nitrite	NO_2	
Carbonate	CO_3	
Oxide	O	
Sulphate	SO_4	2
Sulphide	S	
Sulphite	SO_3	
Phosphate	PO_4	3

Chemical formula

A chemical formula consist of a symbol or symbols showing the number of atoms in one molecule of an element or a compound

Writing chemical formula

Steps to consider when writing the chemical formula

1. Write the symbol for the combining elements and radicals
2. Write the Valency of each element or radical at the top at its top right hand side
3. Exchange the Valencies of the combining elements and radicals by writing them at the bottom right hand side of the element or radical. If the number is 1, do not write it.

   In some formulae, radicals are written in brackets followed by a small sub script digit.

Examples

1. Write the formula for each of the following compounds
   (a) Sodium chloride
   (b) Potassium carbonate
   (c) Ammonium carbonate
   (d) Aluminium phosphate
   (e) Calcium hydrogen carbonate

Solution

(a) Na Cl

Na^1Cl^1

NaCl

(b) K CO_3

$K^1CO_3^2$

$\mathbf{K_2CO_3}$

(c) NH_4 CO_3

NH_4^1 CO_3^2

$\mathbf{(NH_4)_2CO_3}$

(d) Al PO_4

Al^3 PO_4^3

$\mathbf{AlPO_4}$

(e) Ca HCO_3

Ca^2 HCO_3^1

$\mathbf{Ca(HCO_3)_2}$

2. Write the Valency for the elements and radical for each of the following:
   (a) $CaCO_3$
   (b) Fe_2O_3
   (c) NH_3

Solution

	Element /radical	Valency
(a)	Ca	2
	CO_3	2
(b)	Fe	3
	O	2
(c)	N	3
	H	1

Exercise

1. Write the Valency for the elements and radicals in each of the following:
   (a) $Al_2(SO_4)_2$
   (b) $(NH_4)_2CO_3$
   (c) $NaHSO_4$
2. Calcium nitrate has the ions Ca^{2+} and NO_3^-. Write the formula of the compound formed when the two ions combine.

State symbols

State symbols are letters that are used to show the physical state of substances in the equation

State symbols are placed in brackets after the name or formula of each substance in the equation

Examples

Physical state	State symbol
Gas or vapour	(g)
Liquid or molten	(l)
Solid or precipitate	(s)
Aqueous solution	(aq)

Equations

An equation is a chemical sentence which describes what is happening in a chemical reaction

An equation can be represented in the form:

A + B → C + D

(Reactants) (Products)

Interpretation: A reacts with B to form C and D.

A→ B + C

Interpretation: A decomposes into B and C.

The plus (+) sign on the left hand side means "react with"

The plus (+) sign on the right hand side means "and"

The arrow (→) between the reactants and products means "to form"

Reactants are substances that take part in a chemical reaction. They are normally written at the left hand side of a chemical equation

Products are new substances that are formed from the reaction. They are normally written at the right hand side of a chemical equation

[A] Word equations

Word equations are chemical equations written in words

Examples

1. Write down the word equation for each of the following reactions including state symbols
   (a) Magnesium metal reacts with oxygen gas to form magnesium oxide
   (b) Hydrogen sulphide gas reacts with oxygen gas to form sulphur and water
   (c) Iron (II) chloride solution and hydrogen gas are produced when iron reacts with dilute hydrochloric acid
   (d) Iron reacts with chlorine gas to form iron (II) chloride

Solution

(a) $\text{Magnesium}_{(s)} + \text{Oxygen}_{(g)} \rightarrow \text{Magnesium oxide}_{(s)}$

(b) $\text{Hydrogen sulphide}_{(s)} + \text{Oxygen}_{(g)} \rightarrow \text{Sulphur}_{(s)} + \text{Water}_{(l)}$

(c) $\text{Iron}_{(s)} + \text{Hydrochloric acid}_{(aq)} \rightarrow \text{Iron (II) chloride}_{(aq)} + \text{Hydrogen}_{(g)}$

(d) $\text{Iron}_{(s)} + \text{Chlorine}_{(g)} \rightarrow \text{Iron (II) chloride}_{(s)}$

Exercise

1. Write down word equations including state symbols for each of the reactions:
   (a) Mercury oxide decomposes into mercury and oxygen
   (b) Hydrogen gas reacts with oxygen gas to form water
   (c) Sodium metal reacts with water to produce a solution of sodium hydroxide and hydrogen gas
   (d) Calcium oxide dissolves in water to produce calcium hydroxide solution.

[B] Equations with symbols

Writing balanced chemical equations

Balancing the equation is the process of making the number of each type of atom equal on both sides of the equation

Never change the chemical formula of compounds when balancing the chemical equations. You can only add numbers in front of the chemical formula.

Examples

1. Write down balanced chemical equations including state symbols for each of the following word equations
   (a) Mercury oxide$_{(s)}$ $\rightarrow$ Mercury$_{(l)}$ + Oxygen$_{(g)}$
   (b) Hydrogen$_{(g)}$ + Oxygen$_{(g)}$ $\rightarrow$ Water$_{(l)}$
   (c) Magnesium$_{(s)}$ + Oxygen$_{(g)}$ $\rightarrow$ Magnesium oxide$_{(s)}$
   (d) Sodium$_{(s)}$ + Water$_{(l)}$ $\rightarrow$ Sodium hydroxide$_{(aq)}$ + Hydrogen$_{(g)}$
   (e) Calcium oxide$_{(s)}$ + Hydrochloric acid$_{(aq)}$ $\rightarrow$ Calcium chloride$_{(aq)}$ + Water$_{(l)}$

Solution

(a) $2HgO_{(s)} \rightarrow 2Hg_{(l)} + O_{2(g)}$

(b) $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$

(c) $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

(d) $2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$

(e) $CaO_{(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)}$

Exercise

1. Balance the following equations

   (a) $H_{2(g)} + Cl_{2(g)} \rightarrow HCl_{(g)}$

   (b) $NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

   (c) $P_{(s)} + O_{2(g)} \rightarrow P_2O_{5(s)}$

   (d) $H_2S_{(g)} + O_{2(g)} \rightarrow H_2O_{(l)} + SO_{2(g)}$

   (e) $CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$

   (f) $Fe_{(s)} + Cl_{2(g)} \rightarrow FeCl_{3(s)}$

   (g) $KMnO_{4(s)} + HCl_{(aq)} \rightarrow KCl_{(aq)} + MnCl_{2(aq)} + H_2O_{(l)} + Cl_{2(g)}$

[C] Ionic equations

Ionic equations show only the ions involved in a chemical reaction

Ions not taking part in the reaction (spectator ions) are cancelled out in the construction of ionic equations

Steps to consider when writing the ionic equation

1. Construct a balanced chemical equation
2. Split only soluble ionic compounds (compounds in aqueous state) into ions . Insoluble ionic compounds, elements and covalent compounds remain unchanged.
3. Cancel out spectator ions. These are ions that appear on both the left and right hand side of the equation.
4. Rewrite the equation without the spectator ions

Ion	Formula of ion	Valency
Ammonium ion	NH_4^+	1
Chloride ion	Cl^-	1
Hydrogen ion	H^+	1
Hydroxide ion	OH^-	1
Nitrate ion	NO_3^-	1
Potassium ion	K^+	1
Silver ion	Ag^+	1
Sodium ion	Na^+	1

Ion	Formula of ion	Valency
Aluminium ion	Al^{3+}	3
Carbonate ion	CO_3^{2-}	2
Lead (II) ion	Pb^{2+}	2
Phosphate ion	PO_4^{3-}	3
Sulphate ion	SO_4^{2-}	2
Barium ion	Ba^{2+}	2
Copper (II) ion	Cu^{2+}	2
Calcium ion	Ca^{2+}	2

Example

1. Write the ionic equations for the reactions below

(a) $Ba(NO_3)_{2(aq)} + Na_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + 2NaNO_{3(aq)}$

(b) $CuCO_{3(s)} + 2HCl_{(aq)} \rightarrow CuCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

(c) $NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

Solution

(a) $Ba(NO_3)_{2(aq)} + Na_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + 2NaNO_{3(aq)}$

$Ba^{2+}{}_{(aq)} + \cancel{2NO_3^-}{}_{(aq)} + \cancel{2Na^+}{}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)} + \cancel{2Na^+}{}_{(aq)} + \cancel{2NO_3^-}{}_{(aq)}$

$Ba^{2+}{}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)}$

(b) $CuCO_{3(s)} + 2HCl_{(aq)} \rightarrow CuCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

$CuCO_{3(s)} + 2H^+{}_{(aq)} + \cancel{2Cl^-}{}_{(aq)} \rightarrow Cu^{2+}{}_{(aq)} + \cancel{2Cl^-}{}_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

$CuCO_{3(s)} + 2H^+{}_{(aq)} \rightarrow Cu^{2+}{}_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

(c) $NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

$\cancel{Na^+}{}_{(aq)} + OH^-{}_{(aq)} + H^+{}_{(aq)} + \cancel{Cl^-}{}_{(aq)} \rightarrow \cancel{Na^+}{}_{(aq)} + \cancel{Cl^-}{}_{(aq)} + H_2O_{(l)}$

$OH^-{}_{(aq)} + H^+{}_{(aq)} \rightarrow H_2O_{(l)}$

Exercise

1. Write the ionic equations for the following reactions

(a) $AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$

(b) $Ca(OH)_{2(aq)} + H_2SO_{4(aq)} \rightarrow CaSO_{4(aq)} + H_2O_{(l)}$

Determining the number of atoms in a compound

Examples

1. What is the total number of atoms present in a molecule of lead (II) nitrate, $Pb(NO_3)_2$?

Data	Solution
Pb = 1 N = 2 O = 3 x 2 = 6	Total number = 1 +2 + 6 = 9 atoms

2. Find the total number of atoms in ammonium phosphate, $(NH_4)_3PO_4$

Data	Solution
N = 3 H = 4 x 3 = 12 P = 1 O = 4	Total number = 3 + 12 + 1 + 4 = 20 atoms

3. Find the total number of atoms in three moles of ammonium sulphate, $3(NH_4)_2SO_4$

Data	Solution
N = 3 x 2 = 6 H = 3 x 4 x 2 = 24 S = 3 O = 3 x 4 = 12	Total number = 6 + 24 + 3 + 12 = 45 atoms

4. How many different atoms are in sodium ethanoate, CH_3COONa?

Solution

4 different atoms i.e. C, H, O and Na.

Exercise

1. Calculate the total number of atoms in calcium hydrogen sulphate, $Ca(HSO_4)_2$
2. How many different atoms are in urea?

Stoichiometric calculations

Relative atomic mass

Symbol: A_r

Units: It has no units

Definition: Relative atomic mass of an element is the mass of one atom of an element compared to $\frac{1}{12}$ the mass of carbon $-$ 12 isotope.

Relative atomic masses of some elements

Element	Symbol	A_r
Hydrogen	H	1
Carbon	C	12
Nitrogen	N	14
Oxygen	O	16
Sodium	Na	23
Magnesium	Mg	24

Element	Symbol	A_r
Aluminium	Al	27
Sulphur	S	32
Chlorine	Cl	35.5
Calcium	Ca	40
Iron	Fe	56
Copper	Cu	64

Relative molecular mass

Symbol: M_r

Units: It has no units

Definition: Relative molecular mass of a compound is the mass of one molecule of the compound or element compared with $\frac{1}{12}$ the mass of carbon $-$ 12 isotope.

Relative molecular mass can also be defined as the sum of relative atomic masses

Examples

1. Find the relative molecular, M_r, of the following:
   (a) Hydrogen chloride, HCl
   (b) Carbon dioxide, CO_2
   (c) Sodium sulphate, Na_2SO_4
   (d) Copper (II) sulphate -5-water, $CuSO_4.5H_2O$
   (e) Three moles of ammonium carbonate, $3(NH_4)_2CO_3$

Solution

(a) M_r of HCl = [(1 x 1) + (1 x 35.5)]

= 1 + 35.5

= 36.5

(b) M_r of CO_2 = [(1 x 12) + (2 x 16)]

= 12 + 32

= 44

(c) M_r of Na_2SO_4 = [(2 x 23) + (1 x 32) + (4 x 16)]

= 46 + 32 + 64

= 142

(d) M_r of $CuSO_4.5H_2O$ = [(1 x 64) + (1 x 32) + (4 x 16) + (10 x 1) + (5 x 16)]

= 64 + 32 + 64 + 10 + 80

= 250

(e) M_r of $3(NH_4)_2CO_3$ = [(6 x 14) + (24 x 1) + (3 x 12) + (9 x 16)]

= 84 + 24 + 36 + 144

= 288

2. It is known that one form of carbon has the molecules of formulaC_{60}. What is the relative molecular mass, M_r, of these molecules?

Solution

The formula C_{60} means that the molecule contains 60 carbon atoms bonded together.

$M_r$of C_{60} = 60 x 12

= 720

Exercise

1. Work out the relative molecular mass of the following:

(a) $2H_2$

(b) O_2

(c) $2H_2O$

(d) $Ca(HCO_3)_2$

Molar mass

Symbol: MM

SI unit: gram per mole, g/mol

Definition: Molar mass is the mass of one mole of a substance

Molar mass can also be defined as relative molecular mass expressed in grams per mole

Examples

	Symbol / Formula	Relative molecular mass	Molar mass
One mole of potassium	K	(1x39) = 39	39g/mol
One mole of sodium	Na	(1x23) = 23	23g/mol
One mole of ammonia	NH_3	[(1x14) + (3x1)] =17	17g/mol
One mole of carbon dioxide	CO_2	[(1x12) + (2x16)] = 44	44g/mol

Mole

Symbol: n

Unit: mole, mol

Definition: The mole is the amount of substance which contains as many elementary entities as they are in 12.00g of carbon-12 isotope

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

$n = \frac{m}{MM}$

n = mole [mol]

m = mass [g]

MM = molar mass [g/mol]

Examples

1. How many moles of potassium are there in 3.9g of potassium?

Data	Solution
n =? m = 3.9g MM = 39g/mol	$n = \frac{m}{MM}$ $n = \frac{3.9g}{39g/mol}$ n = 0.1mol

2. Find the mass of 0.2moles of ammonia molecules, NH_3.

Data	Solution
m =? n = 0.2mol MM = 17g/mol	m = n x MM m = 0.2mol x 17g/mol m = 3.4g

3. Calculate the molar mass of 2.5moles of a substance X weighing 100g.

Data	Solution
MM =? n = 2.5mol m = 100g	$MM = \frac{m}{n}$ $MM = \frac{100g}{2.5mol}$ MM = 40g/mol

4. How many moles of hydrogen atoms does 3.2g of methane, CH_4, contain?

Solution

MM for CH_4 = [(1 x 12) + (4 x 1)] = 16g/mol

$n(CH_4) = \frac{m}{MM}$

$n = \frac{3.2g}{16g/mol}$

n = 0.2mol

Number of moles of hydrogen in CH_4 = number of hydrogen atoms in CH_4 x *n*

= 4 x 0.2mol

= 0.8mol

Exercise

1. Calculate the number of moles in 20g of calcium carbonate, $CaCO_3$
2. If you need 2.5 moles of sodium hydroxide, NaOH, what mass of sodium hydroxide do have to weigh?
3. Calculate the molar mass of 0.5 moles of substance X weighing 28g.
4. How many moles of oxygen molecules are there in 64g of oxygen, O_2?

Avogadro's number

Alternative term: Avogadro's constant

Symbol: N_A

$N_A = 6.02 \times 10^{23}$ moles

Definition: Avogadro's number is the number of particles in exactly one mole of a pure substance

12.00g of carbon contains as many as 6.02×10^{23} atoms

1 mole of any element has mass equivalent to its mass number and contains 6.02×10^{23} particles.

Example: 1 mole of Ca = 40g = 6.02×10^{23} particles

1 mole of Fe = 56g = 6.02×10^{23} particles

1 mole of C = 12g = 6.02×10^{23} particles

1 mole of Mg = 24g = 6.02×10^{23} particles

1 mole of N = 14g = 6.02×10^{23} particles

1 mole of N_2 = 28g = 6.02×10^{23} particles

1 mole of H = 1g = 6.02×10^{23} particles

1 mole of H_2= 2g = 6.02×10^{23} particles

1 mole of Na = 23g = 6.02×10^{23} particles

Elementary entities include atoms, molecules, ions, electrons, protons and neutrons

1 mole of atoms = $6.02 \text{ x } 10^{23}$ atoms

1 mole of molecules = $6.02 \text{ x } 10^{23}$ molecules

1 mole of ions = $6.02 \text{ x } 10^{23}$ ions

1 mole of electrons = $6.02 \text{ x } 10^{23}$ electrons

1 mole of protons = $6.02 \text{ x } 10^{23}$ protons

1 mole of neutrons = $6.02 \text{ x } 10^{23}$ neutrons

Examples

1. How many atoms of iron (Fe) are there in 25g of iron?

Solution

56g $\rightarrow$ $6.02 \text{ x } 10^{23}$ atoms

25g $\rightarrow$ x

$$x = \frac{25\text{g x } 6.02 \text{ x } 10^{23} \text{ atoms}}{56\text{g}}$$

x = $2.89 \text{ x } 10^{23}$ atoms

2. If 1 mole of carbon is 12g, calculate the mass of 1 atom of carbon 12

Solution

$6.02 \text{ x } 10^{23}$ atoms $\rightarrow$ 12g

1 atom $\rightarrow$ x

$$x = \frac{12\text{g x 1 atom}}{6.02 \text{ x } 10^{23} \text{ atoms}}$$

x = $1.99 \text{ x } 10^{-23}$ g

3. What is the mass of 1.2 x 10^{23} atoms of calcium?

Solution

6.02 x 10^{23} atoms $\rightarrow$ 40g

1.2 x 10^{23}atoms $\rightarrow$ x

x = $\frac{1.2 \text{ x } 10^{23} \text{ atoms x } 40\text{g}}{6.02 \text{ x } 10^{23} \text{ atoms}}$

x = 7.97g

4. How many hydrogen atoms are there in 1 mole of hydrogen molecules, H_2?

Solution

1 mole of atoms $\rightarrow$ 6.02 x 10^{23} atoms

1 mole of H_2 molecules $\rightarrow$ x

But

1 mole of H_2 molecules $\rightarrow$ 2 atoms

6.02 x 10^{23} molecules $\rightarrow$ x

x = $\frac{6.02 \text{ x } 10^{23} \text{ molecules x 2 atoms}}{1 \text{ mol of } H_2 \text{ molecules}}$

x = 1.204 x 10^{24} atoms

5. How many oxygen atoms are in 1.6g of sulphur trioxide, SO_3?

Solution

$M_r$of oxygen = 3 x 16 = 48

$M_r$ofSO_3 = [(1 x 32) + (3 x 16)] = 80

n (SO_3) = $\frac{m}{MM}$

n = $\frac{1.6g}{80g/mol}$

n = 0.02mol

Number of oxygen atoms in 1.6g of SO_3= number of oxygen atoms in SO_3 x n x N_A

$= 3 \times 0.02\text{mol} \times 6.02 \times 10^{23}$ atoms

$= 3.612 \times 10^{22}$ atoms

Exercise

1. How many atoms of magnesium are there in 5g of magnesium?
2. What is the mass of 3.01×10^{23} atoms of carbon?
3. How many nitrogen atoms are there in 7g of nitrogen molecules, N_2?

Calculations from equations

Mole to mole calculations

Examples

1. How many moles of chlorine are required to react 2.5 moles of calcium to produce calcium chloride?

Solution

$Ca_{(s)} + Cl_{2(g)} \rightarrow CaCl_{2(s)}$

1 mol Ca $\rightarrow$ 1 mol Cl_2

2.5mol Ca $\rightarrow$ x

$x = \frac{2.5 \text{ mol Ca} \times 1 \text{ mol } Cl_2}{1 \text{ mol Ca}}$

x = 2.5 mol

2. How many moles of carbon dioxide (CO_2) will be produced by complete combustion of 2 moles of glucose ($C_6H_{12}O_6$) according to the equation?

$C_6H_{12}O_{6(aq)} + 6O_{2(g)} \rightarrow 6CO_{2(g)} + 6H_2O_{(l)}$

Solution

1 mol $C_6H_{12}O_6$ $\rightarrow$ 6 mol CO_2

2 mol $C_6H_{12}O_6$ $\rightarrow$ x

$x = \frac{2 \text{ mol } C_6H_{12}O_6 \times 6 \text{ mol } CO_2}{1 \text{ mol } C_6H_{12}O_6}$

x = 12 mol CO_2

Exercise

1. How many moles of ammonia can be produced from 8 moles of nitrogen gas?

$N_2 \quad + \quad 3H_2 \quad \rightarrow \quad 2NH_3$

2. How many moles of oxygen are needed to burn 1.8 moles of ethanol, C_2H_5OH?

$C_2H_5OH \quad + \quad 3O_2 \quad \rightarrow \quad 2CO_2 \quad + \quad 3H_2O$

Mole to mass calculations

Examples

1. What mass of hydrogen can be produced by reacting 6 moles of aluminium with hydrochloric acid?

$2Al_{(s)} \quad + \quad 6HCl_{(aq)} \rightarrow 2AlCl_{3(aq)} \quad + \quad 3H_{2(g)}$

Solution

2 mol Al $\rightarrow$ 6g H_2

6 mol Al $\rightarrow$ x

$x = \frac{6 \text{ mol Al x 6g } H_2}{2 \text{ mol Al}}$

x = 18g H_2

2. How many grams of oxygen are required to react with 0.3 moles of aluminium to produce aluminium oxide?

$4Al_{(s)} \quad + \quad 3O_{2(g)} \quad \rightarrow \quad 2Al_2O_{3(s)}$

Solution

4 mol Al $\rightarrow$ 96g O_2

0.3 mol Al $\rightarrow$ x

$x = \frac{0.3 \text{ mol Al x } 96\text{g } O_2}{4 \text{ mol Al}}$

x = 7.2g O_2

Exercise

1. What mass of magnesium oxide will be produced when 0.5 moles of magnesium burns in oxygen

Mass to mass calculation

Examples

1. Calculate the mass of calcium chloride produced when 40g of calcium carbonate reacts with hydrochloric acid

   $CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

Solution

$100g\ CaCO_3 \rightarrow 111g\ CaCl_2$

$40g\ CaCO_3 \rightarrow x$

$x = \frac{40g\ CaCO_3 \times 111g\ CaCl_2}{100g\ CaCO_3}$

$x = 44.4g\ CaCl_2$

2. Consider the reaction below.

   $UF_4 + 2Mg \rightarrow 2MgF_2 + U$

How many tonnes of uranium can be produced in the above reaction using 24 tonnes of magnesium?

Solution

$48g\ Mg \rightarrow 238g\ U$

$24tonnes\ Mg \rightarrow x$

$x = \frac{24\ tonnes\ Mg \times 238g\ U}{48g\ Mg}$

$x = 119\ tonnes\ U$

Exercise

1. What mass of calcium metal reacts with 9.0g of water according to the equation below?
   $Ca_{(s)} + 2H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$
2. Hydrogen burns in oxygen to form water. The equation for the reaction is:
   $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$
   How much oxygen is needed to burn 1g of hydrogen?
3. Consider the reaction: $CaCO_3 \rightarrow CaO + CO_2$. What mass of lime (CaO) would be produced from 20 tonnes of limestone ($CaCO_3$)?

Limiting reagents

Definition: A limiting reagent is a reactant that is in short supply by the mole ratio and hence it finishes before the other reactants are completely reacted.

A limiting reagent is always smaller or in less quantity compared to the other reactant

A limiting reagent is found by dividing the number of moles of each reactant by its stoichiometric coefficient in a balanced chemical equation.

$$aA \quad + \quad bB \quad \rightarrow \quad cC \quad + \quad dD$$

$R_A = \frac{n_{(A)}}{a}$, $R_B = \frac{n_{(B)}}{b}$

R_A = Reagent A, R_B = Reagent B

The ratio that is smaller is that of a limiting reagent. A limiting reactant has less number of moles compared to the other reactant

A limiting reagent determines the extent to which the chemical reaction can proceed and the amount of products that would be formed. Once the limiting reagent is finished, the reaction stops even if the other reactants are still available in the reaction vessel

For this reason, it is important to identify the limiting reagent before calculating the theoretical yield

In the identification of the limiting reagent,

- The balanced chemical equation and the mole ratio of reactants are used
- You cannot use volumes, concentrations or masses of the reactants since these will easily mislead you.

Note

Excess reagents

An excess reagent is a reactant that remains unreacted at the end of the reaction.

Example

1. 19.5g of zinc and 9.40g of sulphur were heated together

   $Zn_{(s)} \quad + \quad S_{(s)} \quad \rightarrow \quad ZnS_{(s)}$

(a) Which of the two is the limiting reactant?

(b) How many moles of zinc remain unreacted?

(c) How many grams of zinc element remain unreacted?

(d) Calculate the mass of zinc sulphide formed

Solution

(a) $n(Zn) = \frac{m}{MM}$ $\qquad$ $n(S) = \frac{m}{MM}$

$= \frac{19.5g}{65g/mol}$ $= \frac{9.40g}{32g/mol}$

= 0.3mol = 0.29 mol

The limiting reagent: Sulphur

Reason: It has a smaller number of moles compared to zinc.

(b) Number of moles of Zn unreacted = 0.3mol – 0.29mol

= 0.01 mol

(c) m = n x MM

m = 0.01mol x 65g/mol

m = 0.65g of Zn remained unreacted

(d) 32g S $\rightarrow$ 97g ZnS

9.4g S $\rightarrow$ x

$x = \frac{9.40g\ S \times 97g\ ZnS}{32g\ S}$

x = 28.5g ZnS

2. How many grams of hydrogen chloride would be produced from 0.49g of hydrogen and 50g of chlorine?

$H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$

Solution

$n(H_2) = \frac{m}{MM}$ $n(Cl_2) = \frac{m}{MM}$

$= \frac{0.49g}{2.0g/mol}$ $= \frac{50g}{71g/mol}$

= 0.245mol = 0.704 mol

The limiting reagent: Hydrogen

Reason: It has a smaller number of moles compared to chlorine.

$2.0g\ H_2 \rightarrow 73g\ HCl$

$0.49g\ H_2 \rightarrow x$

$x = \frac{0.49g\ H_2\ \times 73g\ HCl}{2.0g\ H_2}$

x = 17.885g HCl

3. How many moles of iron trioxide (Fe_3O_4) can be obtained by reacting 16.8g of iron with 10g of steam?

$3Fe_{(s)} + 4H_2O_{(g)} \rightarrow Fe_3O_{4(s)} + 4H_{2(g)}$

Solution

$n(Fe) = \frac{m}{MM}$ $\quad\quad$ $n(H_2O) = \frac{m}{MM}$

$n = \frac{16.8g}{56g/mol}$ $\quad\quad$ $n = \frac{10g}{18g/mol}$

$n = 0.3mol$ $\quad\quad$ $n = 0.56\ mol$

$R = \frac{n(Fe)}{3}$ $\quad\quad$ $R = \frac{n(H_2O)}{4}$

$n = \frac{0.3mol}{3}$ $\quad\quad$ $n = \frac{0.56mol}{4}$

$n = 0.1mol$ $\quad\quad$ $n = 0.14mol$

Limiting reagent: Iron (Fe)

3mol Fe $\rightarrow$ 1mol Fe_3O_4

0.3mol Fe $\rightarrow$ x

$x = \frac{0.3\ mol\ Fe \times 1mol\ Fe_3O_4}{3\ mol\ Fe}$

$x = 0.1mol\ Fe_3O_4$

Exercise

1. Aluminium reacts with Sulphuric acid, which is the acid in automobile battery, according to the equation below

$2Al_{(s)} + 3H_2SO_{4(aq)} \rightarrow Al_2(SO_4)_{3(aq)} + 3H_{2(g)}$

If 20.0g of aluminium are put in a solution containing 115g of Sulphuric acid;

(a) Which is the limiting reagent

(b) How many moles of hydrogen gas will be produced

(c) How many moles and grams of the reactant in excess will remain after the reaction has stopped?

2. A mixture of 8.0g of hydrogen with 8.0g of oxygen is ignited

$2H_2 + O_2 \rightarrow 2H_2O$

What is the mass of water formed?

Gas volume (Molar volume of gases)

Avogadro's law

The law states that: the volume of a gas is directly proportional to the number of moles of the gas molecules present if the pressure and temperature are constant

Number of moles $= \frac{\text{Volume}}{\text{Molar volume}}$

$n = \frac{V}{Vm}$

$V = n \times Vm$

n = number of moles [mol]

V = volume [cm^3] or [dm^3]

Vm = molar volume [cm^3/mol] or [dm^3/mol]

Room temperature and pressure (r.t.p): The volume of one mole of any gas is $24dm^3$ or $24000cm^3$ at r.t.p

Standard temperature and pressure (s.t.p): The volume of one mole of any gas is $22.4dm^3$ or $22400cm^3$ at s.t.p

Example

1. Calculate the number of moles of carbon dioxide molecules present in $240cm^3$ of gas at r.t.p

Solution

$n = \frac{V}{Vm}$

$n = \frac{240cm^3}{24000cm^3/mol}$

n = 0.01 mol

2. What is the volume of the following gases at s.t.p?

(a) 2.8×10^{-3}mol of Nitrogen, N_2

(b) 3.2g of oxygen, O_2

Solution

(a) $V = n \times Vm$

$V = 2.8 \times 10^{-3}$ mol x $22.4dm^3$/mol

$V = 0.0627dm^3$

(b) $n = \frac{m}{MM}$

$n = \frac{3.2g}{32g/mol}$

n = 0.1mol

V = n x Vm

V = 0.1mol x $22.4dm^3/mol$

V = $2.24dm^3$

Exercise

1. How many moles of molecules of each of the following gases would contain

(a) $12dm^3$ of hydrogen at r.t.p

(b) $100cm^3$ of carbon dioxide at s.t.p

Calculations from equations

Examples

1. What volume of hydrogen measured at s.t.p is produced when 0 .35g of Lithium reacts with water?

$$2Li_{(s)} + 2H_2O_{(l)} \rightarrow 2LiOH_{(aq)} + H_{2(g)}$$

Solution

$n_{(Li)} = \frac{m}{MM}$

$n = \frac{0.35g}{7g/mol}$

n = 0.05mol

2 mol Li → 1mol H_2

0.05mol Li → x

$x = \frac{0.05 \text{ mol Li x 1 mol } H_2}{2 \text{ mol Li}}$

x = 0.025mol H_2

V = n x Vm

V = 0.025mol x $22.4dm^3/mol$

V = $0.56dm^3$

2. What mass of ammonium sulphate is required to produce 5.6dm^3 of ammonia at s.t.p by the reaction shown below

$2NaOH_{(aq)} + (NH_4)_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)} + 2NH_{3(g)}$

Solution

$n\,(NH_3) = \frac{V}{Vm}$

$n = \frac{5.6\ dm^3}{22.4\ dm^3\ /mol}$

$n = 0.25mol\ NH_3$

$1mol\ (NH_4)_2SO_4 \rightarrow 2mol\ NH_3$

$x \rightarrow 0.25mol\ NH_3$

$x = \frac{1\ mol\ (NH_4)_2SO_4\ \ x\ 0.25\ mol\ NH_3}{2\ mol\ NH_3}$

$x = 0.125mol\ (NH_4)_2SO_4$

m = n x MM

m = 0.125mol x 132g/mol

$m = 16.5g\ (NH_4)_2SO_4$

Exercise

1. What volume of carbon dioxide measured at s.t.p will be produced when 21.0g of sodium hydrogen carbonate ($NaHCO_3$) is completely decomposed according to the equation?

   $2NaHCO_{3(s)} \rightarrow Na_2CO_{3(s)} + CO_{2(g)} + H_2O_{(l)}$

2. What volume of oxygen at s.t.p is required for complete combustion of 40cm^3 of butane (C_4H_{10}) according to the equation?

   $2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$

3. Methane burns completely in oxygen according to the equation below.
   $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$
   If 0.2mol of methane is burned completely, what volume of carbon dioxide measured at r.t.p is formed?
4. What volume of sulphur dioxide at r.t.p is given off on heating 9.7g of zinc sulphide, ZnS?
   $2Zn + O_2 \rightarrow 2ZnO + 2SO_2$
5. The following reaction takes place when chlorine is bubbled into excess ammonia solution.

   $3Cl_{2(g)} + 8NH_{3(g)} \rightarrow N_{2(g)} + 6NH_4Cl_{(s)}$

(a) What volume of nitrogen at r.t.p is formed when $72cm^3$ of chlorine is bubbled in ammonia solution?

(b) How many moles of nitrogen are there in this volume at r.t.p?

(c) How many grams of nitrogen are in these number of moles

6. What volume of oxygen at r.t.p is required for the complete combustion of $25.0cm^3$ of hexane according to the equation?

$2C_6H_{14} + 19O_2 \rightarrow 12CO_2 + 14H_2O_{(l)}$

Volume ration

Gas volume calculations are easy especially when the reactants and products are all gases.

Avogadro's law

The law states that: equal number of volumes of gases at the same temperature and pressure contain the same number of moles.

The significance of Avogadro's law is that the ratio of gas volumes in a reaction is the same as the number of moles in the equation: e.g.

$H_{2(g)}$	+	$Cl_{2(g)}$	$\rightarrow$	$2HCl_{(g)}$
1mol		1mol		2mol
1 volume		1 volume		2 volumes
$1cm^3 1cm^3$			$2cm^3$	

Example

1. Consider the reaction below

$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$

(a) What volume of oxygen is needed to react with $40cm^3$ of methane, CH_4?

(b) What volume of carbon dioxide would be produced?

Solution

(a) 1 volume $CH_4 \rightarrow$ 2 volumes O_2

$40cm^3$ $CH_4 \rightarrow x$

$$x = \frac{40cm^3\ CH_4\ x\ 2\ volumes\ O_2}{1\ volume\ CH_4}$$

$x = 80cm^3\ O_2$

(b) Volume of $CO_2 = 40cm^3$

2. When $100cm^3$ of hydrogen is sparked with $100cm^3$ of oxygen at 110^oC, steam is produced.

$2H_{2(g)} \quad + \quad O_{2(g)} \quad \rightarrow \quad 2H_2O_{(g)}$

(a) Which reactant is in excess?

(b) What volume of steam would be produced

Solution

(a) $2H_{2(g)} \quad + \quad O_{2(g)} \quad \rightarrow \quad 2H_2O_{(g)}$

2 : 1

$\frac{100cm^3}{2} : \frac{100cm^3}{1}$

50mol 100mol

Oxygen is in excess because it has a bigger number of moles

(b) 2 volumes $H_2 \quad \rightarrow \quad$ 2 volumes H_2O

$100cm^3 \quad H_2 \quad \rightarrow \quad x$

(c) $x = \frac{100cm^3\ H_2\ \text{x 2 volumes}\ H_2O}{2\ \text{volumes}\ H_2}$

$x = 100cm^3\ H_2O$

Relative molecular mass of gases

Density = $\frac{\text{mass}}{\text{volume}}$

Mass = Density x volume

Example

1. The density of a gas is $0.71g/dm^3$ at r.t.p. What is the mass of the gas?

Solution

Mass = Density x volume

$= 0.71g/dm^3 \text{ x } 24dm^3$

$= 17.04g$

Exercise

1. The mass of 1 dm^3 of a gas at s.t.p is 12.37g. What is the mass of the gas?
2. The density of a gaseous oxide of carbon is 1.15g/dm^3 at r.t.p. What is the mass of 1 mole of the gas?

Hydrocarbon analysis

Example

1. 20cm^3 of a hydrocarbon requires 100cm^3 of oxygen for complete combustion and produces 60cm^3 of carbon dioxide. What is the formula of the hydrocarbon?

Solution

C_xH_y + O_2 → CO_2 + H_2O

20cm^3 100cm^3 60cm^3

1 mol 5mol 3mol

C_3H_8 + $5O_2$ → $3CO_2$ + $4H_2O$

Formula: C_3H_8

2. 20cm^3 of hydrocarbon gas requires 100cm^3 of oxygen for complete combustion. 40cm^3 of carbon dioxide was formed and 40cm^3 of excess oxygen remained. What is the formula of the hydrocarbon?

Solution

C_xH_y + O_2 → CO_2 + H_2O

20cm^3 100cm^3 – 40cm^3 = 60cm^3 40cm^3

2mol 6mol 4mol

C_4H_8 + $6O_2$ → $4CO_2$ + $4H_2O$

Formula: C_4H_8

Exercise

1. 10cm^3 of a hydrocarbon gas reacts with 90cm^3 of oxygen to form 60cm^3 carbon dioxide. What is the formula of the hydrocarbon?
2. 20cm^3 of hydrocarbon gas requires 100cm^3 of oxygen for complete combustion. 40cm^3 of carbon dioxide was formed and 30cm^3 of excess oxygen remained. What is the formula of the hydrocarbon?

Concentration

Definition: Concentration is the amount of solute dissolved in a unit volume of the solution.

Units: g/dm^3

mol/dm^3

Formula: Concentration (g/dm^3) = $\frac{\text{mass of solute (g)}}{\text{volume of solution (dm}^3)}$

Concentration (mol/dm^3) = $\frac{\text{number of moles of solute (mol)}}{\text{volume of solution (dm}^3)}$

The concentration expressed in mol/dm^3 is called molarity.

Molarity is the number of moles of solute in one liter of solution

$M = \frac{n}{v}$

Note

M = Molarity [mol/dm^3]

n = moles [mol]

V = volume [dm^3]

Example

1. A solution of glucose contains 0.45g of glucose in $0.075dm^3$ of solution. Calculate the concentration of glucose solution in g/dm^3.

Solution

Concentration (g/dm^3) = $\frac{\text{mass of solute (mol)}}{\text{volume of solution (dm}^3)}$

$= \frac{0.45g}{0.075dm^3}$

$= 6.0g/dm^3$

2. Find the concentration in mol/dm^3 of a solution of sodium hydroxide if it contains 3.5g of NaOH in $100cm^3$ of solution.

Solution

$n = \frac{m}{MM}$

$n = \frac{3.5g}{40g/mol}$

$n = 0.0875mol$

$$\text{Concentration } (mol/dm^3) = \frac{\text{number of moles of solute (mol)}}{\text{volume of solution } (dm^3)}$$

$$= \frac{0.0875\ mol}{0.1\ dm^3}$$

$$= 0.875mol/dm^3$$

Exercise

1. What is the concentration of 0.5g of sodium hydroxide contained in 250 cm^3 of solution in:
   (a) g/dm^3
   (b) mol/dm^3

Converting Molarity (concentration inmol/dm^3) into g/dm^3

$$\text{Concentration } (mol/dm^3) = \frac{\text{concentration } (g/dm^3)}{\text{molar mass } (g/mol)}$$

$$\text{Concentration } (g/dm^3) = \text{Concentration}(mol/dm^3) \times \text{Molar mass } (g/mol)$$

Example

1. 4.5g of anhydrous sodium carbonate (Na_2CO_3) was dissolved in distilled water and made up to 500cm^3. Express the concentration of the solution in terms of :
   (a) g/dm^3
   (b) mol/dm^3

Solution

(a) $$\text{Concentration } (g/dm^3) = \frac{\text{mass of solute (g)}}{\text{volume of solution } (dm^3)}$$

$$= \frac{4.5g}{0.5\ dm^3}$$

$$= 9.0g/dm^3$$

(b) $$\text{Concentration } (mol/dm^3) = \frac{\text{concentration } (g/dm^3)}{\text{molar mass } (g/mol)}$$

$$= \frac{9.0g/dm^3}{106g/mol}$$

$$= 0.085mol/dm^3$$

Exercise

1. 1.5dm^3 of hydrogen chloride gas measured at room temperature and pressure was absorbed in aqueous ammonia solution and 50cm^3 of ammonia solution was neutralized completely

$$NH_{3(aq)} + HCl_{(aq)} \rightarrow NH_4Cl_{(aq)}$$

Calculate;

(a) The molarity of the ammonia solution

(b) The mass of ammonium chloride expected to be produced in the reaction

(c) The percentage yield of the ammonium chloride if 2.68g of ammonium chloride was obtained after evaporating the solution

Dilution

This is the process of adding more solvent to a solution in order to reduce the concentration. After dilution, the volume of the solution increases where as the concentration decreases. The concentration is inversely proportion to the volume change when the concentration changes from initial M_1 to new concentration M_2. This principle gives us the dilution law:

$M_1V_1=M_2V_2$ where;

M_1= initial concentration

V_1=initial volume

M_2=new concentration after dilution

V_2=new volume after dilution

The dilution law and the above formula can be used to solve stoichiometric dilution problems.

Examples

1. $0.25dm^3$ of concentrated sulphuric acid of concentration of 18.0M is diluted to give a solution of 2.0M. What is the new volume of the diluted solution?

   $M_1V_1=M_2V_2$

   M_1= 18M

   $V_1=0.25dm^3$

   M_2=2.0M

   V_2=?

$18M x 0.25dm^3 = 2.0M x V_2$

$V_2 = 2.25dm^3$

2. $50cm^3$ of 1.5M potassium hydroxide solution is diluted to $200cm^3$. Calculate the new concentration of the solution?

   $M_1V_1 = M_2V_2$

   $M_1 = 1.5M$

   $V_1 = 50cm^3$

   $M_2 = ?$

   $V_2 = 200cm^3$

   $1.5M x 50cm^3 = M_2 x 200cm^3$

   $M_2 = 0.375M$

3. What volume of water must be added to $30cm^3$ of 2M nitric acid solution to produce o.2M solution?

   $M_1V_1 = M_2V_2$

   $M_1 = 2M$

   $V_1 = 30cm^3$

   $M_2 = 0.2M$

   $V_2 = ?$

$2M x 30cm^3 = 0.2M x V_2$

   $V_2 = 300cm^3$

Since the final volume is $300cm^3$ and the initial volume is $30cm^3$ of 2M nitric acid volume of water that must be added; $300cm^3-30cm^3=270cm^3$ must be added.

4. The concentration of a stock solution of hydrochloric acid is 12M. Calculate the volume of the stock solution needed to prepare $250cm^3$ of 2.15M solution of hydrochloric acid.

   $M_1V_1=M_2V_2$

   $M_1= 12M$

   $V_1=?$

   $M_2=2.15M$

   $V_2=250cm^3$

   $12M x V_1=2.15M x 250cm^3$

   $V_1=44.79cm^3$

5. Calculate the volume of 4M hydrochloric acid required to make $2.5dm^3$ of 0.2M hydrochloric acid.

   $M_1V_1=M_2V_2$

   $M_1= 4M$

   $V_1=?$

   $M_2=0.2M$

   $V_2=2.5dm$

   $4M x V_1=0.2M x 2.5dm^3$

   $V_1=0.125dm^3$

6. Differentiate between a concentrated solution and a diluted solution.

   =A concentrated solution contains more solute than a diluted solution

Calculations from equations

Example

1. Calculate the mass of calcium hydroxide required to neutralize $2.5dm^3$ of 0.5M HCl acid.

$Ca(OH)_{2(aq)} \quad + \quad 2HCl_{(aq)} \quad \rightarrow \quad CaCl_{2(aq)} \quad + \quad 2H_2O_{(l)}$

Solution

$M_{(HCl)} = \frac{n}{V}$

$n = M \times V$

$n = 0.5mol/dm^3 \times 2.5dm^3$

$n = 1.25mol\ HCl$

$1\ mol\ Ca(OH)_2 \rightarrow 2mol\ HCl$

$x \rightarrow 1.25mol\ HCl$

$x = \frac{1\ mol\ Ca(OH)_2\ \times 1.25\ mol\ HCl}{2\ mol\ HCl}$

$x = 0.625mol\ Ca(OH)_2$

$m\ (Ca(OH)_2) = n \times MM$

$m = 0.625\ mol \times 74g/mol$

$m = 46.25g\ Ca(OH)_2$

2. 0.4M of sodium hydroxide was made to react completely with $50cm^3$ of 0.1M Sulphuric acid. What was the volume of sodium hydroxide used?

Solution

$2NaOH_{(aq)} \quad + \quad H_2SO_{4(aq)} \quad \rightarrow \quad Na_2SO_{4(aq)} \quad + \quad 2H_2O_{(l)}$

$n(H_2SO_4) = M \times V$

$n = 0.1mol/dm^3 \times 0.05dm^3$

$n = 0.005mol\ H_2SO_4$

$2mol\ NaOH \rightarrow 1mol\ H_2SO_4$

$x \rightarrow 0.005mol\ H_2SO_4$

$x = \frac{2\ mol\ NaOH \times 0.005\ mol\ H_2SO_4}{1\ mol\ H_2SO_4}$

$x = 0.01mol\ NaOH$

$V\ (NaOH) = \frac{n}{M}$

$V = \frac{0.01\ mol}{0.4\ mol\ /dm^3}$

$V = 0.025dm^3 = 25.0cm^3$

Exercise

1. Calculate the volume of 1M HCl which can be neutralized by 0.29g of magnesium hydroxide, $Mg(OH)_2$

   $Mg(OH)_{2(aq)}\ +\ 2HCl_{(aq)}\ \rightarrow\ MgCl_{2(aq)}\ +\ 2H_2O_{(l)}$

2. What volume of $0.1mol/dm^3$ of sulphuric acid would completely neutralize $20cm^3$ of $0.2mol/dm^3$ of sodium hydroxide?

Percentage by mass of an element in a compound

Formula: % by mass = $\frac{\text{number of atoms of element x } A_r \text{ of element}}{M_r \text{ of compound}}$ x 100%

Example

1. Calculate the % by mass of each element present in sodium carbonate, Na_2CO_3

Solution

(a) $M_{r\ of}\ Na_2CO_3$ = [(2 x 23) + (1 x 12) + (3 x 16)]
= 46 + 12 + 48
= 106

% by mass of Na $= \frac{2 \text{ x } 23}{106}$ x 100%

$= \frac{46}{106}$ x 100%

= 43.3%

(b) % by mass of C = $\frac{1 \text{ x } 12}{106}$ x 100%

= 11.3%

(c) % by mass of O = $\frac{3 \text{ x } 16}{106}$ x 100%

= 45.3%

Percentage by mass of H_2O in a compound = $\frac{\text{number of moles of } H_2O \text{ x } M_r \text{ of } H_2O}{M_r \text{ of compound}}$ x 100%

Example

1. Calculate the percentage by mass of water in sodium carbonate crystals, $Na_2CO_3.10H_2O$

Solution

$M_r of Na_2CO_3.10H_2O$ = [(2 x 23) + (1 x 12) + (3 x 16) + (20 x 1) + (10 x 16)]

= 46 + 12 + 48 + 20 + 160

= 286

% by mass of H_2O in $Na_2CO_3.10H_2O$ = $\frac{10 \text{ x } 18}{286}$ x 100%

= 62.9%

Exercise

1. Calculate the percentage by mass of water in copper (II) sulphate crystals, $CuSO_4.5H_2O$

Mass of element in sample

Formula: Mass of element in sample = $\frac{\text{Mass of element in compound}}{M_r \text{ of compound}}$ x sample mass

Example

1. Calculate the mass of copper (Cu) in 32g of copper (II) sulphate

Solution

Mass of Cu in $CuSO_4$ = $\frac{64}{160}$ x 32g

= 12.8g

Exercise

1. Calculate the mass of iron in 10.0g of sample of rust, Fe_2O_3

Empirical formula

Symbol: E.F

Definition: Empirical formula is a formula which shows the lowest ratio of the different atoms in a compound

Example

1. Find the empirical formula of a compound with the composition of 80% copper and 20% sulphur

Solution

Symbols of elements	Cu	S

Percentage by mas Molar mass	$\frac{80}{64}$	$\frac{20}{32}$
Number of moles	1.25	0.625
Divide by smallest number	$\frac{1.25}{0.625}$	$\frac{0.625}{0.625}$
Mole ratio	2	1
E.F = Cu_2S		

2. 30g of silicon oxide contains 14g of silicon. Find the empirical formula of the compound

Solution

Mass of O = 30g – 14

= 16g

Symbols of elements	Si	O
Percentage by mass Molar mass	$\frac{14}{28}$	$\frac{16}{16}$
Number of moles	0.5	1
Multiply both sides by 2	0.5 x 2	1 x 2
Mole ratio	1	2
E.F = SiO_2		

3. A compound was found to contain 29.4% calcium, 23.5% sulphur and 47.1% oxygen. What is the empirical formula of the compound?

Solution

Symbols of elements	Ca	S	O
Percentage by mass Molar mass	$\frac{29.4}{40}$	$\frac{23.5}{32}$	$\frac{47.1}{16}$
Number of moles	0.735	0.734	2.944
Divide by smallest number	$\frac{0.735}{0.734}$	$\frac{0.734}{0.734}$	$\frac{2.944}{0.734}$
Mole ratio	1	1	4
E.F = $CaSO_4$			

4. A hydrated salt has the following percentage composition iron 20.15%, sulphur 11.51%, oxygen 23.02% and water 45.32%. find out its empirical formula

Solution

Symbols of elements	Fe	S	O	H_2O
Percentage by mass Molar mass	$\frac{20.15}{56}$	$\frac{11.51}{32}$	$\frac{23.02}{16}$	$\frac{45.32}{18}$
Number of moles	0.35	0.35	1.43	2.51
Divide by smallest number	$\frac{0.35}{0.35}$	$\frac{0.35}{0.35}$	$\frac{1.43}{0.35}$	$\frac{2.51}{0.35}$
Mole ratio	1	1	4	7
E.F = $FeSO_4.7H_2O$				

Exercise

1. A hydrocarbon contains 85.7% carbon and 14.3% hydrogen. Work out the empirical formula.
2. A compound X contains 50% sulphur and 50% oxygen, by mass. What is the empirical formula of the compound?

3. An experiment showed that 13.88g of calcium chloride were obtained from the combination of 5g calcium with chlorine. What is the empirical formula of calcium chloride?
4. Analysis of an organic compound showed that 5.4g of carbon combined with 0.9g of hydrogen and 0.8g of oxygen. What is the empirical formula of this organic compound?

Molecular formula

Symbol: M.F

Definition: Molecular formula is a formula which shows the actual number of each kind of atom in a compound

Formula: Molecular formula = $(\text{emperical formula})_n$

$$\text{M.F} = (\text{EF})_n$$

$$n = \frac{\text{Relative molecular mass}}{\text{Relative emperical formula mass}}$$

Example

1. The empirical formula of a compound is C_2H_4O. Its relative molecular mass is 88. Find the molecular formula

Solution

$$n = \frac{\text{Relative molecular mass}}{\text{Relative emperical formula mass}}$$

$$n = \frac{88}{44}$$

$n = 2$

$\text{M.F} = (\text{E.F})_n$

$\text{M.F} = (C_2H_4O)_2$

$\text{M.F} = C_4H_8O_2$

2. A compound of carbon has a composition of 15.8% carbon and 84.2% sulphur. Find the empirical formula. If the relative molecular mass of the compound is 76, find its molecular formula

Symbols of elements	C	S
Percentage by mass / Molar mass	$\frac{15.8}{12}$	$\frac{84.2}{32}$
Number of moles	1.317	2.631
Divide by smallest number	$\frac{1.317}{1.317}$	$\frac{2.631}{1.317}$
Mole ratio	1	2
E.F = CS_2		

n = $\frac{\text{Relative molecular mass}}{\text{Relative emperical formula mass}}$

n = $\frac{76}{76}$

n = 1

M.F = $(\mathrm{E.F})_n$

M.F = $(CS_2)_1$

M.F = CS_2

Exercise

1. 4.04g of nitrogen combines with 11.46g of oxygen to produce a compound with relative molecular mass of 92. What is the molecular formula of the compound?
2. A compound contains 40% carbon, 6.7% hydrogen and 53.3% oxygen.
   (a) Find the empirical formula of the compound
   (b) If the relative molecular mass of the compound is 180, work out its molecular formula.
3. The empirical formula of a compound is found to be CH_2. The relative molecular mass of the compound is known to be 42g. Find the molecular formula of the compound.

Percentage yield

Definition: Percentage yield is the ratio of the actual yield to the theoretical yield multiplied by 100%

Formula: Percentage yield = $\frac{\text{Actual yield}}{\text{Theoretical yield}}$ x 100%

Yield

Definition: Yield is the amount of product produced when substances react.

Actual yield

Definition: Actual yield is the amount of product obtained at the end of the reaction

Actual yield is also called experimental yield

Theoretical yield

Definition: Theoretical yield is the maximum amount of the product calculated from the measured amounts of a given reactant using the balanced chemical equation and the mole concept.

Theoretical yield is also called expected yield

Note

The actual yield is usually less than the theoretical yield due to the following factors:

- the reactants may fail to react completely due to the presence of impurities in them
- the reaction may be reversible such that some of the products formed are converted back into the reactants preventing the reaction from going to completion
- there may be evaporation of both reactants and products for the reaction involving volatile chemicals
- there may be loss of samples of reactants or products as they are being transferred from one vessel to another during measurements
- for large scale industrial processes, there may be leakages in pipes carrying the reactants or products of the reaction

Percentage purity

Definition: Percentage purity is the ratio of the amount of a pure substance in the sample to the total amount of the sample multiplied by 100%

Formula: Percentage purity = $\frac{\text{Mass of pure sample}}{\text{Massof impure sample}}$ x 100%

Percentage purity = $\frac{\text{Mass of impure sample – mass of pure sample}}{\text{Mass of impure sample}}$ x 100%

Example

1. Consider the reaction below

   $C_6H_{6\,(l)} + Cl_{2(g)} \rightarrow C_6H_5Cl_{\,(l)} + HCl_{(g)}$

   When 36.8g of benzene (C_6H_6) reacts with an excess of chlorine (Cl_2), the actual yield of chlorobenzene (C_6H_5Cl) is 38.8g. Find the percentage yield of chlorobenzene (C_6H_5Cl)

Solution

78g C_6H_6 $\rightarrow$ 112.5g C_6H_5Cl

36.8g C_6H_6 $\rightarrow$ x

$x = \frac{36.8g\ C_6H_6 \text{ x } 112.5g\ C_6H_5Cl}{78g\ C_6H_6}$

x = 53.08g C_6H_5Cl(Theoretical yield)

Percentage yield $= \frac{\text{Actual yield}}{\text{Theoretical yield}}$x 100%

$= \frac{38.8g}{53.08g}$x 100%

= 73.1%

2. Marble occurs naturally as rock and contains a lot of calcium carbonate

$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

(a) If you started with 25g of marble, what would be the expected mass of calcium chloride?

(b) If the actual yield of calcium chloride obtained is 13.8g, what is the percentage purity of calcium chloride, given that;

Percentage purity $= \frac{\text{Actual yield}}{\text{Theoretical yield}}$x 100%

(c) What is the likely impurity in the sample product?

Solution

(a) 100g $CaCO_3$ $\rightarrow$ 111g $CaCl_2$

25g $CaCO_3$ $\rightarrow$ x

$x = \frac{25g\ CaCO_3 \text{ x } 111g\ CaCl_2}{100g\ CaCO_3}$

x =27.75g $CaCl_2$

(b) Percentage purity $= \frac{\text{Actual yield}}{\text{Theoretical yield}}$ x 100%

$= \frac{13.8g}{27.75g}$ x 100%

= 41.73%

(c) Silicon dioxide, SiO_2

Exercise

1. Methanol can be produced through the reaction of carbon monoxide (CO) and hydrogen (H_2) in the presence of a catalyst.

   $CO_{(g)} \quad + \quad 2H_{2(g)} \quad \rightarrow \quad CH_3OH_{(l)}$

   If 75.0g of carbon monoxide (CO) reacts to produce 68.4g of methanol (CH_3OH), what is the percentage yield of methanol, CH_3OH?

2. Aluminium(Al) reacts with excess copper (II) sulphate ($CuSO_4$) according to the equation given below:

   $2Al_{(s)} \quad + \quad 3CuSO_{4(aq)} \quad \rightarrow \quad Al_2(SO_4)_{3(aq)} \quad + \quad 3Cu_{(s)}$

   If 1.85g of aluminium (Al) reacts and the percentage yield of copper (Cu) is 56.6%, what mass of copper (Cu) is produced?

3. Calculate the percentage yield when 6.4g of copper are heated in air and 7.6g of copper (II) oxide is obtained.
4. 5g of calcium carbonate was reacted with $40cm^3$ of $2.75mol/dm^3$ hydrochloric acid according to the equation below:
   $CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

   If $1.1dm^3$ of carbon dioxide gas measured at r.t.p was collected in a graduated gas jar.

(a) State the practical yield of carbon dioxide gas
(b) State which reactant was in excess
(c) State which reactant was a limiting reactant
(d) Calculate the theoretical yield
(e) Calculate the percentage yield.

The periodic table

The periodic table is a chart of elements placed according to the order of increasing atomic numbers.

Atomic number is the property of elements used to place them in order on the periodic table.

Important features of the periodic table

PERIODS	GROUP																	
	I	II											III	IV	V	VI	VII	O
P 1								H										He
E 2	Li	Be											B	C	N	O	F	Ne
R 3	Na	Mg	**TRANSITION ELEMENTS**										Al	Si	P	S	Cl	Ar
I 4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
O 5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
D 6	Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Ti	Pb	Bi	Po	At	Rn
S 7	Fr	Ra	Ac															

Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

☐	Metals
■	Non- metals

$^{A}_{Z}X$

Key: A = Relative atomic mass / mass number

X = Atomic symbol

Z = Proton (atomic) number

In the periodic table,

1. Elements are arranged in the order of increasing atomic numbers
2. Elements with similar electronic configuration are placed in the same group.
3. Elements with similar chemical properties are placed in the same group

Period

A period is a horizontal row of elements in the periodic table

The periodic table has 7 periods.

Period 1 contains two elements; hydrogen and helium

Periods 2 and 3 contains eight elements each and are called short periods.

Periods 4, 5 and 6 contain eighteen elements each and are called long periods.

Period 7 contains three elements

The period number is equivalent to the number of electron shells e.g. calcium is placed in period 4 of the periodic table because it has 4 electron shells i.e. Ca 2.8.8.2

Ca	2	8	8	2
	1^{st} shell	2^{nd} shell	3^{rd} shell	4^{th} shell

Progression of properties across the period

As we move across the period (from left to right)

1. The metallic nature of elements decreases. They change from metals to non-mental and inert gases.
2. Electro-negativity increases. Electronegativity is the relative ability of an atom to attract the pair of electrons in a covalent bond.
3. The atomic number increases by one between successive elements.
4. The number of shells remains the same while the number of valence electrons increases steadily. The valence of each element is equal to;
   (a) The valence electrons for metals.

Examples

(i) Na 2. 8.1, valence = 1

(ii) Mg 2.8.2, valence = 2

(iii) Al 2.8.3, valence = 3

(b) Eight (8) minus the valence electrons for non-metals

Examples

(i) Cl 2.8.7
Valency = 8 −7
= 1

(ii) O 2.6
Valency = 8 − 6
= 2

(iii) N 2.5
Valency = 8 − 5
= 3

Groups

A group is a vertical column of elements in the periodic table

There are eight groups on the periodic table

Groups are labeled with roman numerals i.e. groups I, II, III, IV, V, VI, VII with the final group labeled O

The position of an element in the periodic table is determined by the number of electrons in the outer most shell.

The number of electrons in the outer most shell of an element is equivalent to its group number e.g. oxygen is placed in group VI of the periodic table because it has 6 electrons in the outer most shell i.e. O 2.6

Some groups have special names

Examples

Group	Special name
I	Alkali metals
II	Alkaline earth metals
VII	Halogens

O	Noble gases/ Inert gases/ rare gases

Progression of properties down the group

As we move from top to bottom (down the group)

1. The metallic nature of elements increases
2. The atomic number increases
3. Elements in the same group have similar chemical properties
4. Elements in the same group have the same number of electrons in the outer most shell.
5. For non-metals, the group number is equal to eight minus the valence.
6. For metals, the group number is equal to the valence.

Position of hydrogen in the periodic table

Hydrogen is placed between group I and group VII because it behaves like group I and group VII elements i.e. it can lose or gain a single electron.

Hydrogen can lose one electron to form hydrogen ion, $H^+$with 1+ charge and can gain one electron to form hydride ion, $H^-$with 1− charge.

Zig-zag diagonal line

The zig-zag diagonal line in the periodic table divides metallic elements from non-metallic elements.

Metalloids

Alternative term: Semi metals

Metalloids are elements near the Zig-zag diagonal line.

Metalloids have the characteristics of both metals and non – metals..

Group properties

Group I elements

Alternative term: Alkali metals

They are called alkali metals because they react with water to form alkalis (alkaline solutions).

Examples of group I elements

Element	Symbol	Atomic number
Lithium	Li	3

Sodium	Na	11
Potassium	K	19
Rubidium	Rb	37
Caesium	Cs	55
Francium	Fr	87

Occurrence of group I elements

Group I elements do not occur naturally as free elements because they are very reactive.

They are found in compounds e.g. rock salt (impure sodium chloride) which is a good source of sodium.

Identification of group I elements

Group I elements and their compounds give characteristic colours in a flame.

Examples

Element	Characteristic colour
Lithium	Red flame
Sodium	Yellow / orange flame
Potassium	Lilac / Pinkish flame

Storage of group I elements

Group I element are stored under oil to prevent them from reacting with atmospheric air or water.

Physical properties of group I elements

1. They are soft metals which can be cut with a knife.
2. They are good conductors of heat and electricity.
3. They have low densities and hence float on water as they react with it. Their densities increase as you go down the group.
4. They have low melting and boiling points. Their melting points decreases as you go down the group.

Name	Symbol	Density, g/cm^3	Melting point, ^oC
Lithium	Li	0.53	180

Sodium	Na	0.78	98
Potassium	K	0.86	64
Rubidium	Rb	1.5	39
Caesium	Cs	1.9	29

Trend in chemical reactivity of group I elements

The reactivity of group I elements increases as you go down the group.

Lithium is the least reactive and francium is the most reactive element in group I.

Chemical properties of group I elements

1. They have a single electron in their outer most shells.

Example

(a) Li 2.1
(b) Na 2.8.1
(c) K 2.8.8.1

2. They lose their single electrons in their outer most shells to form ions with 1+ charge.

Examples

(a) Li – ẽ → Li^+
(b) Na – ẽ → Na^+
(c) K – ẽ → K^+

3. They are powerful reducing agents because they lose electrons.
4. They react with oxygen to form basic oxides

Examples

(a) $4Li_{(s)} + O_{2(g)} \rightarrow 2Li_2O_{(s)}$
(b) $4Na_{(s)} + O_{2(g)} \rightarrow 2Na_2O_{(s)}$
(c) $4K_{(s)} + O_{2(g)} \rightarrow 2K_2O_{(s)}$

5. They react with water to form metal hydroxides and hydrogen gas.

Examples

(a) $2Li_{(s)} + 2H_2O_{(l)} \rightarrow 2LiOH_{(aq)} + H_{2(g)}$

(b) $2Na_{(s)}$ + $2H_2O_{(l)}$ → $2NaOH_{(aq)}$ + $H_{2(g)}$
(c) $2K_{(s)}$ + $2H_2O_{(l)}$ → $2KOH_{(aq)}$ + $H_{2(g)}$

Note

- The metal hydroxide solutions formed are all strong alkalis with PH values of around 14
- Lithium reacts violently with water
- Sodium reacts very violently with water, sometimes with an explosion
- Potassium reacts explosively with water

6. They react with group VII elements to form salts

Examples

(a) $2Li_{(s)}$ + $F_{2(g)}$ → $2LiF_{(s)}$
(b) $2Na_{(s)}$ + $Cl_{2(g)}$ → $2NaCl_{(s)}$
(c) $2K_{(s)}$ + $Br_{2(l)}$ → $2KBr_{(aq)}$

Group VII elements

Alternative term: Halogens

They are called Halogens because they react with group I elements to form salts

They term halogen means salt former

Examples of group VII elements

Element	**Symbol**	**Atomic number**
Fluorine	F	9
Chlorine	Cl	17
Bromine	Br	35
Iodine	I	53
Astatine	At	85

Occurrence of group VII elements

Group VII elements do not occur naturally in a free state, instead they exist as diatomic molecules meaning two atoms chemically combined.

Physical Properties of Group VII Elements

1. They exist as **diatomic** covalent molecules (meaning two atoms chemically combined)

Examples

Element	Molecular formula
Fluorine	F_2
Chlorine	Cl_2
Bromine	Br_2
Iodine	I_2

2. They exist as coloured,non- metallic elements

Examples

Element	Colour
Fluorine	Pale yellow
Chlorine	Yellowish – green
Bromine	Reddish-brown
Iodine	Black

3. They show a gradual change in their physical states at room temperature and pressure.

Examples

Element	Physical state at room temperature and pressure
Fluorine	Gas
Chlorine	Gas
Bromine	Liquid
Iodine	Solid

4. Their melting and boiling points increases as you go down the group.

Element	Melting point/°C	Boiling point /°C
Fluorine	-220	-188
Chlorine	-101	-35
Bromine	-7	59
Iodine	114	184

5. Their densities increases as you go down the group.
6 Their compounds can either be ionic or covalent. If they combine with a metal the compound is ionic and if they combine with another nonmetal the compound is covalent.

Trend in Chemical reactivity of Group VII Elements

The reactivity of group VII elements increases as you go up the group.

Astatine is the least reactive halogen while fluorine is the most reactive halogen

Chemical properties of group VII elements

1. They all have seven electrons in their outer shells and hence have similar chemical properties.

Examples

(a) F 2.7
(b) Cl 2.8.7
(c) Br 2.8.18.7

2. They gain a single electron to form ions with 1− charge

Examples

(a) $F + \tilde{e} \rightarrow F^-$
(b) $Cl + \tilde{e} \rightarrow Cl^-$
(c) $Br + \tilde{e} \rightarrow Br^-$
(d) $I + \tilde{e} \rightarrow I^-$

3. They are oxidizing agents because they accept /gain electrons.
4. They react with group I elements to form salts

Examples

(a) $2\ Li_{(s)} + F_{2(g)} \rightarrow 2LiF_{(s)}$
(b) $2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$
(c) $2K_{(s)} + Br_{2(g)} \rightarrow 2KBr_{(aq)}$

5. They displace each other in chemical reactions involving their halide ions i.e. the group VII Ions. The more reactive halogen will displace the less reactive halogen from its aqueous salt solution.

Examples

(a) $2NaCl_{(aq)} + F_{2(g)} \rightarrow 2NaF_{(aq)} + Cl_{2(g)}$
(b) $2NaBr_{(aq)} + Cl_{2(g)} \rightarrow 2NaCl_{(aq)} + Br_{2(1)}$
(c) $2KI_{(aq)} + Br_{2(l)} \rightarrow 2KBr_{(aq)} + I_{2(s)}$
(d) $NaCl_{(aq)} + Br_{2(l)} \rightarrow$ No reaction

Uses of group VII elements

1. Fluorine is used in fluoride tooth paste to help prevent tooth decay.
2. Chlorine is used to sterilize drinking water because it kills harmful micro-organisms.
3. Iodine is used as an additive in table salt to prevent goiter in human beings.

Harmful Effects of Group VII Elements

Compounds of group VII elements are known to be responsible for the depletion of the Ozone layer.

Note

Ozone is an allotrope of oxygen made up of three oxygen atoms (O_3)

Group O elements

Alternative term: Noble gases / Inert gases / Rare gases

They are called noble gases or inert gases or rare gases because they are chemically unreactive and therefore do not form compounds.

Group O elements are chemically unreactive because they have full outer most electron shells. Their outer most shells are completely filled.

Group O elements are gases at room temperature and pressure.

Group O elements exist as unreactive monatomic elements with very low melting and boiling points.

They consist of single atoms.

Examples of Group O elements.

Element	Symbol	Atomic number
Helium	He	2
Neon	Ne	10
Argon	Ar	18
Krypton	Kr	36
Xenon	Xe	54
Radon	Rn	86

Group O elements have eight electrons in the outer most shell except for helium which has only two electrons.

Examples

(a) He 2
(b) Ne 2.8
(c) Ar 2.8.8
(d) Kr 2.8.18.8
(e) Xe 2.8.18.18.8

Uses of group O elements

1. Helium is used to fill weather balloons because of low density.
2. Neon is used to fill coloured glowing tubes used in advertisements because it glows red hot in an electric current.
3. Argon is used to fill light bulbs to provide an inert atmosphere to prevent the oxidation of the filament.

Transition elements

Transition elements are found in the centre block of the periodic table. They are found between group II and III of the periodic table and through periods 4 and 6

They are all metals.

Examples of transition elements

Element	Symbol	Atomic number

Manganese	Mn	25
Iron	Fe	26
Copper	Cu	29
Zinc	Zn	30
Silver	Ag	47
Gold	Au	79
Mercury	Hg	80

Physical properties of transition elements

1. They have high densities
2. They have high melting and boiling points
3. They are good conductors of heat and electricity
4. They are solids at room temperature and pressure except mercury which is a liquid
5. They are ductile i.e. they can be drawn into wires
6. They are malleable i.e. they can be hammered into thin sheets.

Chemical properties of transition elements

1. They are catalysts
2. They have variable Valencies and form positively charged ions.

Examples

(a) Copper has a valency of 1 or 2 and forms the ions $Cu^{+}$and Cu^{2+}respectively.
(b) Iron has a valence of 2 or 3 and forms the ions $Fe^{2+}$and Fe^{3+} respectively.

3. They form coloured compounds depending on the valence used.

Examples

(a) Copper (II) compounds are blue
(b) Iron (II) compounds are pale green
(c) Iron (III) compounds are reddish brown

4. They are reducing agents because they lose electrons.

Uses of transition elements

1. They are used to make electric cables because they are good conductors of electricity.
2. They are used to make pots and pans because they are good conductors of heat.
3. They are used to make alloys

Examples

(a) Brass is an alloy of copper and Zinc

(b) Bronze is an alloy of copper and tin

4. They are used as catalysts in the industry to speed up reactions.

Examples

(a) Iron is used as a catalyst in Haber process.
(b) Nickel is used as a catalyst in the hydrogenation of oil to make margarine
(c) Platinum is used as a catalyst in contact process.

Example

1. Using the following evidence, state the group numbers for the elements A, B, C, D and E on the periodic table:

(I) An oxide A_2O exists and is strongly basic
(II) B form a liquid covalent chloride, BCl_3
(III) The oxide of C is C_2O_3
(IV) D produces an ion D^-
(V) E exists as an unreactive monoatomic element.
(a) Using the given letters, write the formula for the compound formed between C and D

Solution

(I) Group I
(II) Group V
(III) Group III
(IV) Group VII
(V) Group O
(a) CD_3

Exercise

1. The diagram below shows part of the periodic table.

				He
C	N	O	F	Ne
		S	Cl	Ar
			Br	Kr

(a) Answer these questions using only the elements shown in the diagram.
Write down the symbol for an element which

(i) Has five electrons in the outer most shell
(ii) Has diatomic molecules
(iii) Reacts with sodium to form sodium bromide
(iv) Is a noble gas.
(v) Has a lower proton than fluorine
(b) Why is argon very unreactive

2. Use the periodic table to answer this question.
(a) Name the element in group II and Period 3 on the periodic table.
(b) State whether the element named in (a) is a metal or non-metal.
(c) Suggest the formula of the compound formed between the element named in (a) and sulphur, S.
(d) Group VII of the periodic table contains fluorine and chlorine. Explain why these elements have similar chemical properties.

3. An element has atomic number of 16.

(a) Use the periodic table to name the element and give the symbol.

(b) Explains why

(i) The element is placed in group VI of the periodic table
(ii) The element has a valence of two in its compound with magnesium.
(iii) An ion of this element has two negative charges.

4. An element is in group I of the periodic table. Another element is in group VII.
(a) Suggest two ways in which the properties of these elements must be different.
(b) Two elements are in group I of the periodic table one is placed in period 2 and the other is in period 3. State two ways in which these elements must be chemically similar.

5. Use the periodic table to help you answer this question.
(a) Name the element in group V which is in the same period as Lithium.
(b) Calculate the number of protons and neutrons in one atom of the element you have named in (a)
(c) (i) Name the element with a relative atomic amass of 31
(ii) Suggest one property, physical or chemical of this element.

6. Some elements have only seven electrons in their outer most shells.
(a) (i) Name three of these elements

(ii) In which of the group of the periodic table are they placed?

(b) Use your knowledge of physicals and chemical properties of these elements to just placing them in the same group of the periodic table.

7. Caesium, Lithium, Potassium and sodium are all in group I of the periodic table
(a) Place these metals in order of reactivity with water, most reactive metal first.
(b) Name the chemical products of the reactions between lithium and water and between sodium and water.
(c) (i) What would you expect to see if small pieces of Caesium were dropped onto water? How would the PH of the resulting solution be different from the PH of water?
(ii) Write the full chemical equation for the reaction between Caesium and water. Include state symbols.
8. Chlorine, bromine and iodine are placed in this order in group VII of the periodic table.
(a) State four ways in which the physical or chemical properties of chlorine, bromine and iodine are similar.
(b) (i) Describe the trends in physical properties of chlorine, bromine and iodine.
(ii) How is the trend in chemical reactivity of chlorine, bromine and iodine shown by the displacement reactions? Give an equation for a reaction in which one element displaces another from its compound.

9. The diagram below shows a table of elements taken from a larger classification of elements given on the periodic table.

I	II		III	IV	V	VI	VII	O
		hydrogen						helium
Lithium	beryllium		boron	Carbon	Nitrogen	oxygen	fluorine	neon
Sodium	magnesium		aluminum	Silicon	phosphorous	sulphur	chlorine	argon

(a) Which group contains
(i) Halogens
(ii) Alkali metals
(b) From the diagram above, choose
(i) A metal from period 2
(ii) A non-metal with a valence of 2
(c) Hydrogen is difficult to classify into a group as it can be compared with both chlorine and sodium.
(i) How many electrons are there in an atom of hydrogen?
(ii) What is the valence of a hydrogen atom?
(iii) State one way in which hydrogen is similar to chlorine
(iv) State one way in which hydrogen is similar to sodium.

10. Use the periodic table to answer this question

(a) Give the symbol of:

(i) A non-metal used to sterilize water

(ii) An element which forms diatomic molecules

(iii) An element which reacts with water to give an alkaline solution

(iv) An element which forms an ion of the type x^{2-}

(b) Oxygen, sulphur and selenium are in group VI of the periodic table. At room . . temperature, oxygen is a gas and sulphur is a solid.

(i) Predict whether selenium is a liquid, a solid or a gas at room temperature

(ii) The trend in reactivity of group VI is similar to that in group VII. Suggest the most reactive element in group VII.

Acids

Definition: An acid is a chemical substance which when dissolved in water produces hydrogen ions,H^+, as the only positively charged ions.

An acid can also be defined as a proton donor.

Note

The hydrogen ions give an acid its characteristic properties. All acids act alike because they contain hydrogen ions

Types of acids

Acids can be classified into two categories

1. **Mineral acids**

Mineral acids are also called inorganic acids

They are acids that are prepared from the minerals obtained from the earth

Generally all mineral acids are strong acids

Examples of mineral acids	Formula
Hydrochloric acid	HCl
Nitric acid	HNO_3
Sulphuric acid	H_2SO_4
Phosphoric acid	H_3PO_4

2. **Organic acids**

Organic acids are naturally occurring acids

They are found in organic matter of living things

They are prepared from organic sources such as plants

Generally organic acids are weak acids

Examples of organic acids	Formula
Methanoic acid	$HCOOH$
Ethanoic acid	CH_3COOH
Propanoic acid	C_2H_5COOH
Butanoic acid	C_3H_7COOH

Ionization of acids

Ionization is the process of forming ions

An ion is a charged particle

When acids dissolve in water, they produce hydrogen ions, H^+, as the only positively charged ions

Examples of ionization of acids

(a) $HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$

(b) $HNO_{3(aq)} \rightarrow H^+_{(aq)} + NO_3^-{}_{(aq)}$

(c) $CH_3COOH_{(aq)} \rightarrow H^+_{(aq)} + CH_3COO^-{}_{(aq)}$

(d) $H_2SO_{4(aq)} \rightarrow 2H^+_{(aq)} + SO_4^{2-}{}_{(aq)}$

(e) $H_3PO_{4(aq)} \rightarrow 3H^+_{(aq)} + PO_4^{3-}{}_{(aq)}$

Strength of acids

Strength of an acid is the measure of its ability to produce hydrogen ions

It can also be defined as the ability to donate protons

Acids can be grouped into two classes

1. **Weak acids**

A weak acid is an acid that partially ionizes when dissolved in water

Weak acids have reversible ionization $\rightleftharpoons$

Examples of weak acids

(a) Ethanoic acid

$$CH_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_3COO^-_{(aq)}$$

(b) Carbonic acid

$$H_2CO_{3(aq)} \rightleftharpoons 2H^+_{(aq)} + CO^{2-}_{3(aq)}$$

2. **Strong acids**

A strong is an acid that completely ionizes when dissolved in water

Examples of strong acids

(a) Hydrochloric acid

$$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

(b) Nitric acid

$$HNO_{3(aq)} \rightarrow H^+_{(aq)} + NO^-_{3(aq)}$$

(c) Sulphuric acid

$$H_2SO_{4(aq)} \rightarrow 2H^+_{(aq)} + SO^{2-}_{4(aq)}$$

(d) Phosphoric acid

$$H_3PO_{4(aq)} \rightarrow 3H^+_{(aq)} + PO^{3-}_{4(aq)}$$

Basicity of an acid

Alternative term: Protocity of an acid.

Definition: Basicity of an acid is the number of moles of hydrogen ions, H^+, produced from one mole of an acid

1. Monobasic acid

Alternative term: Monoprotic acid

A monobasic acid is an acid that produces one hydrogen ion, $1H^+$

Example of monobasic acids	Formula	Equation	Basicity
Hydrochloric acid	HCl	$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$	1
Nitric acid	HNO_3	$HNO_{3(aq)} \rightarrow H^+_{(aq)} + NO^-_{3(aq)}$	1
Ethanoic acid	CH_3COOH	$CH_3COOH_{(aq)} \rightarrow H^+_{(aq)} + CH_3COO^-_{(aq)}$	1

2. Dibasic acid

Alternative term: Diprotic acid

A dibasic acid is an acid that produces two hydrogen ions, $2H^+$

Examples of dibasic acid	Formula	Equation	Basicity
Carbonic acid	H_2CO_3	$H_2CO_{3(aq)} \rightarrow 2H^+_{(aq)} + CO^{2-}_{3(aq)}$	2
Sulphuric acid	H_2SO_4	$H_2SO_{4(aq)} \rightarrow 2H^+_{(aq)} + SO^{2-}_{4(aq)}$	2

3. Tribasic acid

Alternative term: Triprotic acid

A tribasic is an acid is an acid that produces three hydrogen ions, $3H^+$

Example of tribasic acid	Formula	Equation	Basicity
Phosphoric acid	H_3PO_4	$H_3PO_{3(aq)} \rightarrow 3H^+_{(aq)} + PO^{3-}_{4(aq)}$	3

Physical properties of acids

1. Acids have PH numbers less than 7
2. Acids turn blue litmus paper red
3. Acids have a sour taste

Chemical properties of acids

1. Acids react with **reactive metals** to form a salt and hydrogen gas, H_2

Examples

(a) $2Na_{(s)} + 2HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_{2(g)}$

(b) $Ca_{(s)} + H_2SO_{4(aq)} \rightarrow CaSO_{4(aq)} + H_{2(g)}$

Note

Metals below hydrogen in the reactivity series cannot react with acids

2. Acids react with bases and alkalis to form a salt and water only. The reaction between an acid and a base is called **neutralization**.

Examples

(a) $NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

(b) $CuO_{(s)} + H_2SO_{4(aq)} \rightarrow CuSO_{4(aq)} + H_2O_{(l)}$

Ionic equations for neutralization reactions

$$\mathbf{OH^- + H^+ \rightarrow H_2O}$$

3. Acids react with carbonates or hydrogen carbonates to form a salt, water and carbon dioxide gas

Examples

(a) $CaCO_{3(s)} + H_2SO_{4(aq)} \rightarrow CaSO_{4(aq)} + H_2O_{(l)} + CO_{2(g)}$

(b) $NaHCO_{3(s)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

Uses of acids in our every daily life

Name of acid	Every day use
Benzoic acid	Its salts are used to preserve food
Carbonic acid	To make carbonated drinks
Ethanoic acid	It is the main compound of vinegar used in the kitchen, is a liquid containing 3-6% acetic acid. It is used in pickles and in many food preparations
Hydrochloric acid	To clean metals before electroplating / household cleaning / leather processing / swimming pool maintenance
Nitric acid	Production of fertilizers, explosives, etching and dissolution of metals (purification and extraction of gold)
Sulphuric acid	To make detergent and fertilizers such as super phosphate, ammonium sulphate is used in batteries, which are used in cars, etc. Tannic acid is used in the manufacture of ink and leather
Tartaric acid	Manufacturing of soft drinks, provide tartness to food, as an emetic (a substance to induce vomiting)

Bases

Definition: A base is a substance which reacts with an acid to form a salt and water

A base can also be defined as an oxide or hydroxide of a metal or a proton acceptor

A base neutralizes an acid

A soluble base is called an **alkali**

All alkalis act alike because they contain hydroxide ions

An alkali is a substance which when dissolved in water produces the hydroxide ions, OH^-, as the only negatively charged ions

Examples of alkalis

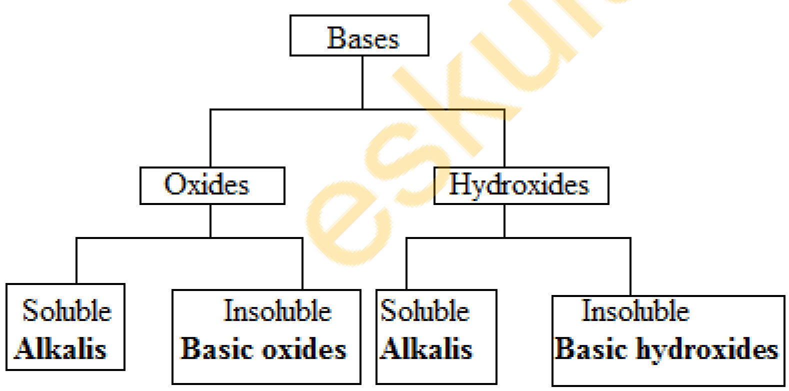
Name	Formula	Special name
Sodium hydroxide	$NaOH$	Caustic soda
Potassium hydroxide	KOH	Caustic potash

Ammonium hydroxide	NH_4OH	Aqueous ammonia
Calcium hydroxide	$Ca(OH)_2$	Lime water

Examples of insoluble bases

Name	Formula
Copper (II) oxide	CuO
Zinc oxide	ZnO
Iron (III) oxide	Fe_2O_3
Lead (II) hydroxide	$Pb(OH)_2$

Simple classification on bases



Note

All alkalis are bases but not all bases are alkalis

Ionization of alkalis

When alkalis dissolve in water, they produce hydroxide ions, OH^-, as the only negatively charged ions.

Examples of ionization of alkalis

(a) $NaOH_{(aq)} \rightarrow Na^{+}_{(aq)} + OH^{-}_{(aq)}$

(b) $KOH_{(aq)} \rightarrow K^{+}_{(aq)} + OH^{-}_{(aq)}$

(c) $NH_4OH_{(aq)} \rightarrow NH_4^{+}{}_{(aq)} + OH^{-}_{(aq)}$

(d) $Ca(OH)_{2(aq)} \rightarrow Ca^{2+}_{(aq)} + 2OH^{-}_{(aq)}$

Strength of alkalis

Strength of an alkali is the measure of its ability to produce hydroxide ions

Alkalis can be classified into two categories; weak and strong alkalis.

1. **Weak alkalis**

A weak alkali is an alkali that partially ionizes when dissolved in water

Weak alkalis have reversible ionization $\rightleftharpoons$

Example of a weak alkali

(a) Ammonium hydroxide

$NH_4OH_{(aq)} \rightleftharpoons NH_4^{+}{}_{(aq)} + OH^{-}_{(aq)}$

2. **Strong alkalis**

A strong alkali is an alkali that ionizes completely when dissolved in water

Examples of strong alkalis

(a) Sodium hydroxide

$NaOH_{(aq)} \rightarrow Na^{+}_{(aq)} + OH^{-}_{(aq)}$

(b) Potassium hydroxide

$KOH_{(aq)} \rightarrow K^{+}_{(aq)} + OH^{-}_{(aq)}$

(c) Calcium hydroxide

$Ca(OH)_{2(aq)} \rightarrow Ca^{2+}_{(aq)} + 2OH^{-}_{(aq)}$

Physical properties of alkalis

1. Alkalis have PH numbers greater than 7
2. Alkalis turn red litmus paper blue
3. Alkalis have a bitter taste and feel soapy or slippery between fingers

Chemical properties of bases and alkalis

1. Alkalis and bases react with acids to form a salt and water only

Examples

(a) $NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

(b) $CuO_{(s)} + H_2SO_{4(aq)} \rightarrow CuSO_{4(aq)} + H_2O_{(l)}$

2. Alkalis reacts with ammonium compounds to form a salt, water and ammonia gas

Examples

(a) $NaOH_{(aq)} + NH_4Cl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)} + NH_{3(g)}$

(b) $Ca(OH)_{2(aq)} + NH_4NO_{3(aq)} \rightarrow Ca(NO_3)_{2(aq)} + H_2O_{(l)} + NH_{3(g)}$

3. Alkalis react with solutions of soluble salts to form precipitates of insoluble hydroxides

Examples

(a) $2NaOH_{(aq)} + CuSO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + Cu(OH)_{2(s)}$

(b) $2NaOH_{(aq)} + FeSO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + Fe(OH)_{2(s)}$

4. Alkalis react with acidic oxides to form a salt and water only

Examples

(a) $2NaOH_{(aq)} + SO_{3(g)} \rightarrow Na_2SO_{4(aq)} + H_2O_{(l)}$

(b) $Ca(OH)_{2(aq)} + SO_{3(g)} \rightarrow CaSO_{4(aq)} + H_2O_{(l)}$

Some neutralization in everyday life

Insect stings: When a bee stings, it injects an acidic liquid into the skin. The sting can be neutralized by rubbing on calamine lotion, which contains zinc carbonate or baking soda, which is sodium hydrogen carbonate.

Wasp stings are alkaline, and can be neutralized with vinegar. Anti-stings and nettle stings contain Methanoic acid

Indigestion: The stomach contains hydrochloric acid. It is a very dilute solution and it is needed for digesting food. But too much of it leads to indigestion, which can be very painful. To cure indigestion, the excess hydrochloric acid can be neutralized with a drink of sodium hydrogen carbonate solution (baking soda) or an indigestion tablet

Soil treatment: Most plants grow best when the PH of the soil is close to 7. If the soil is too acidic, or too alkaline, the plants grow badly or not at all.

Chemicals can be added to the soil to adjust its PH. Most often it is too acidic, so it is treated with quick lime (calcium oxide), slaked lime (calcium hydroxide) or chalk (calcium carbonate). These are all bases and quiet cheap.

Factory waste: Liquid waste from factories often contains acid. If it reaches a river, the acid will kill fish and other river life. This can be prevented by adding slaked lime to the waste, to neutralize it.

Uses of bases in our daily life

Name of base	**Every day use**
Ammonia	Production of fertilizers (ammonium and nitrate salts), used in the manufacture of nitric acid, neutralize the acid (in the petroleum industry) and prevent premature coagulation in natural / synthetic latex
Aluminium hydroxide	Manufacture other aluminium compound and to make gastric medicine (antacid)
Calcium hydroxide	To make cement, limewater, neutralize the acidity of soil and application of sewage treatment.
Sodium hydroxide	Used in the manufacturing of soaps, detergents, and cleaners
Magnesium hydroxide	Suspensions of magnesium hydroxide in water are used as an antacid, used as an antiperspirant armpit deodorant and as a non-hazardous alkali to neutralize acidic wastewater.

The PH scale

Definition: PH is the degree of acidity or alkalinity of a substance

The PH scale ranges from 0 to 14

PH values from 0 to 7 implies acidic medium.

PH values 7 imply a neutral medium e.g. water, H_2O and sodium chloride, NaCl

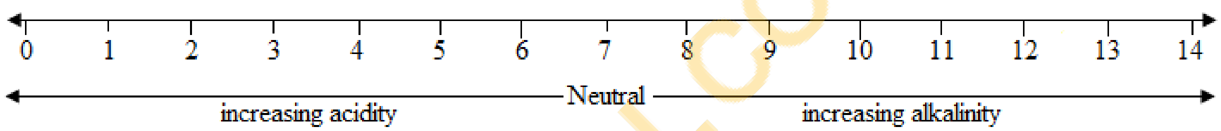
PH values from 7 to 14 implies alkaline medium

The lower the PH value, the stronger the acidity e.g. hydrochloric acid (HCl), sulphuric acid (H_2SO_4) and nitric acid (HNO_3)

The more H^+ ions in a solution, the more acidic it be. In other words, the more H^+ ions there are, the lower the PH number.

The greater the PH value, the stronger the alkalinity e.g. Potassium hydroxide (KOH) and sodium hydroxide (NaOH)

The more OH^- ions in a solution, the more alkaline it be. In other words, the more OH^- ions there are, the higher the PH number.



The ion responsible for acidity is the hydrogen ion, H^+ and the ion responsible for alkalinity is the hydroxide ion, OH^-

PH is a numerical value and has no units

PH is related to the concentration of acids and alkalis.

Significance of PH measurements

Apart from enabling us to determine whether substances are acidic or alkaline, PH values have very important significance and implications in industry, agriculture, pharmacy and medicine

Control of PH in agriculture

Plants grow well in which type of soil. Some plants grow well in acidic soils (PH 6.5), while others grow well in alkaline soils (PH 7.5)

If the ground is too acidic because of acid rain, slaked lime can help to neutralize excess acidity. This process is called liming.

Soil acidity

Formation of acid rain

Acid rain is formed when gaseous acidic oxides dissolve in rain water in the atmosphere

What is destroyed by acid rain?

1. Soil – the soil becomes acidic
2. Crops and vegetation

Examples of acidic rain fall

1. **Nitric acid**

Nitric acid is formed when nitrogen dioxide gas dissolve in rain water in the atmosphere

$NO_{2(g)} + H_2O_{(l)} \rightarrow HNO_{3(aq)}$

2. **Carbonic acid**

Carbonic acid is formed when carbon dioxide gas dissolve in rain water in the atmosphere

$CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(aq)}$

3. **Sulphuric acid**

Sulphuric acid is formed when sulphur trioxide gas dissolve in rain water in the atmosphere

$SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$

All these acids come down as acid rainfall

Acid rain causes pollution

Acidic soils on the farm can be neutralized or avoided by adding lime (Calcium oxide)

Indicators

Definition: An indicator is a chemical substance which has different colours with different substances or solutions

Indicator	**Colour in acid**	**Colour in alkali**
Litmus	Red	Blue
Methyl orange	Red	Yellow

Bromothymol	Yellow	Blue
Phenolphthalein	Colourless	Red / pink

Universal indicator

A universal indicator is a solution or paper which shows different colors in acidic or alkaline conditions

PH scale	1,2,3	4,5	6	7	8	9,10	11,12,13,14
Colour	Red	Orange	Yellow	Green	Blue	Indigo	Purple

Exercise

1. Acid rain is partially caused by Sulphur dioxide

(a) What is damaged by acid rain

(b) How is acid rain formed?

(c) Suggest how this pollution can be avoided

2. The PH of a solution depends upon the ions present in it

(a) State the PH of a neutral solution

(b) Complete the table below

Solution in water	**PH range**
Acids	
Alkalis	

(c) What gas is formed when;

(i) Dilute Sulphuric acid is added to sodium carbonate

(ii) Sodium hydroxide is warmed with ammonium chloride

(iii) Write a balanced chemical equation for either of the reactions in (c)

(d) Explain why sulphuric acid is said to be a strong acid.

3. The figure below shows a PH scale and the PH values of four solutions A, B, C and D.

(a) Which of the solutions is likely to be:

(I) Sodium chloride

(II) Hydrochloric acid

(b) State the ion which is responsible for:

(I) Acidity

(II) Alkalinity

(c) What type of reaction occurs when solution A is mixed with solution D? Write an ionic equation for the reaction that occurs.

4. Analysis of cigarette smoke and cigar smoke were carried out. The results showed that cigarette smoke is acidic and cigar smoke is alkaline with the following PH values:

Smoke	**PH**
Cigarette	5.2
Cigar	8.3

(a) State the ion which is formed when:

(I) Cigarette smoke is absorbed in water and is responsible for the PH value given

(II) Cigar smoke is absorbed in water and is responsible for the PH value given

(b) When cigar smoke is mixed with cigarette smoke, neutralization occurs. What is the meaning of neutralization?

(c) If cigarette smoke is absorbed in universal indicator solution, what colour change would be observed?

(d) Name an alkali that would neutralize cigarette smoke.

(e) Smoking cigarette can be a danger to health. This fact is well known in Zambia. The main danger of cigarette smoking is that it can cause lung cancer. Carbon monoxide is present in cigarette smoke. It does not cause lung cancer, yet is toxic. Describe and explain the poisonous or toxic nature of carbon monoxide.

(f) How would you use blue litmus paper to show that cigarette smoke is acidic?

Oxides

Definition: An oxide is a compound formed when oxygen combines with any other element

Element + Oxygen → Oxide

Classification of oxides

[A] Acidic oxides

They are oxides of non-metals

Acidic oxide	Formula
Carbon dioxide	CO_2
Nitrogen dioxide	NO_2
Sulphur dioxide	SO_2
Sulphur trioxide	SO_3
Silicon dioxide	SiO_2

They are formed when a non – metal reacts with oxygen

Example

(a) $C + O_2 \rightarrow CO_2$
(b) $N_2 + 2O_2 \rightarrow 2NO_2$
(c) $S + O_2 \rightarrow SO_2$
(d) $2SO_2 + 3O_2 \rightarrow 2SO_3$
(e) $Si + O_2 \rightarrow SiO_2$

Characteristics of acidic oxides

1. They dissolve in water to form acids

Examples

(a) $CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(aq)}$ (Carbonic acid)

(b) $NO_{2(g)} + H_2O_{(l)} \rightarrow HNO_{3(aq)}$ (Nitric acid)

(c) $SO_{2(g)} + H_2O_{(l)} \rightarrow H_2SO_{3(aq)}$ (Sulphurus acid)

(d) $SO_{3(g)} + H_2O_{(g)} \rightarrow H_2SO_{4(aq)}$ (Sulphuric acid)

Note

- Dinitrogen tetra oxide forms two acids when dissolved in water

  $N_2O_{4(g)} + H_2O_{(l)} \rightarrow HNO_{2(aq)} + HNO_{3(aq)}$

  (Nitrous acid) (Nitric acid)

- They are called acidic oxides because they react with water to form acids

2. They react with alkalis to form a salt and water only

Examples

(a) $SO_{3(g)} + 2NaOH_{(aq)} \rightarrow Na_2SO_{4(aq)} + H_2O_{(l)}$

(b) $SO_{3(g)} + Ca(OH)_{2(aq)} \rightarrow CaSO_{4(aq)} + H_2O_{(l)}$

(c) $CO_{2(g)} + 2NaOH_{(aq)} \rightarrow Na_2CO_{3(aq)} + H_2O_{(l)}$

[B] **Basic oxides**

They are oxides of metals

Basic oxide	Formula

Sodium oxide	Na_2O
Copper (II) oxide	CuO
Iron (III) oxide	Fe_2O_3
Calcium oxide	CaO
Magnesium oxide	MgO
Potassium oxide	K_2O

They are formed when a metal reacts with oxygen

Example

(a) $4Na + O_2 \rightarrow 2Na_2O$
(b) $2Cu + O_2 \rightarrow 2CuO$
(c) $4Fe + 3O_2 \rightarrow 2Fe_2O_3$
(d) $2Ca + O_2 \rightarrow 2CaO$
(e) $2Mg + O_2 \rightarrow 2MgO$
(f) $4K + O_2 \rightarrow 2K_2O$

Characteristics of basic oxides

1. Some soluble basic oxides dissolve in water to form alkalis

Examples

(a) $Na_2O_{(s)} + H_2O_{(l)} \rightarrow 2NaOH_{(aq)}$

(b) $CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)}$

2. They react with acids to form a salt and water only

Examples

(a) $Na_2O_{(s)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + H_2O_{(l)}$

(b) $CaO_{(s)} + H_2SO_{4(aq)} \rightarrow CaSO_{4(aq)} + H_2O_{(l)}$

[C] Amphoteric oxides

They are oxides of metals

Amphoteric oxide	Formula
Zinc oxide	ZnO
Aluminium oxide	Al_2O_3
Lead (II) oxide	PbO

They are formed when a metal reacts with oxygen

Example

(a) $2Zn + O_2 \rightarrow 2ZnO$
(b) $4Al + 3O_2 \rightarrow 2Al_2O_3$
(c) $2Pb + O_2 \rightarrow 2PbO$

Amphoteric oxides show both basic and acidic properties i.e. they react with both acids and alkalis to form a salt and water

(a) Reaction with acids

Examples

(I) $ZnO_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_2O_{(l)}$

(II) $Al_2O_{3(s)} + H_3PO_{4(aq)} \rightarrow AlPO_{4(aq)} + H_2O_{(l)}$

(III) $PbO_{(s)} + H_2SO_{4(aq)} \rightarrow PbSO_{4(s)} + H_2O_{(l)}$

(b) Reaction with alkalis

Examples

(I) $ZnO_{(s)} + 2NaOH_{(aq)} \rightarrow Na_2ZnO_{2(aq)} + H_2O_{(l)}$

(Sodium zincate)

(II) $Al_2O_{3(s)} + 2NaOH_{(aq)} \rightarrow 2NaAlO_{2(aq)} + H_2O_{(l)}$

(Sodium aluminate)

(III) $PbO_{(s)} + 2NaOH_{(aq)} \rightarrow Na_2PbO_{2(aq)} + H_2O_{(l)}$

(Sodium plumbite)

[D] Neutral oxides

Neutral oxides are oxides of non – metals

Neutral oxide	**Formula**
Carbon monoxide	CO
Nitrogen monoxide	NO
Water (Hydrogen oxide)	H_2O
Dinitrogen oxide	N_2O

Neutral oxides do not show either basic or acidic properties i.e. they do not react with either bases or acids

Salts

Definition: A salt is a chemical substance formed when the hydrogen ions in an acid are replaced by a metal or ammonium ions, NH_4^+

Types of salts

1. **Normal salt**

This is a salt formed when all the hydrogen ions in an acid are replaced by a metal or ammonium ions.

Examples

1. $NaOH_{(aq)} + \underline{H}Cl_{(aq)} \rightarrow \underline{Na}Cl_{(aq)} + H_2O_{(l)}$
2. $Ca_{(s)} + \underline{H}_2SO_{4(aq)} \rightarrow \underline{Ca}SO_{4(aq)} + H_{2(g)}$
3. $2NaOH_{(aq)} + \underline{H}_2SO_{4(aq)} \rightarrow \underline{Na}_2SO_{4(aq)} + 2H_2O_{(l)}$

Normal salt	**Formula**	**Common name**
Sodium chloride	$NaCl$	Common salt/Table salt
Calcium sulphate	$CaSO_4$	
Sodium sulphate	Na_2SO_4	Glauber's salt
Copper (II) sulphate	$CuSO_4$	

Calcium carbonate	$CaCO_3$	Limestone/Marble
Magnesium sulphate	$MgSO_4$	Epsom salt

2. Acid salt

This is a salt formed when part of the hydrogen ions in an acid are replaced by a metal or ammonium ions.

Examples

1. $KOH_{(aq)} + \underline{H}_2SO_{4(aq)} \rightarrow \underline{K}HSO_{4(aq)} + H_2O_{(l)}$
2. $NH_4OH_{(aq)} + \underline{H_2}SO_{4(aq)} \rightarrow \underline{NH_4}HSO_{4(aq)} + H_2O_{(l)}$
3. $NaOH_{(aq)} + \underline{H}_2SO_{4(aq)} \rightarrow \underline{Na}HSO_{4(aq)} + H_2O_{(l)}$
4. $CaO_{(s)} + 2\underline{H}_2CO_{3(aq)} \rightarrow \underline{Ca}(HCO_3)_{2(aq)} + H_2O_{(l)}$

Acid salt	Formula	Common name
Potassium hydrogen sulphate	$KHSO_4$	
Sodium hydrogen sulphate	$NaHSO_4$	
Calcium hydrogen carbonate	$Ca(HCO_3)_2$	
Sodium hydrogen carbonate	$NaHCO_3$	Baking/Bicarbonate salt

Solubility of salts

Solubility is the ability of a salt to dissolve in water.

Facts about solubility of salts in water

Salt	Solubility
Nitrates	All nitrates are soluble
Chlorides	All chlorides are soluble except silver chloride and lead (II) chloride which are insoluble
Sulphates	All Sulphates are soluble except barium sulphate and lead (II) sulphate which are insoluble. Calcium sulphate is slightly soluble.
Carbonates	All carbonates are insoluble except potassium carbonate, sodium carbonate and ammonium carbonate which are soluble.

Preparation of salts

The method chosen to prepare a given salt depends on its solubility and how it can be separated from the mixture of other products.

Methods of preparing salts

1. Neutralization
2. Replacement (Displacement)
3. Synthesis
4. Precipitation (Double Decomposition)

1. **Neutralization**

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

Acid + Base → Salt + Water

Neutralization includes the preparation of a salt by reacting:

(a) a metal hydroxide with a dilute acid.
(b) an insoluble metal oxide with a dilute acid.
(c) a metal carbonate with a dilute acid.

Preparation of soluble salts by neutralization

(a) Metal hydroxide + Acid → Salt + Water

Example: Preparation of sodium chloride

Reagents

- Sodium hydroxide, NaOH
- Hydrochloric acid, HCl

Products

- Sodium chloride, NaCl
- Water, H_2O

Reaction equation

$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

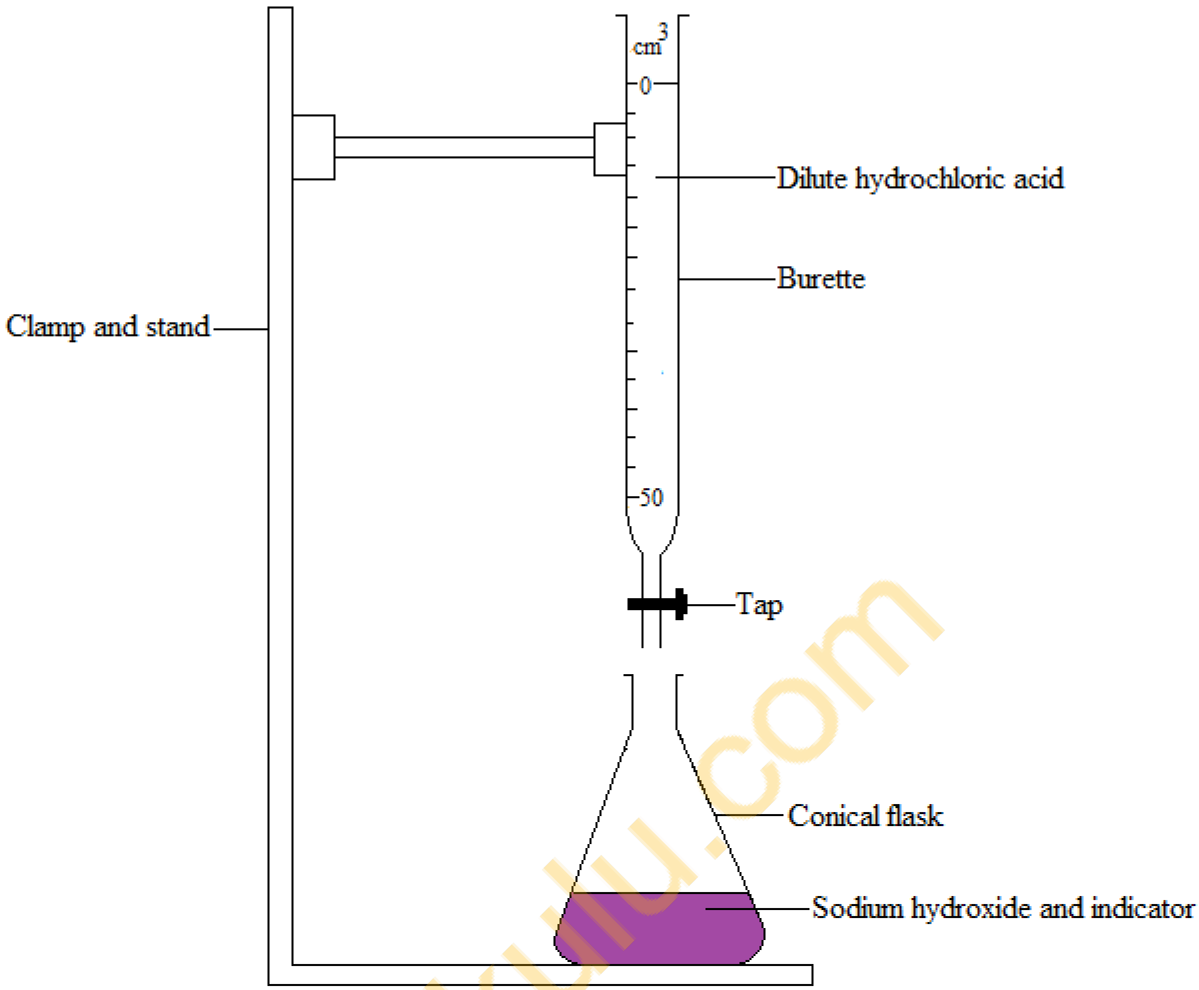
Ionic equation

$OH^-_{(aq)} + H^+_{(aq)} \rightarrow H_2O_{(l)}$

Method of preparation

A preliminary titration is carried out to find the end point with the help of an indicator.

- Using a pipette, measure $25.0cm^3$ of sodium hydroxide and put it into a conical flask.
- Add two or three drops of indicator to sodium hydroxide using a teat pipette.
- Fill the burette to the zero reading with dilutehydrochloric acid.
- Place the conical flask on a white tile below the burette.
- Add dilute hydrochloric acid from the burette to sodium hydroxide in the conical flask, while swirling, until the mixture just changes colour.
- From the titration result, we can know the exact volume of hydrochloric acid needed to react with $25.0cm^3$ of sodium hydroxide.
  Volume of hydrochloric used: $V_a = V_2 - V_1$



The results are recorded in the table below:

Burette reading

	Rough	1	2	3
Final burette reading/cm^3				
Initial burette reading/cm^3				
volume of HCl used/cm^3				
Tick best titration result(√)				

Summary

Tick (√) the best titration results

Using these results, the average volume of hydrochloric acidused was __________cm^3

Volume of sodium hydroxide used was______________________cm^3

A second titration is carried out without the indicator. The exact volume obtained from the preliminary titration is used.

- Using a pipette, measure 25.0 cm^3 of sodium hydroxide and put it into a conical flask. This time no indicator is added.
- Add V_a cm^3 of dilute hydrochloric acid from the burette to sodium hydroxide.
- Evaporate the mixture to obtain a saturated solution.
- Cool the saturated solution to obtain crystals of sodium chloride.
- Filter the mixture to obtain pure crystals sodium chloride.

Note

- The point at which the colour changes is called **end point.**
- At the end point, the volume of acid used is measured at the bottom of the meniscus in the burette. This volume is called **titre.**

Exercise

1. 6.00g of sample containing iron was obtained for analysis. All the iron in it was converted into iron (II) and the solution made up to 1000cm^3.
   P is a solution containing iron (II) ions, prepared as described above.
   Q is 0.02mol/dm^3 potassium manganate (VII) solution
   (a) **Q**was put into the burette
   25cm^3 portions of P was pipetted into a flask. No indicator was required. **Q**was added from the burette until a faint pink colour persisted.
   The results are recorded in the table. Since **Q** is dark, it was suitable to read the upper meniscus.
   Results
   Burette reading

	Rough	**1**	**2**
Final burette reading/cm^3	16.9	33.5	16.5
Initial burette reading/cm^3	0.0	16.9	0.0
volume of Qused/cm^3	16.9	16.6	16.5
Tick best titration result(√)			

Summary
Tick (√) the best titration results

Using these results, the average volume of solution **Q** used was
_____________________________cm^3

Volume of solution P used was_________________________cm^3

The reaction between iron (II) ions and manganate (VII) is given below

$MnO^-_{4(aq)} + 8H^+_{(aq)} + 5Fe^{2+}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + 5Fe^{3+}_{(aq)} + 4H_2O_{(l)}$

(b) Q is 0.02mol/dm^3 of potassium manganate (VII).
Using your answer to (a), calculate the number of moles of potassium manganate (VII) in the average volume of **Q**.

(c) Using your answer to (b) above and the equation given, calculate the number of moles of iron (II) in 25cm^3.

(d) Using your answer to (c), calculate the mass of iron in 25cm^3
(Molar mass of iron is 56g/mol)

(e) Using your answer to (d), calculate the mass of iron in 1000cm^3.

(f) Using your answer to (e), calculate the percentage of iron in the original sample

2. Sodium phosphate, Na_3PO_4, is a soluble salt, used as water softener in washing powders.

   It is made by reacting dilute phosphoric acid, H_3PO_4, with an alkali.

   (a) Give the formula of the ions present in sodium phosphate.

   (b) (i) Name the alkali which reacts with phosphoric acid to make sodium phosphate.

   (ii) Write an equation for this reaction.

   (c) Given solutions of phosphoric acid and an alkali, a suitable indicator and standard laboratory apparatus, explain how you would obtain crystals of sodium phosphate.

   (d) Sodium carbonate also reacts with phosphoric acid to make sodium phosphate.

   (i) Name the gas formed during this reaction.

   (ii) Describe the test for this gas.

   (e) Calcium sulphate is used to make fertilizers.

   (i) Deduce the formula of calcium phosphate.

   (ii) What essential plant element does calcium phosphate provide?

(b) Insoluble metal oxide + Acid → Salt + Water

Example: Preparation of copper (II) sulphate

Reagents

- Copper (II) oxide, CuO
- Dilute sulphuric acid, H_2SO_4

Products

- Copper (II) sulphate, $CuSO_4$
- Water, H_2O

Reaction equation

$CuO_{(s)} + H_2SO_{4(aq)} \rightarrow CuSO_{4(aq)} + H_2O_{(l)}$

Method of preparation

Copper (II) sulphate crystals $CuSO_4.5H_2O$ are prepared from copper (II) oxide and sulphuric acid.

- Put dilute sulphuric acid in a beaker, warm the acid but don't boil.
- Add an excess of copper (II) oxide to the acid in the beaker and stir until the reaction is over. Excess copper (II) oxide is added to ensure that all the sulphuric acid is converted to copper (II) sulphate.
- Filter off excess copper (II) oxide to have a blue solution of copper (II) sulphate.
- Evaporate the filtrate to obtain a saturated solution.
- Cool the filtrate in an ice bath to allow crystals of copper (II) sulphate to form.

Exercise

1. Copper (II) sulphate crystals, $CuSO_4.5H_2O$, are soluble in water. Describe how these crystals could be prepared from a named acid and a suitable compound of copper.

(c) Metal carbonate + Acid → Salt + Water + Carbon dioxide

Example: Preparation of Magnesium sulphate

Reagents

- Magnesium carbonate, $MgCO_3$
- Dilute sulphuric acid, H_2SO_4

Products

- Magnesium sulphate, $MgSO_4$
- Water, H_2O
- Carbon dioxide, CO_2

Reaction equation

$MgCO_{3(s)} + H_2SO_{4(aq)} \rightarrow MgSO_{4(aq)} + H_2O_{(l)} + CO_{2(g)}$

Method of preparation

- Put dilute sulphuric acid in a beaker, warm the acid but don't boil.
- Add an excess of magnesium carbonate to the acid to ensure that all the sulphuric acid is converted to magnesium sulphate.
- Stir the mixture and filter to have a clear solution of magnesium sulphate.
- Evaporate the mixture to obtain a saturated solution.
- Cool the filtrate (saturated solution) in an ice bath to allow crystals of magnesium sulphate to form.

Exercise

1. The following procedure was used by Darlington Naosa junior to prepare magnesium sulphate crystals: "Dilute sulphuric acid was put in a beaker and an excess of magnesium carbonate was added to the acid. The mixture stirred and filtered. The filtrate obtained was evaporated to half its original volume. The filtrate was then cooled in an ice bath".

(a) State why ;-

(I) An excess of magnesium carbonate was added to dilute sulphuric acid.

(II) The reaction mixture was filtered.

(III) The filtrate was evaporated to half its original volume.

(IV) The filtrate was cooled.

2. Epsom salt is a commercial or trade name for hydrated magnesium sulphate crystals, $MgSO_4.7H_2O$
The crystals are formed from a solution of magnesium sulphate, $MgSO_{4(aq)}$ by a suitable process.
A pupil wishes to prepare Epsom salt, in its pure, dry form by starting with an insoluble solid reactant X with a known dilute acid Y.
(a) Give the name and formula of the acid Y

(b) Give the formulae of any three possible reactants to be used as solid X in this experiment

(c) When X is reacted with Y, a gas which turns lime water milky is produced.

(I) Use this information to construct a balanced chemical equation for the reaction of X and Y, including state symbols.

(d) The pupil deliberately adds more of the solid X to $100cm^3$ of $1.0mol/dm^3$ acid.

(I) Explain why the pupil chooses to use the solid in excess

(II) What separation technique must the pupil use at this stage to obtain the solution of magnesium sulphate?

(e) The crystals of Epsom salt, $MgSO_4.7H_2O$ are then collected from the solution of $MgSO_{4(aq)}$. The following steps are involved (not in their order)

1. Filter
2. Dry crystals
3. Heat solution
4. Allow solution to cool slowly

Arrange these four steps in their logical order for the dry, pure crystals of $MgSO_4.7H_2O$ to be collected from $MgSO_{4(aq)}$.

(f) Excess of the solid x is used or added to $100cm^3$ of $0.1mol/dm^3$ acid Y.

(I) Calculate the mass of the crystals of $MgSO_4.7H_2O$ expected to be produced in the experiment

(II) Calculate the percentage yield if 20.0g of $MgSO_4.7H_2O$ is actually collected

Hydrated salt

This is a salt that contains water of crystallization. They contain a fixed amount of water in their crystal lattice. This is called water of crystallization. The water of crystallization is part of the structure. If this water is removed, by heating for example, the colour and shapes of the crystals may change.

Examples of hydrated salts

Name of crystallized salt	**Formula**
Copper (II) sulphate -5 water	$CuSO_4.5H_2O$
Sodium carbonate -10 water	$Na_2CO_3.10H_2O$
Cobalt (II) chloride - 6 water	$CoCl_2.6H_2O$

Iron (II) sulphate - 7 water	$FeSO_4.7H_2O$

Anhydrous salts

A salt which has lost its water of crystallization is called an **anhydrous salt**.

When water is added to an anhydrous salt, the salt becomes **hydrated**.

For example, when blue copper (II) sulphate crystals are heated, stem is produced and a pale- blue or white powder.

$CuSO_4.5H_2O_{(s)}$ → $CuSO_{4(s)}$ + $5H_2O_{(g)}$

(Hydrated copper sulphate) (Anhydrous copper sulphate) (Steam)

When water is added to anhydrous copper (II) sulphate heat is produced and a blue solution is formed:

$CuSO_{4(s)} + 5H_2O_{(l)} \rightarrow CuSO_4.5H_2O_{(aq)}$ + heat

This process is called **hydration.**

Efflorescence

This is the loss of water of crystallization to the atmosphere.

Example

Crystals of sodium carbonate – 10 − water become Powderly when exposed to air.

$Na_2CO_3.10H_2O_{(s)} \rightarrow Na_2CO_3.H_2O_{(s)} + 9H_2O_{(l)}$

Deliquescence

This is the absorption of water from the atmosphere to form a solution. Calcium chloride is a deliquescent salt. It is used as a drying agent in desiccators. A desiccator is a piece of equipment used to dry substances.

Hygroscopic

A hygroscopic substance absorbs water from the air but does not the change its state.

Anhydrous cobalt chloride is a hygroscopic salt.

Water changes anhydrous cobalt chloride from blue to pink.

$CoCl_{2(s)} + 6H_2O_{(l)} \rightarrow CoCl_2.6H_2O_{(s)}$ + heat

(Blue) (Pink)

This reaction is often used as a test for the presence of water. The process can be reversed by heating the pink hydrated salt:

$CoCl_2.6H_2O_{(s)} \rightarrow CoCl_{2(s)} + 6H_2O_{(g)}$

(Pink) (Blue)

Concentrated sulphuric acid is also hygroscopic. It can be used to dehydrate blue crystals of hydrated copper (II) sulphate forming the pale blue anhydrous salt.

2. **Replacement (Displacement)**

This is a method where the hydrogen ions in an acid are replaced by a metal. It can also be defined as a reaction in which one element displaces another from a compound

Reactive metal + Acid → Salt + Hydrogen

Example: Preparation of iron (II) sulphate

Method

- Add iron fillings to warm dilute sulphuric acid in a beaker until no more hydrogen gas is evolved.
- Filter off the solution when the reaction is complete to remove excess iron.

Reaction equation

$Fe_{(s)} + H_2SO_{4(aq)} \rightarrow FeSO_{4(aq)} + H_{2(g)}$

Note

- Iron should be in excess so that all the acid is used up.
- Air must be excluded to prevent oxidation of iron (II) sulphate.

Exercise

1. Name one salt that can be prepared by the reaction of a metal with a dilute acid

3. **Synthesis**

This method involves the direct combination of elements for binary salts.

In this method, a salt is prepared directly from its elements i.e. a metal and a halogen.

Metal + Halogen → Salt

Example: Preparation of iron (II) chloride

Iron (II) chloride can be prepared by passing chlorine gas over heated iron.

Reaction equation

$Fe_{(s)} + Cl_{2(g)} \rightarrow FeCl_{2(s)}$

4. **Precipitation**

Precipitation is the formation of an insoluble product and may occur on mixing two solutions.

Precipitation is an example of double decomposition. In double decomposition, two solutions are mixed to form an insoluble salt and a soluble solution.

Soluble salt + soluble salt → insoluble salt + soluble salt

Soluble solution + soluble solution → insoluble solid + soluble solution

In double decomposition reactions, cations and anions are exchanged.

Precipitation is also an example of ionic association which is the attraction of oppositely charged ions to one another to form a solid called **precipitate** abbreviated as **ppt.**

Preparation of insoluble salts by precipitation (Double decomposition)

Example 1: Preparation of silver chloride

Reagents

- Silver nitrate, $AgNO_3$
- Sodium chloride, NaCl (Alternatively hydrochloric acid, HCl)

Products

- Silver chloride, AgCl
- Sodium nitrate, $NaNO_3$

Reaction equation

$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$

Ionic equation

$Ag^{+}_{(aq)} + Cl^{-}_{(aq)} \rightarrow AgCl_{(s)}$

Method of preparation

- Mix silver nitrate solution with sodium chloride solution in a beaker.
- A white precipitate of silver chloride forms.
- Allow the precipitate to settle.
- Filter off the precipitate and wash it with distilled water to remove any amount of sodium nitrate left.
- Dry the precipitate on the filter paper.
- A pure dry sample of silver chloride forms.

Exercise

1. Silver chloride is an insoluble salt.

   (a) Describe how you would prepare a pure dry sample of silver chloride naming the reagents used.

   (b) Write an ionic equation for the reaction involved.

Example 2: Preparation of lead (II) chloride

Reagents

- Lead (II) nitrate, $Pb(NO_3)_2$
- Sodium chloride, NaCl (Alternatively, hydrochloric acid, HCl)

Products

- Lead (II) chloride, $PbCl_2$
- Sodium nitrate, $NaNO_3$

Reaction equation

$Pb(NO_3)_{2(aq)} + 2NaCl_{(aq)} \rightarrow PbCl_{2(s)} + 2NaNO_{3(aq)}$

Ionic equation

$Pb^{2+}{}_{(aq)} + 2Cl^{-}{}_{(aq)} \rightarrow PbCl_{2(s)}$

Method of preparation

- Mix lead (II) nitrate solution with sodium chloride solution in a beaker.
- A white precipitate of lead (II) chloride forms.
- Allow the precipitate to settle.
- Filter off the precipitate and wash it with distilled water to remove any amount sodium nitrate left.
- Dry the precipitate on the filter paper.
- A pure dry sample off lead (II) chloride forms.

Example 3: Preparation of lead (II) iodide

Reagents

- Lead (II) nitrate, $Pb(NO_3)_2$
- Potassium iodide, KI

Products

- Lead (II) iodide, PbI_2
- Potassium nitrate, KNO_3

Reaction equation

$Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \rightarrow PbI_{2(s)} + 2KNO_{3(aq)}$

Ionic equation

$Pb^{2+}_{(aq)} + 2I^{-}_{(aq)} \rightarrow PbI_{2(s)}$

Method of preparation

- Mix lead (II) nitrate solution with potassium iodide solution in a beaker.
- A yellow precipitate of lead (II) iodide forms.
- Allow the precipitate to settle.
- Filter off the precipitate and wash it in distilled water to remove any amount of potassium nitrate left.
- Dry the precipitate on the filter paper.
- A pure dry sample of lead (II) iodide forms.

Example 4: Preparation of lead (II) sulphate

Reagents

- Lead (II) nitrate, $Pb(NO_3)_2$
- Sodium sulphate, Na_2SO_4 (Alternatively, sulphuric acid, H_2SO_4)

Products

- Lead (II) sulphate, $PbSO_4$
- Sodium nitrate, $NaNO_3$

Reaction equation

$Pb(NO_3)_{2(aq)} + Na_2SO_{4(aq)} \rightarrow PbSO_{4(s)} + 2NaNO_{3(aq)}$

Ionic equation

$Pb^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow PbSO_{4(s)}$

Method of preparation

- Mix lead (II) nitrate solution with sodium sulphate solution in a beaker
- A white precipitate of lead (II) sulphate forms.
- Allow the precipitate to settle.
- Filter off the precipitate and wash it with distilled water to remove any amount sodium nitrate left.
- Dry the precipitate on the filter paper.
- A pure dry sample of lead (II) sulphate forms.

Example 5: Preparation of barium sulphate

Reagents

- Barium nitrate, $Ba(NO_3)_2$
- Sodium sulphate, Na_2SO_4 (Alternatively, sulphuric acid, H_2SO_4)

Products

- Barium sulphate, $BaSO_4$
- Sodium nitrate, $NaNO_3$

Reaction equation

$Ba(NO_3)_{2(aq)} + Na_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + 2NaNO_{3(aq)}$

Ionic equation

$Ba^{2+}{}_{(aq)} + SO_4{}^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)}$

Method of preparation

- Mix barium nitrate solution with sodium sulphate solution in a beaker.
- A white precipitate of barium sulphate forms.
- Allow the precipitate to settle.
- Filter off the precipitate and wash it with distilled water to remove any amount of sodium nitrate left.
- Dry the precipitate on the filter paper.
- A pure dry sample of barium sulphate forms.

Exercise

1. Barium sulphate ($BaSO_4$) is an insoluble salt which is prepared by precipitation.

a) Using sodium sulphate as one of the reactants:

(i) Name the other reactant you would use to prepare barium sulphate

(ii) Write a balanced chemical equation for the reaction. Include state symbols.

(iii) Write an ionic equation for the reaction.

b) Briefly explain how you would obtain a fairly pure dry sample of the salt.

Note

Silver chloride, lead (II) chloride, lead (II) iodide, lead (II) sulphate and barium sulphate do not form crystals and therefore cannot be crystallized.

Qualitative analysis tests

Identification of ions.

Test for anions in solution

Anion in solution	**Test method**	**Positive test result**
Carbonate ion, CO_3^{2-}	Add any dilute acid	Effervernsce occurs and bubbles of carbon dioxide gas are produced. $CO_3^{2-}{}_{(aq)} + 2H^+{}_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$
Chloride ion, Cl^-	Add a few drops of dilute nitric acid to make the solution acidified. Then add a few drops of silver nitrate.	A white precipitate of silver chloride forms. $Ag^+{}_{(aq)} + Cl^-{}_{(aq)} \rightarrow AgCl_{(s)}$
Iodide ion, I^-	Add a few drops of dilute nitric acid to make the solution acidified. Then add a few drops of aqueous lead (II) nitrate	A yellow precipitate of lead (II) iodide forms $Pb^{2+}{}_{(aq)} + 2I^-{}_{(aq)} \rightarrow PbI_{2(s)}$
Sulphate ion, SO_4^{2-}	Add a few drops of dilute nitric acid to make the solution acidified. Then add a few drops of barium nitrate.	A white precipitate of barium sulphate forms. $Ba^{2+}{}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)}$
Nitrate ion, NO_3^-	Add a few drops aqueous sodium hydroxide then aluminium foil; warm carefully. Introduce red litmus paper into the test tube.	Moist red litmus paper turns blue. Ammonia gas is produced.

Test for cations in solution

When testing for a cation using either aqueous sodium hydroxide or aqueous ammonia, two observations will help identify the cation present:

- The colour of the precipitate formed on adding a few drops of chemical reagent;
- The solubility of the precipitate in excess chemical reagent

Cation in solution	**Effect of aqueous sodium hydroxide**	**Effect of aqueous ammonia**
Aluminum, Al^{3+}	White ppt, soluble in excess giving a colourless solution	White ppt, insoluble in excess
Ammonium, NH_4^+	Ammonia gas produced on warming.	
Calcium, Ca^{2+}	White ppt insoluble in excess	No ppt or very slight white ppt
Copper (II), Cu^{2+}	Light blue ppt, insoluble in excess	Light blue ppt, soluble in excess giving a dark blue solution
Iron (II), Fe^{2+}	Green ppt, insoluble in excess	Green ppt, insoluble in excess
Iron (III),Fe^{3+}	Red-brown ppt, insoluble in excess	Red-brown ppt insoluble in excess
Zinc, Zn^{2+}	White ppt, soluble in excess giving a colourless solution	White ppt, soluble in excess giving a colourless solution

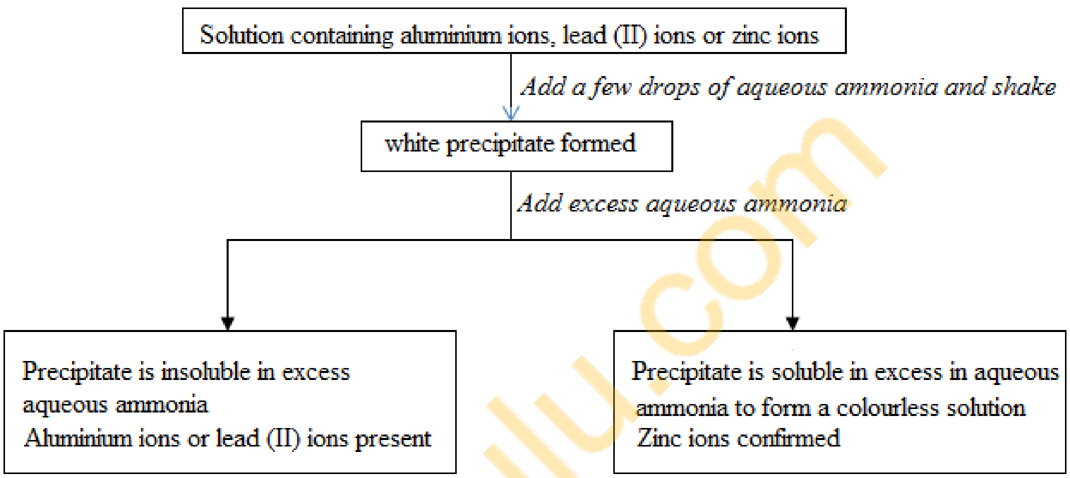
The cations react with hydroxide ions present in aqueous sodium hydroxide or ammonia to form insoluble hydroxides. These hydroxides appear as precipitates.

Example: $Fe^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Fe(OH)_{2(s)}$

Some of these precipitates dissolve in excess aqueous sodium hydroxide or aqueous ammonia to form soluble complex salts. These appear as colourless solutions, or in the case of copper (II) ions in excess aqueous ammonia, a dark blue solution.

Copper (II), iron (II) and iron (III) ions are easily identified by the characteristic colour of their precipitates.

Aluminium, lead (II) and zinc ions all give the same observations when aqueous sodium hydroxide is used. However, only zinc ions will give a white precipitate soluble in excess aqueous ammonia; aluminium and lead ions do not.



To distinguish between aluminium and lead (II) ions, dilute hydrochloric acid *(HCl)* or aqueous potassium iodide can be used:

$$Al^{3+}_{(aq)} \quad + \quad 3Cl^{-}_{(aq)} \quad \rightarrow \quad AlCl_{3(aq)}$$

(from HCl) *(colourless solution)*

$$Pb^{2+}_{(aq)} \quad + \quad 2Cl^{-}_{(aq)} \quad \rightarrow \quad PbCl_{2(s)}$$

(from HCl) *(white precipitate)*

Similar results will be obtained if potassium iodide is used. Aluminium ions will give a colourless solution of aluminium iodide while lead (II) ions will give a yellow precipitate of lead (II) iodide.

Test for gases

When recording observations for gases, it is important to record

- Presence of effervescences, if any
- Colour and smell of the gas
- Chemical test for the gas and test result
- Name of the gas

Gas	Colour and smell of gas	Test method	Positive test result
Oxygen, O_2	Colourless Odourless	Introduce a glowing splint in the test tube containing oxygen.	Relights a glowing splint
Hydrogen, H_2	Colourless Odourless	Introduce a burning splint in the test tube containing hydrogen.	Burns with a pop sound.
Carbon dioxide, CO_2	Colourless Odourless	Bubble carbon dioxide into lime water (Calcium hydroxide)	Turns lime water milky
Ammonia, NH_3	Colourless Pungent smell	-	Turns damp red litmus paper blue.
Chlorine, Cl_2	Yellow green Pungent smell	-	Turns most blue litmus paper red, then bleached. Bleaches the moist blue litmus paper.
Sulphur dioxide, SO_2	Colourless Pungent smell	-	Turns moist blue litmus red Turns orange acidified potassium dichromate (VI) green

Oxidation and Reduction

Oxidation and reduction can be defined in terms of oxygen, hydrogen, electrons and oxidation number.

[A] Oxidation
Oxidation in terms of oxygen

Oxidation is the addition of oxygen to a chemical substance.

Examples

(I) $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

Magnesium has been oxidized to magnesium oxide since oxygen has been added

(II) $4FeO_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_{3(s)}$

Iron (II) oxide (FeO) has been oxidized to iron (III) oxide (Fe_2O_3) since oxygen has been added.

Oxidation in terms of hydrogen

Oxidation is the removal of hydrogen from a chemical substance.

Examples

(I) $2NH_{3(g)} \rightarrow N_{2(g)} + 3H_{2(g)}$

Ammonia (NH_3) has been oxidized to nitrogen (N_2) since hydrogen has been removed.

(II) $H_2S_{(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)} + S_{(s)}$

Hydrogen sulphide (H_2S) has been oxidized to sulphur (S) since hydrogen has been removed.

Oxidation in terms of electrons.

Oxidation is the loss of electrons from a chemical substance.

Examples

(I) $Cu \rightarrow Cu^{2+} + 2e$

Copper atoms (Cu) have been oxidized to copper (II) ions (Cu^{2+}) since two electrons have been lost.

(II) $Al \rightarrow Al^{3+} + 3\tilde{e}$

Aluminum atoms (Al) have been oxidized to aluminum ions (Al^{3+}) since three electrons have been lost.

(III) $2Cl^- \rightarrow Cl_2 + 2\tilde{e}$

Chloride ions (Cl^-) have been oxidized to chlorine molecules (Cl_2) since electrons have been lost.

Oxidation number.

Rules to consider when assigning the oxidation number

1. The oxidation number of neutral particles like atoms and molecules is equal to Zero.

Examples

Cu^0	Oxidation number = 0	Al^0	Oxidation number = 0	Cl_2^0	Oxidation number = 0

Na^0	Oxidation number = 0	Fe^0	Oxidation number = 0	H_2^0	Oxidation number = 0
Mg^0	Oxidation number = 0	Zn^0	Oxidation number = 0	O_2^0	Oxidation number = 0

2. The oxidation number of hydrogen in all compounds except metallic halides is +1
3. The oxidation number of oxygen in all compounds except in peroxides and in OF_2 is -2.
4. For some metals the oxidation number is equal to the group number on the periodic table or the valence.
5. The oxidation number of a simple ion is equal to the charge it carries.

Examples

Ion	**Formula**	**oxidation number**
Aluminium ion	Al^{3+}	+3
Calcium ion	Ca^{2+}	+2
Chloride ion	Cl^{-}	-1
Nitride ion	N^{3-}	-3
Oxide ion	O^{2-}	-2
Sodium ion	Na^{+}	+1

6. In neutral compounds, the sum of individual elements present is equal to zero.

Examples

a) Find the oxidation number of

(i) S in $\underline{S}O_2$

(ii) S in $H_2\underline{S}O_4$

Solution

(i) x + (-2) 2 = 0

x + (-4) = 0

x – 4 = 0

x = +4

(ii) (1x2) + x + (-2) 4 =0

2 + x – 8 = 0

x – 6 = 0

x = +6

7. The sum of all oxidation numbers of all elements in a complex ion is equal to the charge on the ion.

Examples

a) Find the oxidation number of

(i) C in $H\underline{\textbf{C}}O_3^-$

(ii) S in $\underline{\textbf{S}}O_4^{2-}$

Solution

(i) $1 + x + (-2)3 = -1$

$x + 1 - 6 = -1$

$x - 5 = -1$

$x = +5 - 1$

$x = +4$

(ii) $x + (-2)4 = -2$

$x - 8 = -2$

$x = +8 - 2$

$x = +6$

Oxidation in terms of oxidation number

Oxidation is the increase in the oxidation number of a substance.

Examples

(I) $Cu^o \rightarrow Cu^{2+} + 2\tilde{e}$

Copper atoms (Cu) have been oxidized to copper (II) ions (Cu^{2+}) because the oxidation number has increased from 0 in Cu to +2 in Cu^{2+}.

(II) $Cl^- \rightarrow Cl^o + \tilde{e}$

A Chloride ion (Cl^-) has been oxidized to chlorine atom (Cl) since the oxidation number has increased from −1 in Cl^- to 0 in Cl.

(III) $S^0_{2(g)} + O^0_{2(g)} \rightarrow S^{+4}O^{-2}_{2(g)}$

Sulphur (S) has been oxidized to sulphur dioxide (SO_2) since the oxidation number has increased from 0 in S to +4 in SO_4.

Oxidizing agents

Alternative term: Oxidants

Definition: An oxidizing agent is a chemical substance which brings about oxidation of another substance but end up being reduced.

Examples of oxidizing agents.

- Oxygen, O_2
- Chlorine, Cl_2
- Hydrogen peroxide, H_2O_2
- Potassium permanganate, $KMnO_4$

- Manganese dioxide (manganese (IV) oxide), MnO_2
- Concentrated Sulphuric acid, H_2SO_4

Characteristics of an oxidizing agent.

An oxidizing agent: -

(a) Supplies or donates oxygen to another substance

Example

$PbO_{(s)} + H_{2(g)} \rightarrow Pb_{(s)} + H_2O_{(g)}$

Lead (II) oxide (PbO) is an oxidizing agent because it has donated oxygen to hydrogen

(b) Removes hydrogen from another substance

Example

$H_2S_{(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)} + S_{(s)}$

Chlorine (Cl_2) is an oxidizing agent because it has removed hydrogen from hydrogen sulphide (H_2S)

(c) Accept electrons from another substance.

(d) Increases the oxidation number of another substance.

Example

$Zn^o + Cu^{2+} \rightarrow Zn^{2+} + Cu^o$

Copper (II) ion (Cu^{2+}) is an oxidizing agent because it has caused the increase in the oxidation number of Zinc from 0 in Zn to +2 in Zn^{2+}.

Test for oxidizing agents.

An oxidizing agent can be tested by using acidified potassium iodide, KI or concentrated hydrochloric acid.

Oxidizing agents liberate iodine from a solution of acidified potassium iodide.

They turn moist starch iodide paper from white to blue. Starch iodide paper contains starch and potassium iodide. The iodide ion is oxidized to form iodine.

$2I^-_{(aq)} \rightarrow I_{2(s)} + 2\tilde{e}$

The iodine reacts with starch to produce a blue colour.

[B] Reduction

Reduction in terms of oxygen

Reduction is the removal of oxygen from a chemical substance.

Examples

(a) $CO_{2(g)} + C_{(s)} \rightarrow 2CO_{(g)}$

Carbon dioxide (CO_2) has been reduced to carbon monoxide (CO) since oxygen has been removed.

(b) $PbO_{(s)} + H_{2(g)} \rightarrow Pb_{(s)} + H_2O_{(g)}$

Lead (II) oxide (PbO) has been reduced to lead (Pb) by the removal of oxygen.

(c) $3CO_{(g)} + Fe_2O_{3(s)} \rightarrow 2Fe_{(s)} + 3CO_{2(g)}$

Iron (III) oxide (Fe_2O_3) has been reduced to iron (Fe) by the removal of oxygen.

Reduction in terms of hydrogen.

Reduction is the addition of hydrogen to a chemical substance**.**

Examples

(a) $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$

Nitrogen (N_2) has been reduced to ammonia (NH_3) since hydrogen has been added.

(b) $C_2H_{4(g)} + H_{2(g)} \rightarrow C_2H_{6(g)}$

Ethene (C_2H_4) has been reduced to ethane (C_2H_6) by the addition of hydrogen

(c) $2C_{(s)} + H_{2(g)} \rightarrow C_2H_{2(g)}$

Carbon (C) has been reduced to ethyne (C_2H_2) by the addition of hydrogen.

Reduction in terms of electrons

Reduction is the gain of electrons.

Examples

(a) $Cu^{2+} + 2\tilde{e} \rightarrow Cu$

Copper (II) ions (Cu^{2+}) have been reduced to copper atoms (Cu) by gaining two electrons.

(b) $Cl_2 + 2\tilde{e} \rightarrow 2Cl^-$

Chlorine molecules (Cl_2) have been reduced to chloride ions (Cl^-) by gaining two electrons.

Reduction in terms of oxidation number

Reduction is the decrease in the oxidation number of a chemical substance.

Examples

(a) $Zn^{2+} + 2\,\tilde{e} \rightarrow Zn^o$

Zinc ions (Zn^{2+}) have been reduced to Zinc atoms (Zn) since the oxidation number has decreased from +2 in Zn^{2+} to 0 in Zn.

Reducing agents

Alternative term: Reductants

Definition: A reducing agent is a chemical substance which brings about reduction but end up being oxidized.

Examples of reducing agents.

- Hydrogen, H_2
- Carbon monoxide, CO
- Carbon, C
- Ammonia, NH_3
- Hydrogen sulphide, H_2S
- Sulphur dioxide, SO_2

Characteristics of a reducing agent.

A reducing agent: -

(a) Removes oxygen from another substance

Example

$3CO_{(g)} + Fe_2O_{3(s)} \rightarrow 2Fe_{(s)} + 3CO_{2(g)}$

Carbon monoxide (CO) is a reducing agent because it has removed oxygen from iron (III) oxide (Fe_2O_3)

(b) Donates or supplies hydrogen to another sustenance

Example

$H_2S_{(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)} + S_{(s)}$

Hydrogen sulphide (H_2S) is a reducing agent because it has donated hydrogen to chlorine.

(c) Donates electrons to another substance

(d) Decreases the oxidation number of another substance

Example

$Zn^o + Cu^{2+} \rightarrow Zn^{2+} + Cu^o$

Zinc atom (Zn) is a reducing agent because it has caused the decrease in the oxidation number of copper (II) ions from +2 in Cu^{2+} to 0 in Cu.

Test for reducing agents.

Reducing agents can be tested by using acidified potassium per manganate or acidified potassium dichromate (VI)

1. Reducing agents change the colour of the solution of potassium permanganate (VII) from purple to colourless.

$MnO_4^-{}_{(aq)} + 8H^+{}_{(aq)} + 5\tilde{e} \rightarrow Mn^{2+}{}_{(aq)} + 4H_2O_{(l)}$

(Purple) (Colourless)

2. Reducing agents change a solution of acidified potassium dichromate (VI) from orange to green.

$Cr_2O_7^{2-}{}_{(aq)} + 14H^+{}_{(aq)} + 6\tilde{e} \rightarrow 2Cr^{3+} + 7H_2O_{(l)}$

(Orange) (Green)

In both tests, the electrons come from the reducing agent.

Redox reactions

Reduction and oxidation reactions are called redox reactions because they occur simultaneously. When one substance is reduced, the other one is oxidized.

Examples of redox reactions

Reduction

(I) $Fe_2O_{3(s)} + 2Al_{(s)} \rightarrow Al_2O_{3(s)} + 2Fe$

Oxidation

Iron (III) oxide (Fe_2O_3) has been reduced to iron (Fe) while aluminium (Al) has been oxidized to aluminium oxide, Al_2O_3.

Reduction

(II) $PbO_{(s)} + H_{2(g)} \rightarrow Pb_{(s)} + H_2O_{(l)}$

Oxidation

Lead (II) oxide (PbO) has been reduced to lead (Pb) while hydrogen (H_2) has been oxidized to water (H_2O).

Oxidation

(III) $2CuO_{(s)} + 2NH_{3(g)} \rightarrow 2Cu_{(s)} + 2N_{2(g)} + 2H_2O_{(l)}$

Reduction

Copper (II) oxide (CuO) has been reduced to copper (Cu) while ammonia (NH_3) has been oxidized to nitrogen (N_2)

Oxidation

(IV) $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$

Reduction

Zinc (Zn) has been oxidized to zinc ion (zn^{2+}) while copper (II) ion (Cu^{2+}) has been reduced to copper (Cu).

Oxidation

(V) $H_2S_{(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)} + S_{(s)}$

Reduction

Hydrogen sulphide (H_2S) has been oxidized to sulphur (S) while chlorine molecules (Cl_2) have been reduced to hydrogen chloride (HCl)

Exercise

1. Which of the following reactions is oxidation and which is reduction?
   (a) $Al^{3+} + 3\tilde{e} \rightarrow Al$
   (b) $Cl_2 + 2\tilde{e} \rightarrow 2Cl^-$
   (c) $2O^{2-} \rightarrow O_2 + 2\tilde{e}$
   (d) $Fe^{2+} \rightarrow Fe^{3+} + \tilde{e}$
   (e) $Cu^+ + \tilde{e} \rightarrow Cu$
2. State whether the underlined substance has been oxidized or reduced.
   (a) $\underline{CO_2} + C \rightarrow 2CO$
   (b) $2FeO_3 + \underline{2Al} \rightarrow Al_2O_3 + 2Fe$
   (c) $2CuO + \underline{2NH_3} \rightarrow 2Cu + 3N_2 + 2H_2O$
3. Indicate reduction and oxidation in the following redox reactions.

(a) $Ag_2O_{(s)} + H_{2(g)} \rightarrow 2Ag_{(s)} + H_2O_{(l)}$

(b) $CO_{(g)} + CuO_{(s)} \rightarrow Cu_{(s)} + CO_{2(g)}$

(c) $Zn_{(s)} + Pb(NO_3)_{2(aq)} \rightarrow Pb_{(s)} + Zn(NO_3)_{2\,(aq)}$

(d) $MnO_{2(s)} + 4HCl_{(aq)} \rightarrow MnCl_{2(aq)} + 2H_2O_{(l)} + Cl_{2(g)}$

(e) $Zn^{2+}{}_{(aq)} + Mg_{(s)} \rightarrow Mg^{2+}{}_{(aq)} + Zn_{(s)}$

4. Find the oxidation number of:
   (a) Manganese in MnO_4^-
   (b) Carbon in H_2CO_3
   (c) Chromium in $Cr_2O_7^{2-}$

Rates of chemical reactions

Definition: Rate of chemical reaction is the speed at which the reaction takes place

Formula: Rate of reaction = $\frac{\text{change in amount of reactants/products}}{\text{time taken}}$

Measuring the rate of reaction

The rate of reaction can be measured by measuring:

- how quickly a product is obtained
- how quickly a reactant is used up

Example

Consider the reaction below;

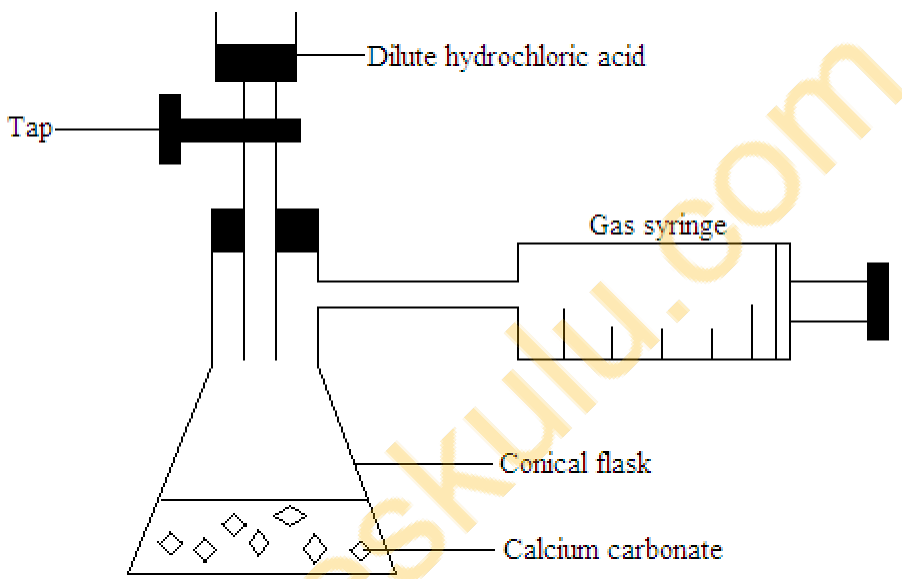
$CaCO_{3(s)}$ + $2HCl_{(aq)}$ $\rightarrow$ $CaCl_{2(aq)}$ + $H_2O_{(l)}$ + $CO_{2(g)}$

In the reaction above, the rate of reaction can be measured by measuring:

- The volume of carbon dioxide over time
- The decrease in mass of the system due loss of carbon dioxide

[A] Measuring the rate of reaction by measuring the volume of the gas produced

A graduated syringe is used to measure the volume of carbon dioxide gas formed over time



The total volume of carbon dioxide given off at one minute interval is recorded

The graph of total volume of carbon dioxide against time is plotted

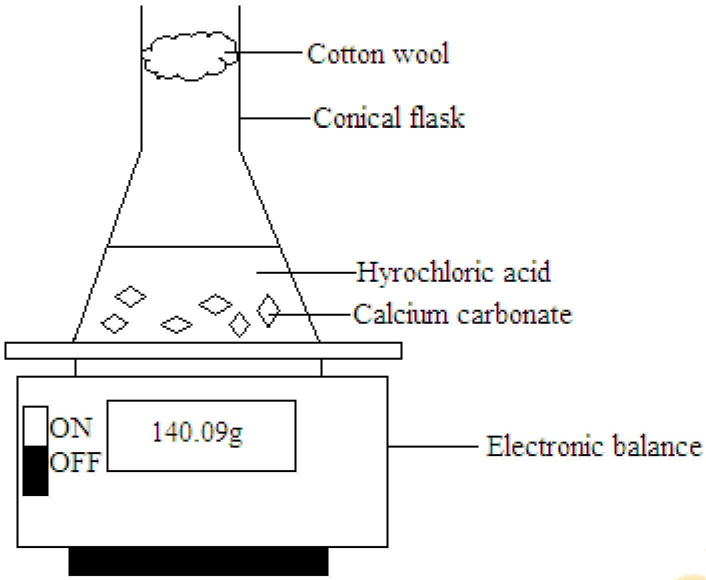
The gradient of the graph is calculated

The gradient of the graph is equal to the rate of reaction.

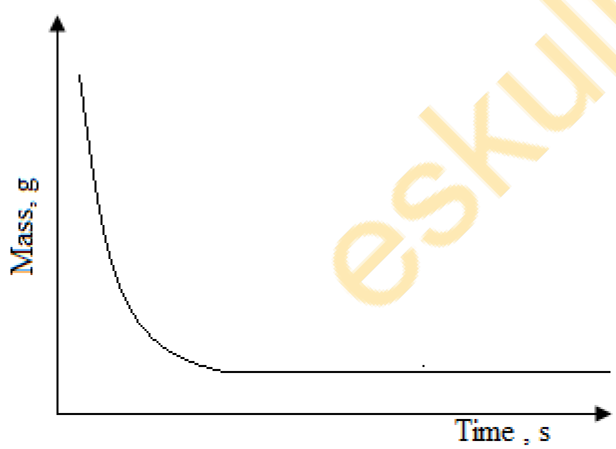
Rate of reaction = $\frac{\text{volume of carbon dioxide produced}}{\text{time taken}}$

[B] Measuring the rate of reaction by measuring the decrease in mass of a system due to loss of product

A mass balance is used to follow the loss in mass of a system



The mass readings will drop over time as the carbon dioxide gas formed escapes. The mass readings are taken at one minute intervals and plotted against time



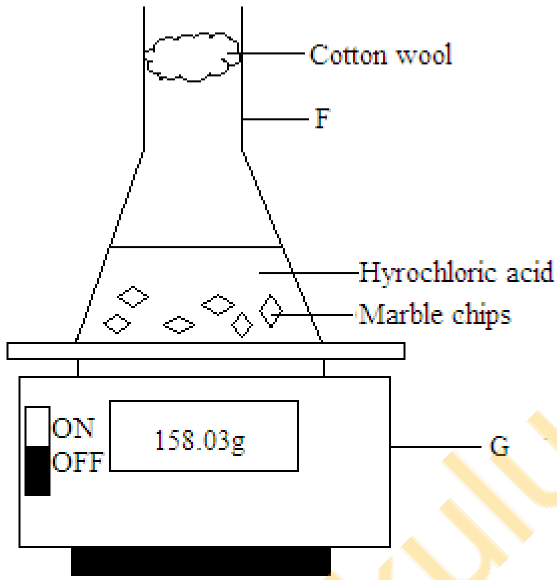
The gradient of the graph at various points of the curve will give the rate of reaction. The reaction is fastest at the start because the gradient of the graph is the highest. The value of the gradient decreases with time and finally becomes zero. This means that as the reaction proceeds, the reaction slows down and finally comes to a stop.

Note

- The cotton wool is used as a stopper. It will allow the escape of carbon dioxide into the atmosphere and prevent the solution inside the conical flask from splashing out.
- The mass decreases as the reaction proceeds because of the loss and escape of carbon dioxide into the atmosphere
- When the curve levels off, it means the reaction has stopped

Exercise

1. The rate of reaction between dilute hydrochloric acid and marble chips (calcium carbonate) was investigated using the following apparatus



(a) Name the piece of apparatus labeled F and G
(b) Why was the cotton wool placed in the neck of the apparatus F?
(c) Name two compounds left in the solution at the end of the experiment
(d) Why did the mass decrease as the reaction proceeds?

The collision theory

The collision theory states that: For a reaction to take place, the particles of the reacting substances must move and collide with each other with a certain amount of kinetic energy.
The number of collisions taking place per unit time depends on the number of particles.
If the particles are increased, the number of collisions also increases.

Factors affecting the rate of reaction

- Concentration
- Temperature
- Pressure

- Surface area (size of particles)
- Catalyst

1. **Effects of concentration on the rate of reaction**

Concentration refers to the reactants in solution.

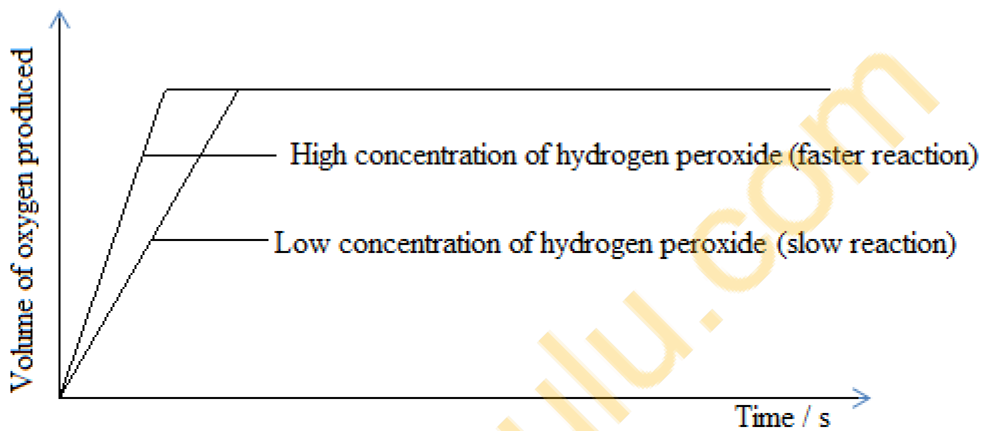
When concentration is increased, the rate of reaction also increases. This is because the number of particles in the solution increases and collides with each other effectively.

On the other hand, when concentration is reduced, the rate of reaction also reduces. This is because the number of particles in the solution reduces and do not collide effectively.

Graphical representation of concentration

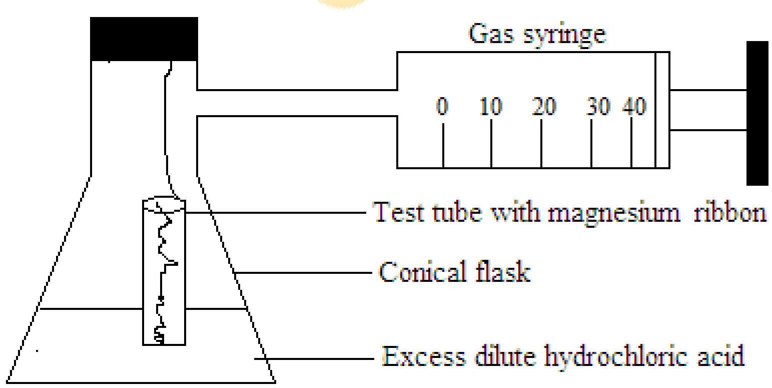
Example: Decomposition of hydrogen peroxide.

$$2H_2O_{2(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$$



Exercise

1. The rate of reaction between magnesium ribbon and an excess of dilute hydrochloric acid could be measured using the apparatus shown below.



(a) What was the purpose of the test tube?

(b) How do you get the reaction to start?

(c) The volume of hydrogen produced was recorded every minute as shown in the table below

Time / minutes	0	1	2	3	4	5	6	7
Volume of hydrogen gas / cm^3	0	14	23	31	38	40	40	40

(I) Plot the results on the graph and draw the graph

(II) What was the total volume of hydrogen produced when the reaction was over

(III) Why did the reaction stop?

(IV) How do you make the reaction go faster?

2. Effects of temperature on the rate of reaction

Temperature is the measure of the average kinetic energy of the particles

When temperature is increased, the rate of reaction also increases. This because the particle gain kinetic energy and move faster and collide effectively

On the other, when temperature is reduced, the rate of reaction also reduces. This is because the particles lose kinetic energy and move slower and do not collide effectively

Note

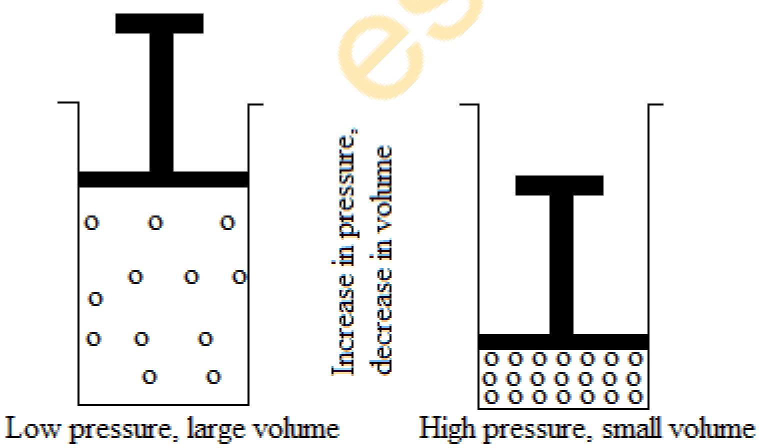
Temperature increases the rate of reactions for endothermic reactions

3. Effects of pressure on the rate of reaction

Pressure becomes a dominant factor in reactions involving gases

When pressure is increased, the rate of reaction also increases. This is because the volume reduces forcing the gas particles closer together and collides effectively

On the other hand, when pressure is reduced, the rate of reaction also reduces. This is because the volume increases and the gas particles are further apart and do not collide effectively



4. Effects of surface area (particle size) on the rate of reaction

Particle size usually refers to particles of a solid reactant.

The rate of reaction is faster when the size of particles is small. This is because a small sized particle has a large surface area.

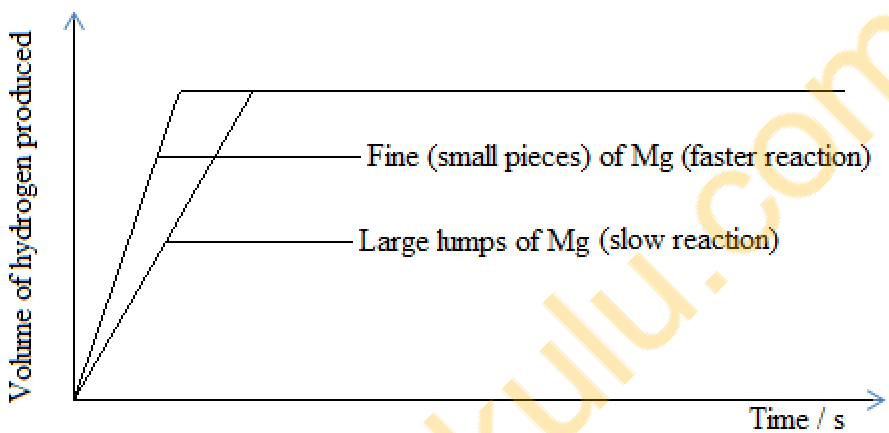
Note

- When a reactant is in solid state, the reaction takes place on the surface of the solid. By breaking up the solid into smaller pieces, the surface area is increased giving a greater area for collisions to occur. This results in an increase of the rate of reaction.
- This explains why mixtures of saw dust, fine products of flour mills and combustion of gases can cause an explosion due to the large surface area.

Graphical representation of surface area

Example: Reaction of magnesium with dilute hydrochloric acid.

$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

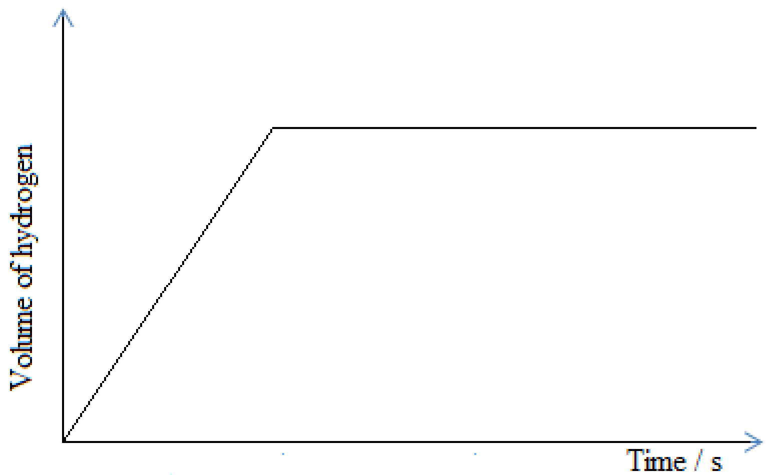


Note

The reaction stops when the reactant in fewer quantities is used up, and hence curve levels off.

Exercise

1. A lump of marble was used to react with dilute hydrochloric acid.
   (a) What is the chemical name for marble?
   (b) State three different ways in which the reaction could have been made more rapidly
   2. The diagram below shows the volume of hydrogen produced over a period of time when 4g of calcium pellets are carefully added to an excess of dilute hydrochloric acid at 25^oC



(a) Draw on the same diagram the curve you would expect if the experiment were repeated at the same temperature with:

(I) 4g of calcium pellets but with excess acid of half the concentration of the original acid. Label this curve A

(II) Excess acid of the original concentration but with 4g of powdered calcium. Label this curve B

(b) Write a balanced chemical equation for the reaction including state symbols

(c) Write an ionic equation for the reaction in (b) above.

3. The table below shows the results of the production of hydrogen in the reaction o between zinc granules and dilute hydrochloric acid at 25^{o}C.

Volume (cm^3)	0	5.0	10.0	13.0	15.5	18.5	19.5	19.5	19.5
Time (Min)	0	0.6	1.5	2.3	3.0	4.4	5.7	7.0	7.5

(a) Plot a graph of volume against time for the reaction

(b) Calculate the average rate of the reaction during the first 3 minutes

(c) Use the graph to determine the volume of gas collected after 5 minutes

(d) On the same graph, sketch a curve you would expect if powdered zinc were used instead of zinc granules. Label this curve P

(e) How long did it take for the reaction to end?

(f) Mention factors you would employ to increase the rate of production of the gas

(g) Explain briefly how each of the mentioned factors above increases the rate of production of the gas.

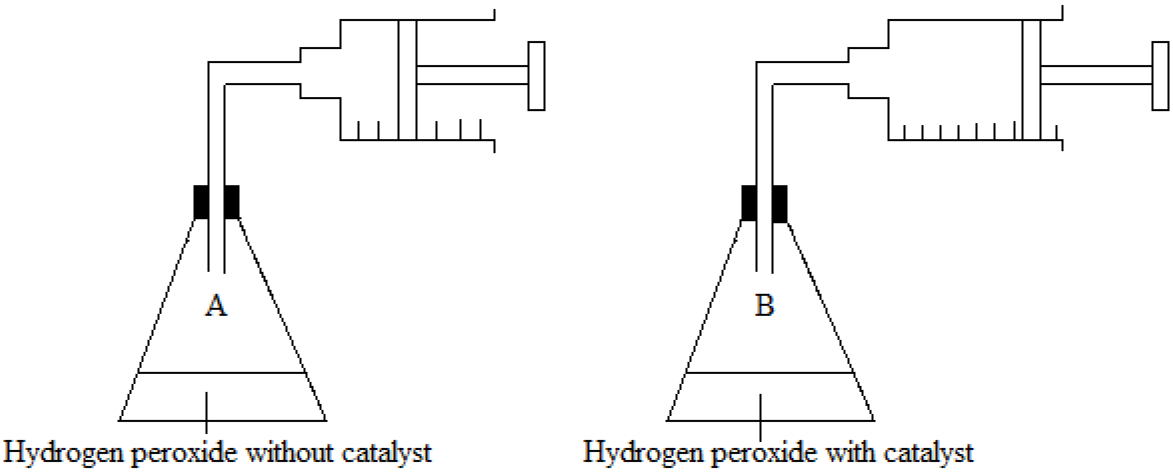
5. Effects of a catalyst on the rate of reaction

A catalyst is a chemical substance which alters the rate of reaction but remains chemically unchanged at the end of the reaction

A catalyst usually speeds up the reaction by lowering the activation energy of the reaction.

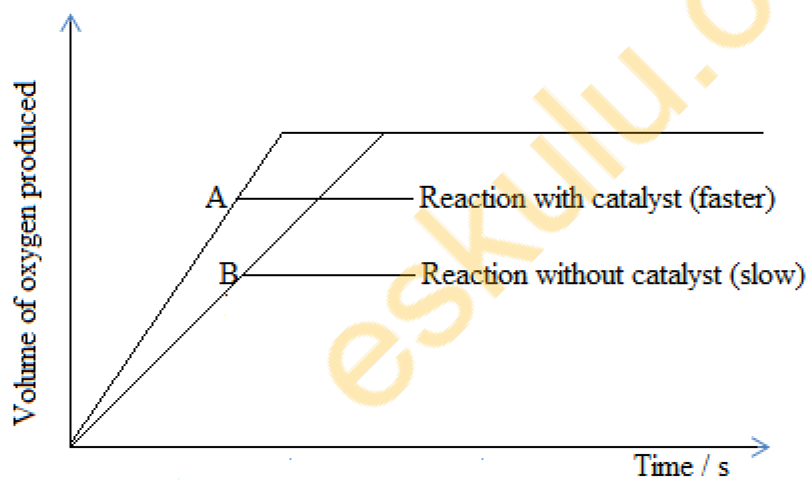
For example, the volume of oxygen produced from the decomposition of hydrogen peroxide (H_2O_2) can be measured using a gas syringe with a catalyst manganese (IV) oxide, a black solid.

The addition of manganese (IV) oxide (MnO_2) speeds up the reaction and increases the volume of oxygen formed within a short time.



($0.1\ mol/dm^3 H_2O_2$ without MnO_2catalyst)($0.1\ mol/dm^3 H_2O_2$ with MnO_2catalyst)

Graphical representation of a catalyst



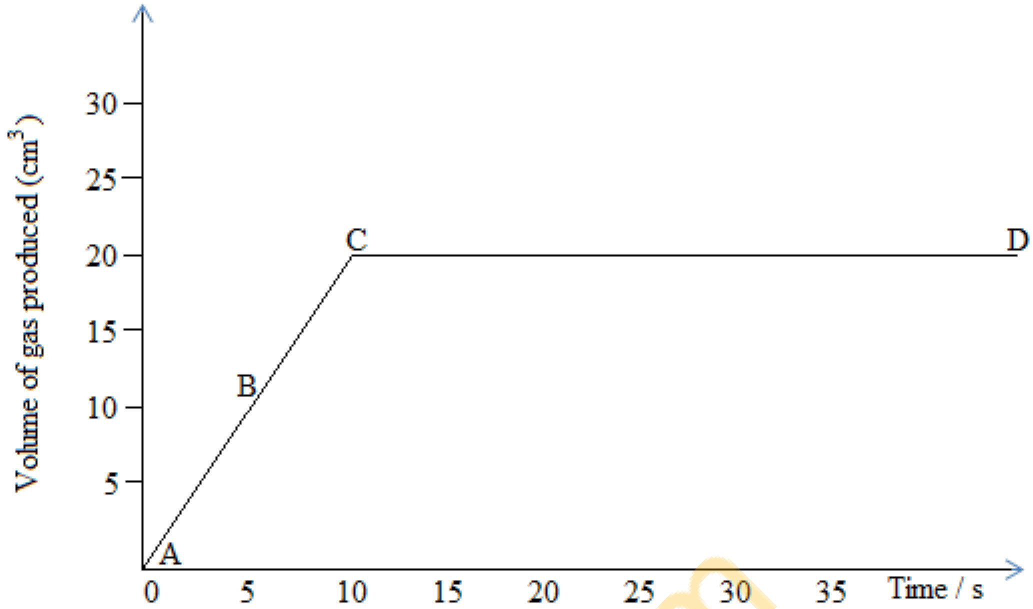
Exercise

1. The decomposition of aqueous hydrogen peroxide solution is catalyzed by a black solid. A gas is evolved.

(a) Name the gas

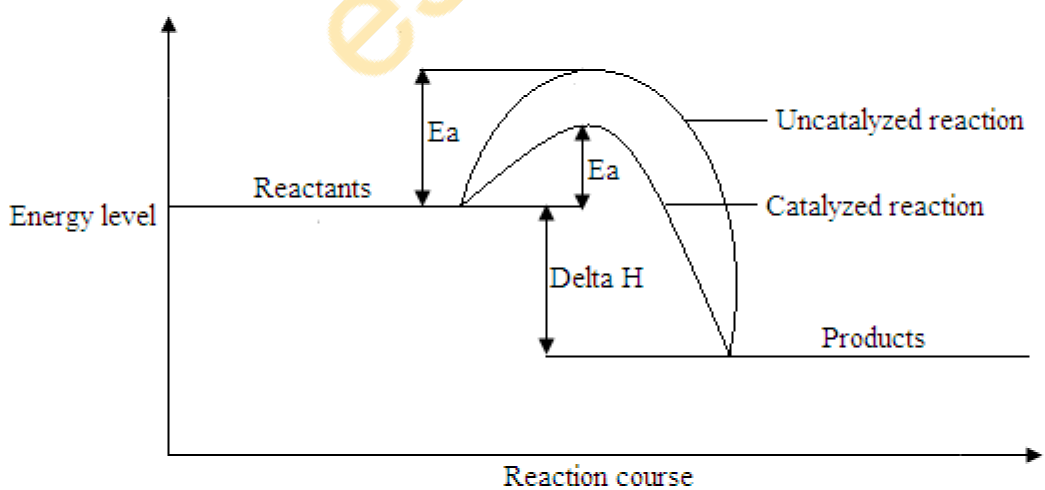
(b) What is a catalyst and name the catalyst suitable for this reaction?

(c) The small amount of catalyst is mixed with $20cm^3$ of hydrogen peroxide. The total volume of gas produced is measured in every 5 seconds. The results are plotted as shown below



(I) What is the final volume of the gas produced?
(II) Why does the curve level off after 10 seconds?
(III) Choose at which point where the reaction is fastest A, B, C or D?

Activation energy
Symbol: E_a
Definition: Activation energy is the minimum energy required to start a reaction. As a result, a catalyst allows a reaction to go by a different pathway with lower activation energy allowing more collisions for a successful reaction.



Activation energy is usually the energy barrier because if any collision is not energetic enough, the reaction will be futile

Some catalysts slow down the reactions; these are called inhibitors (negative catalysts)

Characteristics of a catalyst

1. It catalyzes both the forward and reverse reaction
2. It undergoes physical change
3. It remains chemically unchanged at the end of the reaction
4. It is only needed in very small amounts
5. It is poisoned or rendered useless in the presence of impurities

Examples of some catalysts used for important reactions

Catalyst	**Reaction catalyzed**
Aluminium oxide	Cracking of alkanes
Iron	Haber process: Production of ammonia
Manganese (IV) oxide	Decomposition of hydrogen peroxide
Vanadium pentoxide	Contact process: Manufacture of sulphuric acid
Nickel	Hydrogenation of vegetable oils

Chemical equilibrium

Equilibrium is the point where the rate of forward reaction is equal to the rate of reverse reaction.

Reversible reactions

Symbol: $\rightleftharpoons$

Definition: A reversible reaction is a reaction that proceeds in the forward and backward directions

It follows that if the reaction is exothermic, then the reverse reaction will be endothermic and vice versa

Example: $A + B \rightleftharpoons C + D$

The double headed arrow shows that the reaction is a reverse reaction

Characteristics of equilibrium reactions

1. Rate of forward reaction is equal to rate of backward reaction
2. Concentration of reactants and products remain constant at equilibrium
3. An equilibrium can only be established in a closed system
4. There is no loss or gain of external materials by the reacting system

Le Chartelier's Principle

The principle states that: When a chemical equilibrium is externally disturbed by a change in one of the factors upon which it depends, the equilibrium will shift in a direction so as to offset the change.

Factors that affect the position of equilibrium

1. **Temperature changes**

Temperature variations changes the position of the equilibrium of either endothermic or exothermic reactions

An increase in temperature favors the forward reaction of endothermic reaction while a decrease in temperature will shift the equilibrium backwards

Example

$NH_4Cl_{(s)} \rightleftharpoons NH_{3(g)} + HCl_{(g)}$

In thermal decomposition of ammonium chloride, temperature increase cause equilibrium to shift to the right producing ammonia and hydrogen chloride. On the other hand, when temperature is reduced, equilibrium shifts to the left and the backward reaction is favored producing ammonium chloride

2. **Concentration changes**

Generally, an increase in concentration of the reactants of an equilibrium reaction favors the forward reaction. This is because the equilibrium will adjust itself so as to offset the effect of adding more reactants. On the other hand, if the concentration of products is increased, the backward reaction will be favored so that the reactants are produced to restore the balance

3. **Pressure changes**

In gaseous systems, an equilibrium reaction is followed by volume change. Therefore, the equilibrium is affected by change in pressure.

Example

$N_{2(g)} + 3N_{2(g)} \rightleftharpoons 2NH_{3(g)}$

In the production of ammonia, there is a general decrease in volume and a consequent increase in pressure. Therefore, an increase in pressure will make the equilibrium shift towards the reduction of volume. This means the forward reaction is favored and more ammonia is produced. Consequently, a decrease in pressure leads to the production of nitrogen and hydrogen; it favors the backward reaction

Generally, gaseous reactions that lead to reduction in volume are favored by high pressure.

Exercise

1. The gas ammonia is made industrially by Haber process. In process nitrogen and hydrogen gases are reacted together, as in the equation

   $N_{2(g)} + H_{2(g)} \rightleftharpoons NH_{3(g)}$

   (a) Balance the equation
   (b) What does the notation $\rightleftharpoons$ show about the reaction?
   (c) Ammonia is removed from the reaction mixture by cooling until it condenses. How does this affect the formation of ammonia
   (d) How does an increase in pressure affect the position of equilibrium?
   (e) State Le Chartelier's principle
   (f) Mention any two characteristics of an equilibrium reaction

4. **Catalysts**

Catalysts help to attain equilibrium quickly by increasing both the forward and backward reactions the same way. Therefore, a catalyst does not affect the position of equilibrium.

Energy changes

Alternative term: Energetics

Energetics refers to the energy changes that characterize chemical reactions

Terms related to energy changes

Enthalpy

Symbol: H

Definition: This is the total energy in one mole of a substance

It is also the heat content of a reacting system

It depends on the physical state of the compound and varies from compound to compound

Enthalpy change of reaction

Symbol: ΔH (delta H)

Formula: $\Delta H = H_{products} - H_{reactnats}$

Units: Kilojoules per mole, Kj/mol

Definition: This is the difference between the enthalpy of the products and the enthalpy of the reactants

Endothermic reaction

This is a reaction which absorbs heat energy from the surroundings

Examples of endothermic reactions

(a) Photosynthesis

$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$

(b) Photography in which light energy helps to decompose silver salt to silver on the photographic plate. The essential reaction in photography is the reduction of silver ions to metallic silver. With exposure to light energy, the silver salts decompose into ions as follows:

$AgBr \rightarrow Ag^{+} + Br^{-}$

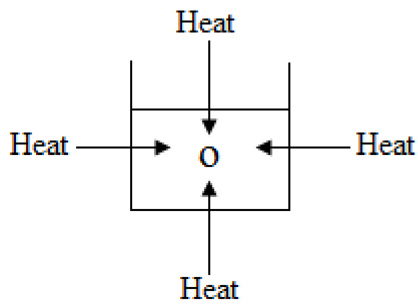
$Ag^{+} + e \rightarrow Ag$

Light is usually absorbed to dissociate silver salts

(c) Dissolving processes such as dissolving ammonium nitrate or potassium nitrate in water.

In endothermic reaction;

- Energy is absorbed from the surroundings



- Reactants have less energy than products
- ΔH is positive
- Temperature of the system falls

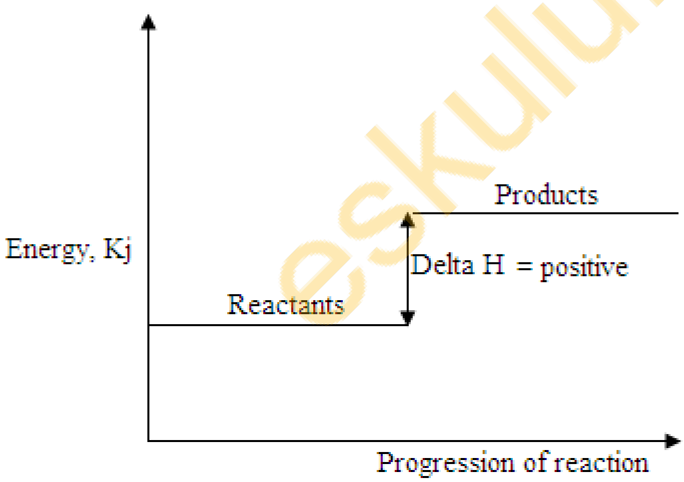
Endothermic reaction as a bond breaking process

It generally requires the absorption of heat energy from the surroundings, hence endothermic.
For example, decomposition of ammonia gas into nitrogen and hydrogen gas.

$2NH_{3(g)} \rightleftharpoons N_{2(g)} + 3N_{2(g)}$ $\Delta H = + 46.5$ Kj / mol

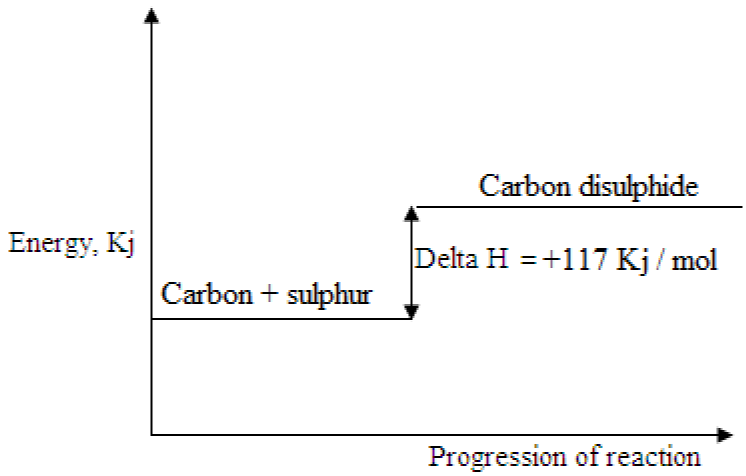
The positive sign in front of the energy value indicate that the reaction is endothermic
Enough amount of heat is required to break the bonds of ammonia gas.
Therefore, endothermic reactions involve bond breaking

Energy level diagram for endothermic reactions



Example of energy level diagram for the formation of carbon disulphide

$$C_{(s)} + 2S_{(s)} \rightarrow CS_{2(s)} \; \Delta H = +117 \text{ Kj/mol}$$



Exercise

1. Consider the reaction below

   $CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$ $\Delta H = +187$ kj/mol

   (a) State whether the reaction is exothermic or endothermic. Give a reason

   (b) Draw the energy level diagram for this reaction

   (c) When calcium oxide reacts with carbon dioxide, will heat be given out or taken in?

   (d) Calculate the energy change when 88g of carbon dioxide reacts with excess of calcium oxide

Exothermic reaction

This is a reaction in which energy is released to the surroundings

Examples of exothermic reactions

(a) All combustion or burning processes e.g. burning of fuels such as coal, oils, wood.

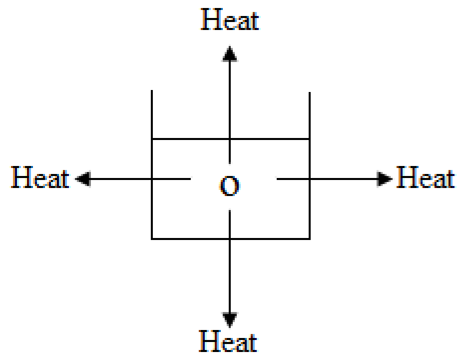
(b) Tissue respiration in all living organisms

$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$ + energy

(c) Dissolving sodium hydroxide crystals in water

In exothermic reactions;

- Energy is released to the surroundings



- Reactants have more energy than products
- ΔH is negative
- Temperature of the system increases

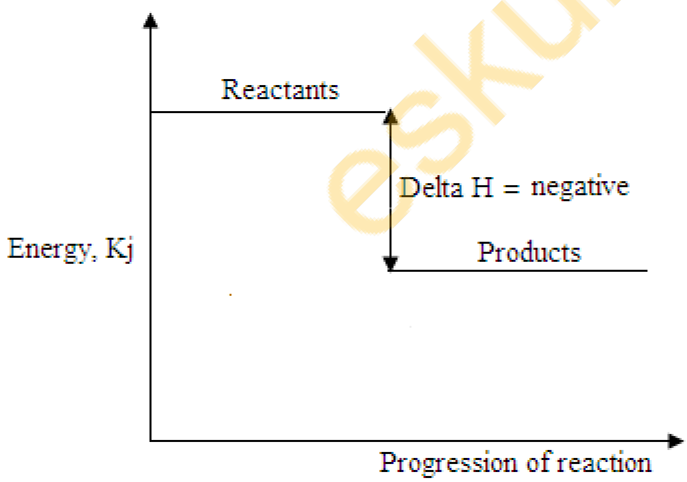
Exothermic reaction as a bond forming process

In general, when bonds are formed, heat is liberated to the surroundings. For example, in a combustion reaction, carbon burns in oxygen to form carbon dioxide as the atoms combine.

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}\ \Delta H = -394\ Kj/mol$

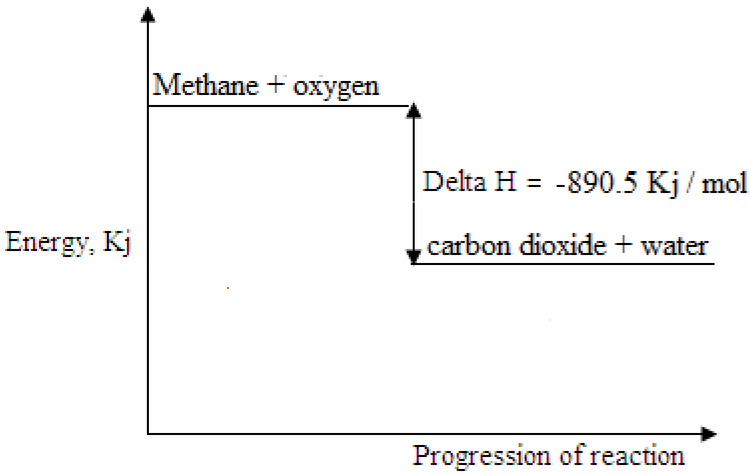
The negative sign in front of the energy value indicate that the reaction is exothermic. Therefore, exothermic reactions involve bond forming and the bonds formed are relatively stronger than bonds broken

Energy level diagram for exothermic reactions



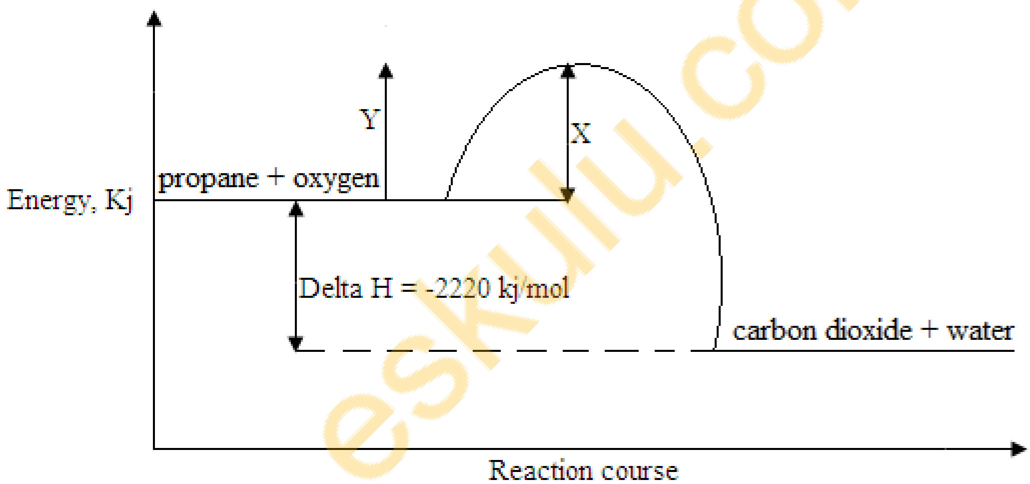
Example of energy level diagram for the combustion of methane

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}\ \ \Delta H = -890.4\ Kj/mol$$



Example

1. The diagram below is an energy level diagram for the combustion of propane in plentiful supply of air



Use the information above to answer the following questions

(a) Given that propane reacts with oxygen to form carbon dioxide and water, write down a balanced chemical equation including state symbols for this reaction.

(b) State the molar enthalpy change of combustion of propane

(c) What is X?

(d) What is the purpose of Y?

(e) How much heat is energy is released to the surroundings when 2.2g of propane is burnt completely in oxygen?

Solution

(a) $C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(l)}$

(b) $\Delta H = -2220$ kj/mol

(c) Activation energy

(d) To start a chemical reaction

(e) $C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(l)} \quad \Delta H = -2220$ kj/mol

$n(C_3H_8) = \frac{m}{MM}$

$n = \frac{2.2g}{44g/mol}$

$n = 0.05mol$

1 mol $C_3H_8 \rightarrow$ -2220 kj /mol

0.05mol $C_3H_8 \rightarrow$ x

$x = \frac{0.05 \text{ mol x } (-2220 \text{ kj/mol})}{1 \text{ mol}}$

$x = -1110$ Kj/mol

2. When carbon is burnt in air, heat is released to the surroundings
   (a) State with reason, whether the reaction is an endothermic or exothermic process.
   (b) When 1.2g of carbon is completely burned in air, 39Kj of heat is liberated
   Calculate the enthalpy of combustion of carbon
   (c) Draw the energy level diagram for the burning of carbon

Solution

(a) Exothermic process

Reason: Heat is released to the surroundings

(b) From the equation: $C + O_2 \rightarrow CO_2$

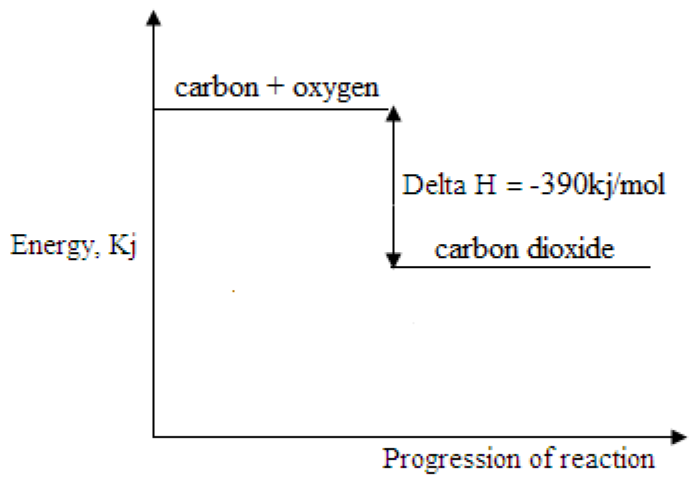
12g C $\rightarrow$ x

1.2g C $\rightarrow$ -39kj

$x = \frac{12g \text{ x } (-39Kj/mol)}{1.2g}$

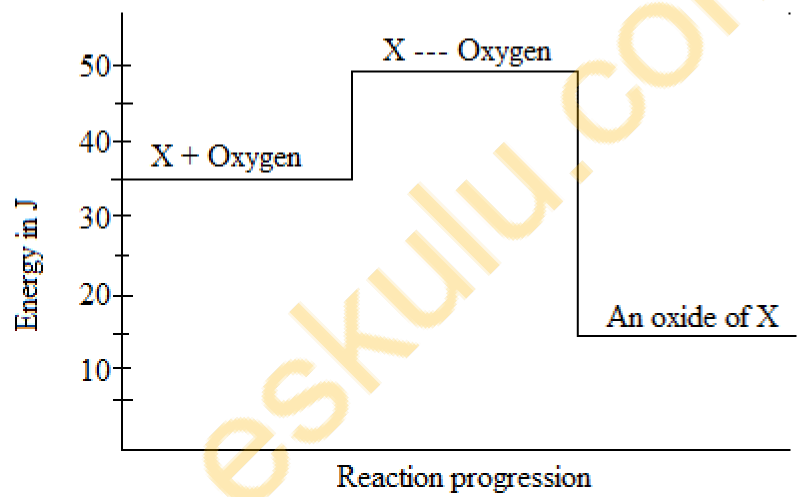
$x = -390$kj/mol

(c)



Exercise

6. Below is the energy level diagram for the combustion of one mole of an unknown fuel X



(a) State the type of reaction represented by the diagram

(b) State the energy levels for:

(I) reactants

(II) products

(c) Calculate the energy change for this reaction

(d) State one example of an endothermic reaction in every day experiences

(e) What scientific term is used to describe the amount of energy required for the reaction to occur?

Bond energy

Alternative term: Bond enthalpy
Definition: This is the amount of energy required to either break the bond or energy released when one mole of bond is formed

Calculating enthalpy change using bond energies
Formula: $\Delta H = \sum$ bond energies of reactants – $\sum$ bond energies of products
Example

1. Calculate the enthalpy change for the reaction between hydrogen and chlorine gas given the following bond energies

Bond	**Bond energy, Kj/mol**
Cl – Cl	243
H – H	436
H – Cl	431

Solution
$H_2 + Cl_2 \rightarrow 2HCl$
$H - H + Cl - Cl \rightarrow 2(H - Cl)$
$\Delta H = \sum$ bond energies of reactants – $\sum$ bond energies of products
= (436 + 243) – (2 x 431)
= 679 – 862
= -183Kj/mol

Exercise

1. Some bond enthalpy terms in kj/mol are shown in the table below

H – H	C – Cl	H – Cl	N ≡ N	N – H
+436	+242	+431	+945	+389

Calculate the enthalpy change for the reaction:
$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$

Electrochemistry
Electrolysis
Electrolysis is the decomposition of an electrolyte by using an electric current
Electrolyte
An electrolyte is a substance which conduct electricity in fused (molten) or in solution and is thereby decomposed

Examples of electrolytes

- Aqueous sulphuric acid
- Aqueous hydrochloric acid
- Aqueous nitric acid
- Aqueous sodium chloride
- Aqueous sodium hydroxide
- Aqueous carbonic acid
- Aqueous ethanoic acid

Strong electrolyte
It is a substance which ionizes completely and produces a lot of ions in solution which are able to carry out an electric current rapidly.
Examples of strong electrolytes

- Aqueous sodium hydroxide
- Aqueous sodium chloride
- Aqueous copper (II) sulphate
- Aqueous hydrochloric acid
- Aqueous sulphuric acid
- Aqueous nitric acid

Weak electrolyte
It is a substance which ionizes partially and it conducts electric current only slightly and therefore undergoes slight decomposition
Examples of weak electrolytes

- Carbonic acid
- Organic acids e.g. ethanoic acid

Non electrolyte
It is a substance which does not conduct electricity I fused or in solution state
Examples of non-electrolytes

- Sugar solution
- Ethanol
- Petrol
- Benzene
- Tetra chloromethane

They do not conduct electricity because they exist only in form of molecules and not capable of ionization
Conductor
It is a solid substance that allows electricity to pass through without decomposing e.g. all metals

Non conductor

Alternative term: Insulator
It is a solid that does not conduct electricity e.g. plastics, wood and glass

Electrodes

They are conductors that allows electricity in and out of an electrolyte

Anode

It is a positively charged electrode
It is an electrode connected to the positive terminal of the power supply

Cathode

It is a negatively charged electrode
It is an electrode connected to the negative terminal of the power supply

Cations

They are positively charged ions

Example

Cation	Formula of cation
Aluminium ion	Al^{3+}
Calcium ion	Ca^{2+}
Hydrogen ion	H^{+}
Magnesium ion	Mg^{2+}
Sodium ion	Na^{+}

Anions

They are negatively charged ions

Example

Anion	Formula of anion
Bromide ion	Br^{-}
Chloride ion	Cl^{-}
Hydroxide ion	OH^{-}
Iodate ion	I^{-}
Sulphate ion	SO_4^{2-}

The ionic theory

The ionic theory states that: The electrolytes contain ions and when no current is passing, the ions wander about randomly in the solution. If the electric circuit is closed, the cathode immediately becomes negatively charged and the anode becomes positively charged. The anode attracts negatively charged ions while the cathode attracts the positively charged ions.

The selective order of discharge of ions

The discharge of ions varies from one electrolyte set up to another. When two or more ions of similar charge are present under similar conditions in a solution e.g. H^+ and Na^+or $OH^-$and SO_4^{2-}, one is preferentially selected for discharge. The ion selected for discharge of competing ions depends on the following factors:

- The position of ions in the electrochemical series
- The concentration of ions in solution
- The nature of electrodes

1. **Position of ions or radicals in the electrochemical series**

Ions are arranged in the order of decreasing order of stability and amount of energy they require for them to get discharged from an aqueous solution when an electric current is made to flow through the solution. The arrangement is called electrochemical series. Note that the electrochemical series is slightly different from the reactivity series of metals.

Most stable ions Increase in stability and require energy least stable ions

$$\longleftarrow$$

$$K^+, Ca^{2+}, Mg^{2+}, Al^{3+}, Zn^{2+}, Fe^{2+}, Pb^{2+}, H^+, Cu^{2+}, Ag^+, SO_4^{2-}, NO_3^-, Cl^-, Br^-, I^-, OH^-$$

The ions of very reactive metals such as potassium, sodium, calcium, magnesium and aluminium are more stable in solution and hence require more energy to be liberated as compared to the ions of less reactive metals such as copper, silver, mercury and gold.

Therefore, when the ion of a more reactive metal is competing for discharge against the ion of a less reactive metal, the less stable ion requiring lower energy for discharge is liberated in preference to the more stable ion. Positive ions will be discharged at the cathode and negative ions at the anode. For example, H^+ will be discharged at the cathode in preference to Na^+ since H^+ are less stable and require lower energy for discharge. Equally, the OH^- will be discharged at the anode in preference to SO_4^{2-}since the $OH^-$are less stable and require lower energy for discharge.

2. **Concentration of ions in aqueous solution**

Concentration has no effect on the selective discharge of metal cations. However, the concentration of anions is the principle factor that determines which anions will be liberated regardless of their position in the electrochemical series. The anion with the highest concentration is selectively discharged from the solution in preference to the one whose concentration is lower regardless of their position in the electrochemical series. For example, in the electrolysis of concentrated sodium chloride, both OH^- and Cl^- are present. The Cl^- are discharged in preference to OH^-owing to the high concentration of Cl^- in solution despite Cl^- requiring more energy than OH^-. This is the only case in which the order of discharge as stated by the electrochemical series is reversed.

3. **Nature of electrodes**

Inert or unreactive electrodes such as platinum and graphite (Carbon) have no effect on the selective discharge of competing ions on their surfaces. However, active electrodes such as copper, mercury and most metals have an effect on the preferential discharge of competing ions. An active electrode selectively discharges more stable ions in preference to less stable ions. For example, in the electrolysis of sodium chloride solution using inert electrodes, H^+ are discharged at the cathode in preference to Na^+. However, when the mercury cathode has been used, Na^+ are selectively discharged to form sodium Amalgam (NaHg) in preference to the H^+.

Electrolysis of lead (II) bromide

Ions present in lead (II) bromide

Pb^{2+} and Br^-

At the cathode

Lead (II) ions, Pb^{2+} , move to the cathode where they gain electrons to become lead atoms

$Pb^{2+} + 2e \rightarrow Pb$

Lead is given off as a liquid

At the anode

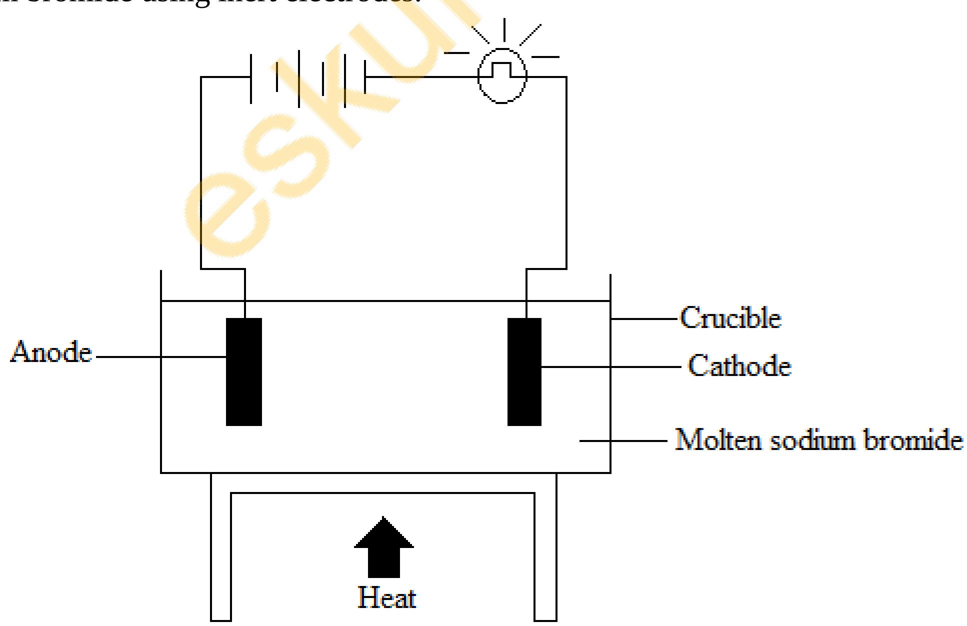
The bromide ions, Br^-, each lose one electron to form bromine atoms. They pair as molecules

$2Br^- \rightarrow Br_2 + 2e$

Bromine is given off as a gas.

Exercise

1. The diagram below shows apparatus which were used by Jane Naosa to electrolyte molten sodium bromide using inert electrodes.



(a) State the ions present in molten sodium bromide

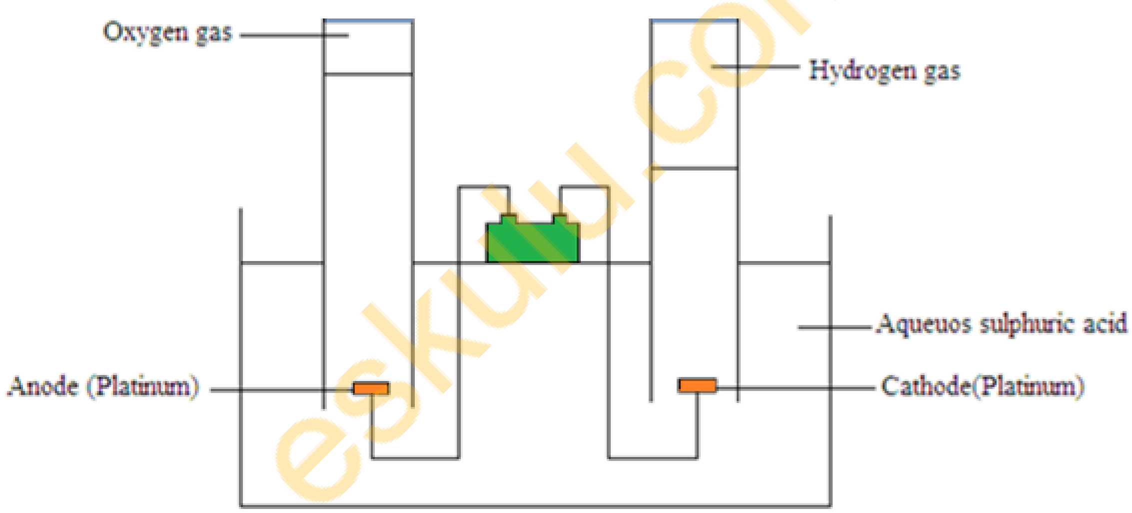
(b) State two suitable materials for the electrodes

(c) Name the product formed at each electrode. Write the equation for the reaction occurring at each electrode
(d) A silvery coloured liquid was seen floating on the electrolyte around the cathode. Name this liquid and explain why it floats on molten sodium bromide.
(e) State what was seen at the anode
(f) Why was molten sodium bromide being heated?
(g) Jane Naosa repeated the electrolysis but using aqueous sodium bromide.
Write the;
(I) Equation for the anode and cathode reactions
(II) Overall equation for this electrolysis

Electrolysis of acidified water

Pure water is a weak electrolyte. It only partially ionizes. Therefore, it is a poor conductor of electricity. The addition of hydrochloric acid or sulphuric acid makes it a strong electrolyte.

Hoffman apparatus



Ions present in aqueous sulphuric acid

$H^+$and SO_4^{2-} : from sulphuric acid

H^+andOH^- : from water

At the cathode

Hydrogen ions, H^+, are attracted to the cathode where they are discharged as hydrogen gas

$2H^+_{(aq)} + 2e \rightarrow H_{2(g)}$

2 volumes of hydrogen gas are produced

At the anode

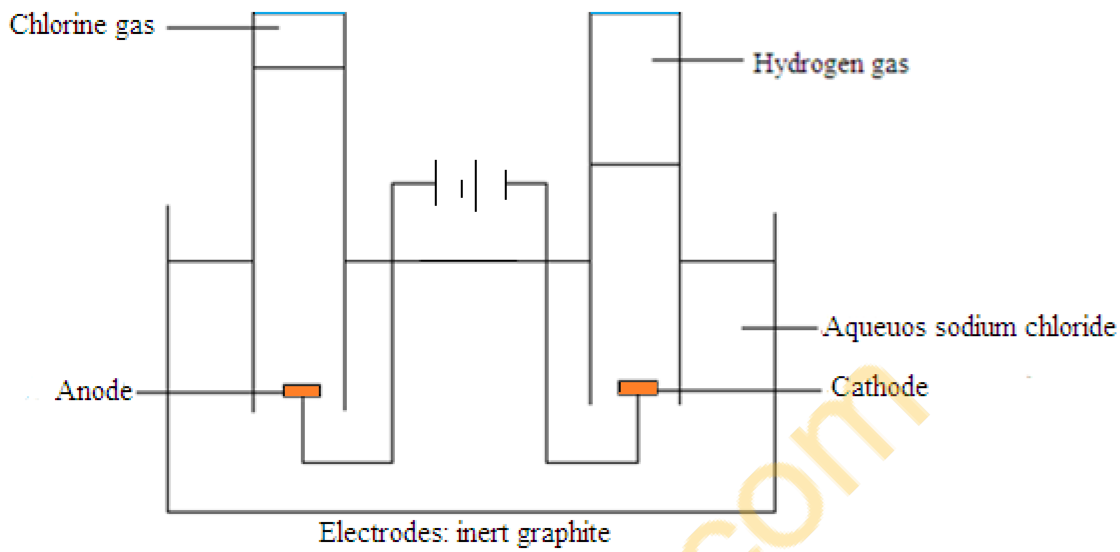
Hydroxide ions, OH^-, and sulphate ions, SO_4^{2-}, are attracted to the anode. But the OH^- are preferred for discharge due to their position in the electrochemical series.

$4OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)} + 4e$

1 volume of oxygen is produced

Electrolysis of sodium chloride

Concentrated sodium chloride is called brine.



Ions present in aqueous sodium chloride

Na^+ and Cl^- : from sodium chloride

H^+ and OH^- : from water

At the anode

Chloride ions, Cl^-, are discharged in preference to hydroxide ions, OH^- due to the concentration factor.

$2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e$

At the cathode

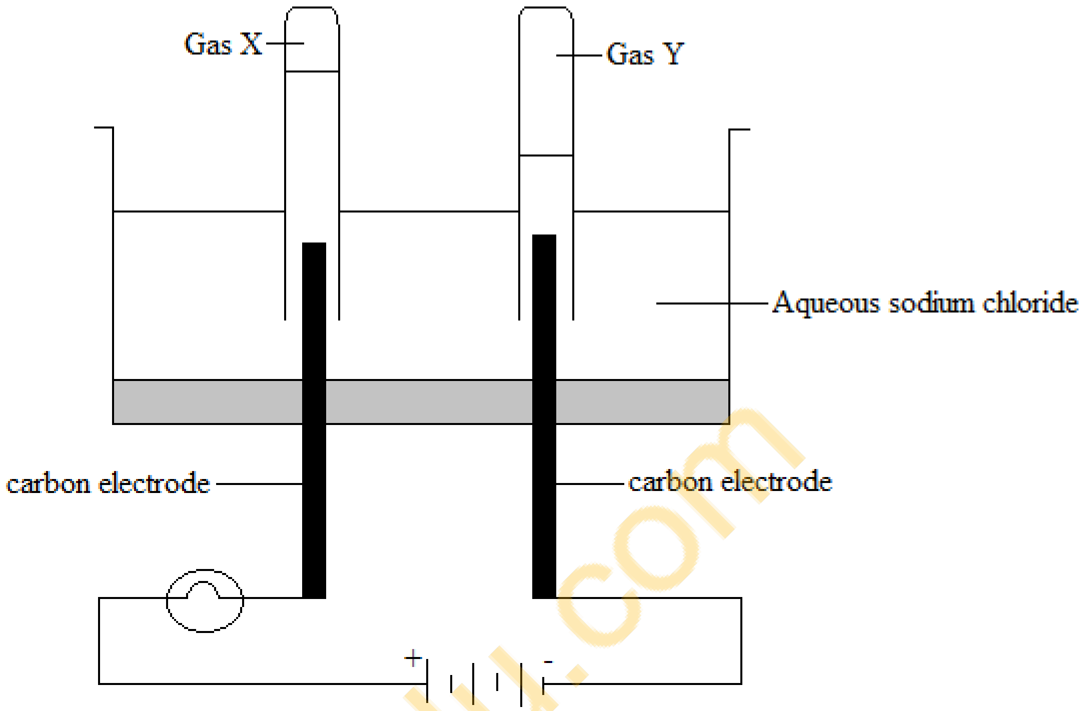
Hydrogen ions, H^+, are discharged in preference to sodium ions, Na^+.

$2H^+_{(aq)} + 2e \rightarrow H_{2(g)}$

The sodium ions, Na^+ , and hydroxide ions, OH^-, that remain solution combine to form sodium hydroxide and electrolyte gradually becomes alkaline due to the presence of sodium hydroxide, NaOH.

Exercise

1. Aqueous sodium chloride was electrolyzed in the apparatus shown below



(a) Name the gas X

(b) How can gas X be identified?

 (I) Test

 (II) Result

(c) Name the gas Y

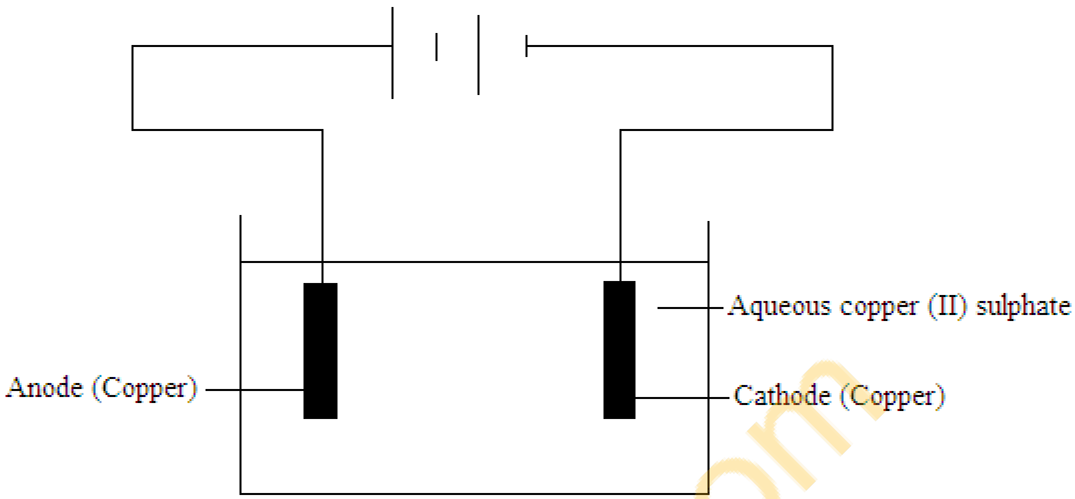
(d) Write the equation for the half – reaction in which Y is produced

(e) Why is the volume of X smaller than that of Y?

(f) The electrolysis of aqueous sodium chloride is used to manufacture important chemicals in industry. Name the major product of this electrolysis.

Electrolysis of aqueous copper (II) sulphate

(a) Using copper electrodes



Ions present in aqueous copper (II) sulphate

Cu^{2+} and SO_4^{2-} : from copper (II) sulphate

H^+ and OH^- : from water

At the cathode

Copper (II) ions gain two electrons and deposited as copper atoms on the cathode.

$Cu^{2+} + 2e \rightarrow Cu$

The copper cathode grows as copper atoms are deposited

At the anode

Copper atoms lose two electrons and enter the solution as copper (II) ions

$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e$

(b) Using an inert electrode e.g. carbon (graphite)

At the cathode

Copper (II) ions, Cu^{2+} are liberated in preference to hydrogen ions, H^+, to form copper metal because Cu^{2+} require less energy for discharge than H^+. Hydrogen ions remain in solution

$Cu^{2+}_{(aq)} + 2e \rightarrow Cu_{(s)}$

At the anode

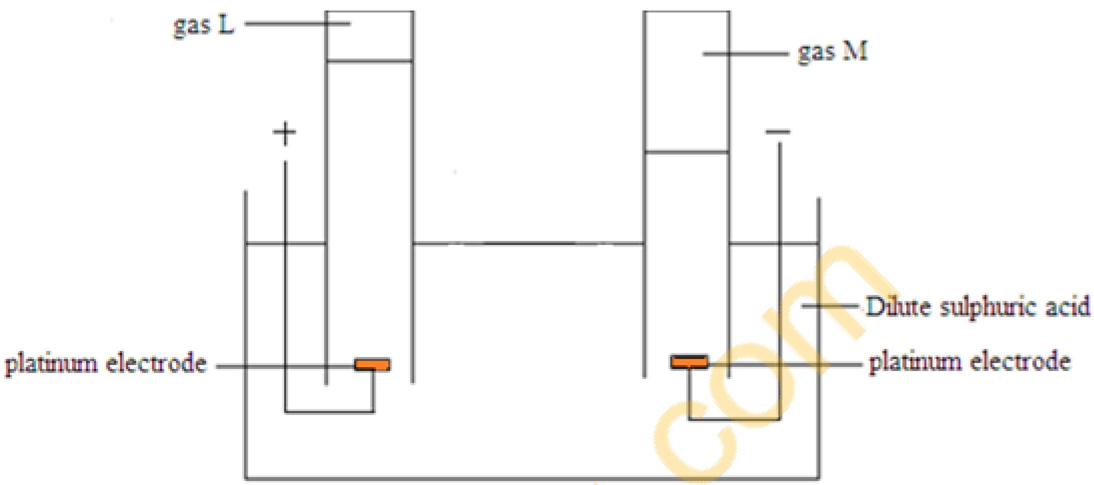
Hydroxide ions, OH^-, from water are discharged in preference to sulphate ions, SO_4^{2-}, to produce oxygen gas becausethey require less energy for discharge and are more concentrated than sulphate ions. The sulphate ions remain in solution

$4OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)} + 4e$

The hydrogen ions and sulphate ions that remain solution combine to form sulphuric acid. As aqueous copper (II) sulphate solution is electrolyzed using inert electrode, its blue colour disappears due to the removal of copper (II) ions and the solution becomes acidic.

Example

1. The diagram below shows the electrolysis of dilute sulphuric acid



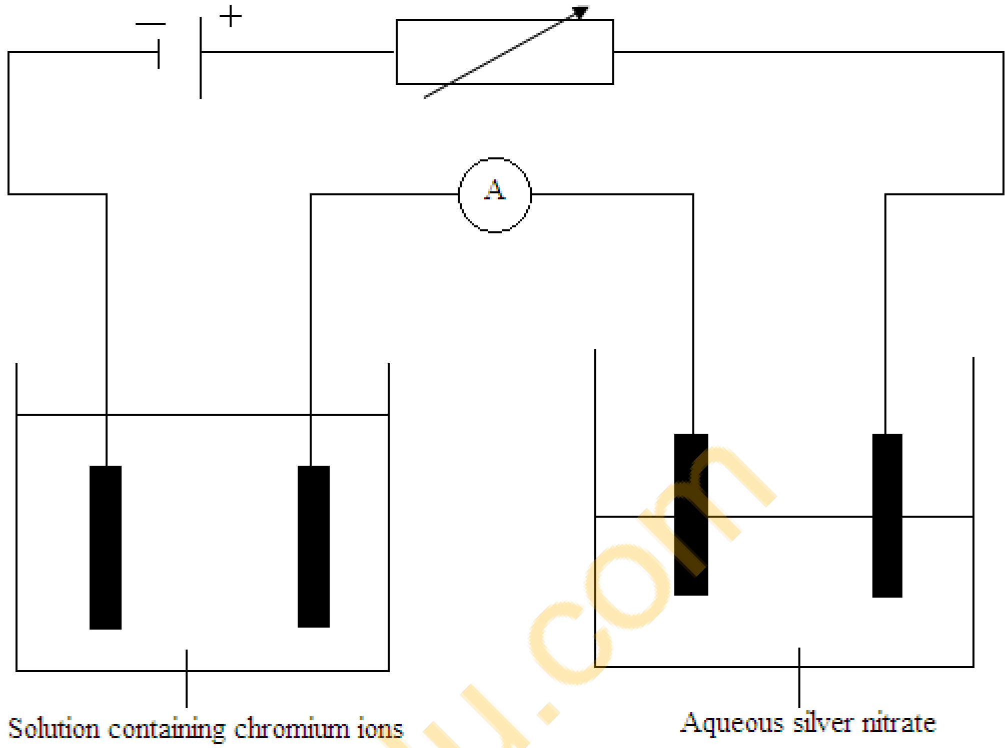
(a) Give the names of the gases labeled
  (I) L
  (II) M
(b) During the electrolysis of dilute sulphuric acid, which ions move;
  (I) Positive electrodes
  (II) Negative electrodes
(c) Give the equations for the reactions occurring at:
  (I) The positive electrode
  (II) Negative electrode

Solution

(a) (I) L = Oxygen
  (II) M = Hydrogen
(b) (I) Hydroxide ions, OH^-
  (II) Hydrogen ions, H^+
(c) (I) $4OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)} + 4e$
  (II) $2H^+_{(aq)} + 2e \rightarrow H_{2(g)}$

Exercise

1. In order to determine the charge on chromium ion, two electrolytic cells were connected in series, one containing an aqueous solution of chromium ions and the other an aqueous solution of silver nitrate as shown in the diagram below.

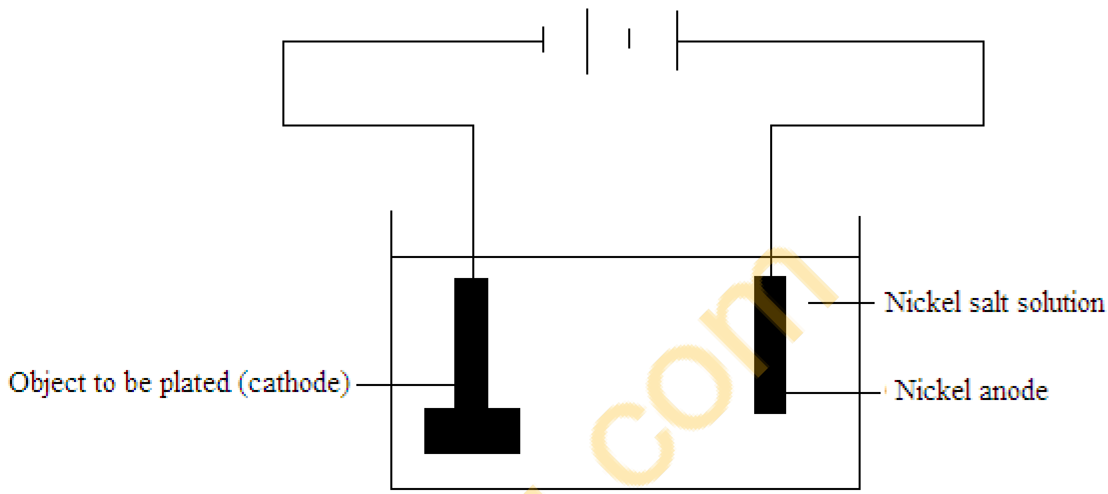


(a) What name is given to the electrode where chromium and silver are deposited?
(b) Write down the half reaction showing the discharge of silver ions
(c) The two aqueous solutions in the diagram above conduct electricity. What name is given to compounds which conduct electricity and are decomposed?
(d) At the end of the experiment, it was found that 2.02g of silver and 0.324g of chromium had been deposited. Calculate the charge on a chromium ion. Show your working.
(e) Suggest the formula of the compound, chromium, in solution.
(f) Chromium is a metal. Explain how the structure of metals enables them to conduct electricity and beaten into thin sheets.

Industrial application of electrolysis

1. Electrolysis is used in extraction of very reactive metals such as potassium, sodium, magnesium, calcium and aluminium
2. Electrolysis is used in the refinery and purification of metals such as copper and zinc.
3. Electrolysis is used in the electroplating metals. Electroplating is the art of covering the surface of a metal with a thin adherent metal coating by means of electrolysis. Electroplating is done to protect the surface of the base metal against corrosion or for a purely decorated effect. Metals that may be used for electroplating include nickel, silver,

gold, chromium, zinc, tin etc. The object to be electroplated is used as the cathode and is immersed in a solution containing ions of the metal to be plated. When an electric current passes through the solution, metal ions that migrated to the cathode are reduced, depositing on the object as free metal. The metal deposited on the object is replaced in the solution by using an anode of the same metal.



Anode reaction

Nickel dissolves at the anode and replenishes the solution

$Ni_{(s)} \rightarrow Ni^{2+}_{(aq)} + 2e$

Cathode reaction

Nickel plated on the object

$Ni^{2+}_{(aq)} + 2e \rightarrow Ni_{(s)}$

4. Sodium hydroxide, chlorine gas and hydrogen gas are manufactured on a large scale by electrolysis of brine in the mercury cathode cell.

Exercise

1. Electrolysis is also used to purify copper.
   (a) Draw a labeled diagram which could be used to purify a block of impure copper
      (I) Describe what happens at each electrode during electrolysis
      (II) Write equations to show the reactions at each electrode

Faraday's laws of electrolysis

Faraday's first law of electrolysis

The law states that: The mass of a substance produced at the electrode during electrolysis is directly proportional to the quantity of electricity passing through the electrolyte.

$m \propto Q$

$m \propto It$

$Q = It$

m = mass[g]

I = current [A]

t = time [s]

Q = charge (quantity of electricity) [Coulombs, C]

Example

1. A current of 0.4A flows for about 1500 seconds. Calculate the quantity of electricity.

Solution

$Q = It$

$Q = 0.4A \times 1500s$

$Q = 675C$

Faraday's second law of electrolysis

The law states that: When the same quantity of electricity is passed through different electrolytes the number of moles of the element deposited is inversely proportional to the charges on the ions of the element.

Note

1 Faraday = 1 mole of electrons = 96500C

$$\frac{\text{Charge on ion of solution B}}{\text{Charge on ion of soltion A}} = \frac{\text{Mass of substance A} \div \text{relative atomic mass of A}}{\text{Mass of substance B} \div \text{relative atomic mass of B}}$$

$$n = \frac{Q \times MM}{m \times F} = q$$

$$m = \frac{A_r \times It}{F \times q}$$

n = number of moles [mol]

m = mass [g]

MM = molar mass

F = faraday's constant [96500C /mol]

A_r = relative atomic mass

I = current [A]

q = number of faradays = magnitude of charge on the ion

Example

1. Calculate the charge on an ion X if 5.4g is deposited by a current of 5A passing for 3 hours 13 minutes. [A_r of X = 27]

Solution

$$q = \frac{Q \times MM}{m \times F}$$

$$q = \frac{It \times MM}{m \times F}$$

$$q = \frac{0.5A \times [(60 \times 60 \times 3)+(60 \times 13)]s \times 27g/mol}{5.4g \times 96500C/mol}$$

q = 3

Charge on ion X = X^{3+} or X^{3-}

2. How many moles of electron are required to produce 2.7g of aluminium by electrolytic reduction?

Solution

$Al^{3+} + 3e \rightarrow Al$

3mol of electron $\rightarrow$ 27g

x $\rightarrow$ 2.7g

x = $\frac{\text{3 mol of electron x 2.7g}}{\text{27g}}$

x = 0.3 mol of electron

Exercise

1. An element X has a relative atomic mass of 88. When a current of 0.5A was passed through the fused chloride of X for 32 minutes 10 seconds, 0.44g was deposited at the cathode. Calculate the number of faradays to liberate 1 mole of X.
2. In an electrolysis experiment, the same quantity of electricity deposited 16g of copper and 6g of titanium. What was the charge on the titanium ion?
   [Assume the charge on the copper ion is 2+]
   A 1+ B 2+ C 3+ D 4+

Calculations involving molar volumes

The volume of one mole of any gas is:

- $24dm^3$ ($24000cm^3$) at r.t.p
- $22.4dm^3$ ($22400\ cm^3$) at s.t.p

Example

1. 0.2 faradays of electricity were passed through a solution of dilute sulphuric acid. Calculate the volume of the gases produced or evolved measured at r.t.p at the anode and cathode.

Solution

At the anode oxygen gas is produced.

$4OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)} + 4e$

1F $\rightarrow$ 1 mol of ẽ

x $\rightarrow$ 4 mol of ẽ

x = $\frac{\text{1 F x 4mol}}{\text{1 mol}}$

x = 4F

4F $\rightarrow$ $24dm^3$

0.2F $\rightarrow$ x

x = $\frac{0.2F \times 24dm^3}{4F}$

x = $1.2dm^3$

At the cathode hydrogen gas is produced

$2H^+_{(aq)} + 2e \rightarrow H_{2(g)}$

1mol of ẽ $\rightarrow$ 1F

2mol of ẽ $\rightarrow$ x

$x = \frac{1F \times 2mol}{1mol}$

x = 2F

2F $\rightarrow$ 24dm^3

0.2F $\rightarrow$ x

$x = \frac{0.2F \times 24dm^3}{2F}$

x = 2.4dm^3

2. What volume of oxygen and hydrogen is produced at r.t.p. when a current of 2A is passed for 6 minutes 26 seconds through a solution containing:
   (a) Hydroxide ions
   (b) Hydrogen ions

Solution

(a) $4OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)} + 4e$

Q= It

Q = 2A x [(6 x 60) + 26]s

Q = 772C

1mol of e $\rightarrow$ 96500C

4mol of e $\rightarrow$ x

$x = \frac{4mol \times 96500C}{1mol}$

x = 386000C

386000C $\rightarrow$ 24dm^3

772C $\rightarrow$ x

$x = \frac{772C \times 24dm^3}{386000C}$

x = 0.048dm^3

(b) $2H^+_{(aq)} + 2e \rightarrow H_{2(g)}$

1mol of e $\rightarrow$ 96500C

2mol of e $\rightarrow$ x

$x = \frac{2mol \times 96500C}{1mol}$

x = 193000C

193000C $\rightarrow$ 24dm^3

772C $\rightarrow$ x

$x = \frac{772C \times 24dm^3}{193000C}$

x = 0.096dm^3

Exercise

1. 0.2F of electricity were passed through a solution of copper (II) sulphate. Calculate the volume of oxygen produced at r.t.p.

Alloys

Definition: An alloy is a mixture of two or more metals or a metal with a non-metal

The combination of alloys is physical

Alloys are harder than the metals from which they are made

Alloying a metal is one way of increasing its strength

Preparation of alloys

The mixture is usually heated under controlled temperature

The molten mixture is then allowed to cool and solidify

Advantages of alloys

1. They are flexible in use
2. They usually have improved appearance
3. They are durable and reliable
4. They have increased resistance to corrosion

Examples of alloys

Alloy	**Typical composition**		**Particular property**	**Typical use**
Brass	Copper, Cu Zinc, Zn	70% 30%	Harder than pure copper, Resistant to corrosion	Used to make musical instruments
Bronze	Copper, Cu Tin, Sn	90% 10%	Harder than pure copper	Used to make trophies Used to make medallions
Mild steel	Iron, Fe Carbon, C	99.7% 0.3%	Stronger and harder than pure iron	Used to make bridges Used to make boring tools and crushing equipment Used to make automobile bodies and parts e.g. car bodies
Stainless steel	Iron, Fe Chromium, Cr Nickel, Ni	70% 20% 10%	Harder than pure iron Does not rust	Used in cutlery e.g. making razor blades, knives because it does not rust
Solder	Tin, Sn Lead, Pb	50% 50%	Has lower melting point than either tin or lead	Used for soldering and welding

Exercise

1. Alloying is one way of preventing rusting of iron
   (a) Name any two alloys containing iron and state one use for each
   (b) State two other methods of rust prevention

Metals

General physical properties of metals

1. Metals are good conductors of heat and electricity
2. Metals are malleable i.e. they can be hammered into thin sheets
3. Metals are ductile i.e. they can be drawn into long wires
4. Metals are sonorous i.e. they produce sound when hammered
5. Metals are lustrous i.e. they have a silver luster surface when freshly cut
6. Metals are solids at room temperature and pressure except for mercury which is a liquid at room temperature and pressure
7. Metals have high melting and boiling points

The reactivity series

Alternative term: Activity series

Definition: Reactivity is a list of metals with the most reactive metal at the top and the least reactive metal at the bottom.

The order of reactivity can be determined by the reaction of the metal with water or steam and acids. In both types of reaction, if a reaction takes place, hydrogen gas is formed.

Metal	**Symbol**	**Hint**	
Potassium	K	Kaunda	Most reactive
Sodium	Na	Now	
Calcium	Ca	Can	
Magnesium	Mg	Make	
Aluminium	Al	All	
Zinc	Zn	Zambians	
Iron	Fe	Free	
Lead	Pb	Provided	
Hydrogen	H	He	
Copper	Cu	Could	
Mercury	Hg	Hung	
Silver	Ag	All	
Gold	Au	Americans	Least reactive

The reactivity series is related to the tendency of metals to form positive ions

Very reactive metals lose their valence electrons easily to form cations

Metals at the top of series lose the electrons more easily and form ions rapidly and are called electropositive.

Metals at the bottom of the series lose electrons with difficulty and do not readily form ions and are said to be less electropositive

The metal higher up in the series will displace the one below it from aqueous solution of its salts

Hydrogen is included in the series although it is a non-metal. It serves as a reference point in the series. Metals above hydrogen will react with dilute acids to give hydrogen gas, while metals below hydrogen will not react with dilute acids to give hydrogen gas.

Reaction of metals with water or steam

Metal	**Symbol**	**Observation / Equation**
Potassium	K	Reacts very violently with cold water to produce potassium hydroxide and hydrogen gas $2K_{(s)} + 2H_2O_{(l)} \rightarrow 2KOH_{(aq)} + H_{2(g)}$
Sodium	Na	Reacts violently with cold water to produce sodium hydroxide and hydrogen gas $2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$
Calcium	Ca	Reacts less violently with cold water to produce calcium hydroxide and hydrogen gas $Ca_{(s)} + 2H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$
Magnesium	Mg	Reacts very slowly with cold water but violently with steam to produce hydrogen gas $Mg_{(s)} + 2H_2O_{(l)} \rightarrow Mg(OH)_{2(aq)} + H_{2(g)}$ $Mg_{(s)} + H_2O_{(g)} \rightarrow MgO_{(s)} + H_{2(g)}$
Aluminium	Al	Does not react with cold water but react slowly with steam to form aluminium oxide and hydrogen gas. $2Al_{(s)} + 3H_2O_{(g)} \rightarrow Al_2O_{3(s)} + 3H_{2(g)}$
Zinc	Zn	Does not react with cold water but hot zinc burns in steam to form zinc oxide and hydrogen gas. $Zn_{(s)} + H_2O_{(g)} \rightarrow ZnO_{(s)} + H_{2(g)}$
Iron	Fe	Does not react with cold water but react slowly with steam to form iron oxide and hydrogen gas. $Fe_{(s)} + H_2O_{(g)} \rightarrow FeO_{(s)} + H_{2(g)}$
Lead Copper Silver Gold	Pb Cu Ag Au	They are below hydrogen in the reactivity series do not react with cold water not even steam

Reaction of metals with dilute hydrochloric acid

All metals above hydrogen in the reactivity series react with dilute hydrochloric acid to form a salt and hydrogen gas

All metals below hydrogen in the reactivity series do not react with dilute acids

Metal	Symbol	Observation / Equation
Potassium	K	Explosive reaction $K_{(s)} + HCl_{(aq)} \rightarrow KCl_{(aq)} + H_{2(g)}$
Sodium	Na	Explosive reaction $Na_{(s)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_{2(g)}$
Calcium	Ca	Reacts vigorously to give calcium chloride and hydrogen gas $Ca_{(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_{2(g)}$
Magnesium	Mg	Reacts rapidly to give magnesium chloride and hydrogen gas $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$
Zinc	Zn	Reacts quiet slow (moderately fast) to give zinc chloride and hydrogen gas $Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$
Iron	Fe	Reacts slowly to give iron (II) chloride and hydrogen gas $Fe_{(s)} + 2HCl_{(aq)} \rightarrow FeCl_{2(aq)} + H_{2(g)}$
Lead	Pb	Slow, and only if the acid is concentrated $Pb_{(s)} + 2HCl_{(aq)} \rightarrow PbCl_{2(aq)} + H_{2(g)}$
Copper Gold	Cu Au	No reaction, even with concentrated acid

Reaction of metals with oxygen

Metal	Symbol	Behavior / Equation
Sodium	Na	Catches fire with only a little heating. Burns fiercely with a bright yellow flame $4Na_{(s)} + 2O_{2(g)} \rightarrow 2Na_2O_{(s)}$
Magnesium	Mg	Catches fire easily. Burns with a blinding white flame $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$
Iron	Fe	Does not burn, but the hot metal glows brightly in oxygen, and gives off yellow sparks. $Fe_{(s)} + O_{2(g)} \rightarrow Fe_3O_{4(s)}$
Copper	Cu	Does not burn, but the hot metal becomes coated with a black substance. $2Cu_{(s)} + O_{2(g)} \rightarrow 2CuO_{(s)}$
Gold	Au	No reaction, no matter how much the metal is heated.

Example

1. Iron, calcium and copper are metals. The table below describes the reactions of these metals with cold water and steam.
   (a) Place a tick (√) if the reaction will take place and a cross (x) if the reaction will not take place

Metal	**Reaction of metal with cold water**	**Reaction of metal with steam**
Copper		
Iron		
Calcium		

   (b) place these metals in order of chemical activity, starting with the most reactive

Solution

(a)	**Metal**	**Reaction of metal with cold water**	**Reaction of metal with steam**
	Copper	x	x
	Iron	x	√
	Calcium	√	√

(b) Calcium, iron, copper

Exercise

2. The list below shows metals arranged in ascending order of reactivity.
   Silver
   Zinc
   Aluminium
   Sodium
   (a) Using metals from the list only, name:
      (I) A metal which can be displaced by copper
      (II) A metal which can react with cold water to produce an alkaline solution
      (III) A metal which forms an amphoteric oxide when burnt
      (IV) A metal whose carbonate does not decompose when heated
      (V) A metal which forms a stable oxide layer
   (b) Write a balanced chemical equation for the reaction between sodium and water.
3. Place in order of chemical reactivity towards the metals calcium, iron, magnesium and zinc, with the most reactive first. Four other metals A, B, C and D are tested with dilute hydrochloric acid and with water
   Metal A reacts with steam but not with cold water

Metal B does not reacts with steam or cold water but does react with dilute hydrochloric acid

Metal C reacts quickly with cold water

Metal D does not react with water or with dilute hydrochloric acid

(a) Place metals A, B, C and D in order with the most reactive first

(b) State between which two letters (metals) hydrogen should be placed in the above series

Displacement of metals from their compounds

A metal high in the reactivity series will always displace a less reactive metal from the solution of its compound

A metal lower in the reactivity series cannot displace the one above it from the solution of its compound.

Example 1: $Fe_{(s)} + CuSO_{4(aq)}$

Iron will displace copper from the blue copper (II) sulphate solution. Green iron (II) sulphate is formed.

$Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{(aq)} + Cu_{(s)}$

Example 2: $Cu_{(s)} + AgNO_{3(aq)}$

Since copper is above silver in the reactivity series, copper will displace silver from silver nitrate solution. The solution will also turn from colourless to blue due to the formation of copper (II) ions.

$Cu_{(s)} + 2AgNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2Ag_{(s)}$

Example 3: $Fe_{(s)} + ZnSO_{4(aq)}$

Iron is lower than zinc in the reactivity series. Since it is less reactive than zinc, no displacement reaction will take place

Exercise

1. What would you see when zinc is added to copper (II) sulphate solution? Write a balanced chemical equation for the reaction.
2. Will copper react with zinc sulphate? Explain why

Recycling of metals

Metals are finite resources. It is essential that we recycle metals that are still useful to us.

Advantages or reasons of recycling metals

1. Better conservation of natural resources, so that reserves last longer. The demand for metal ores will decrease once scrap metal is identified as a viable source of raw material.
2. With recycling, less mining will take place. There will be less air and water pollution caused by mining process
3. More effective waste disposal as scrap metal is recovered. Less landfill space will be needed. This will also solve the problem of litter accumulation

Recycling is sometimes not feasible because of the costs involved. Transportation, sorting through waste and cleaning the scrap metal etc. may cost more than extracting the metal from ores. This is true for some cheaper metals.

Extraction of metals

Most metals occur in the earth's crust as ores

An ore is a compound from which a metal can be extracted.

Methods of extraction

The method of extraction of a metal from the ore depends on its position in the reactivity series

There are three main methods used to extract metals from their ores

1. Electrolysis
2. Reduction
3. Thermal decomposition

Metal	**Symbol**	**Method of extraction**
Potassium Sodium Calcium Magnesium Aluminium	K Na Ca Mg Al	Electrolysis of fused oxides or chlorides
Zinc Iron Lead	Zn Fe Pb	Reduction by coke
Copper Mercury Silver Gold	Cu Hg Ag Au	Thermal decomposition

Stability of compounds

Compounds of very reactive metals like potassium, sodium, calcium, magnesium and aluminium cannot be decomposed by heating or reduction method using hydrogen, carbon or carbon monoxide as reducing agents

Compounds of these metals are very stable and can therefore be extracted by electrolysis

Mild reactive metals like iron, zinc and lead can easily be extracted from their ores by reduction with carbon or carbon monoxide because their compounds are less stable

Less reactive metals like copper and silver can be extracted from their ores by heating

Unreactive metals like gold exists in nature as pure and free metals and can be mined in pure form because their compounds are unstable

Gold always occur in native (Free State)

Silver occurs both in native and combined state

Metal extraction often produces sludge because they contain impurities

Effects of heat on carbonates

(a) Carbonates of group I elements like potassium, sodium, rubidium are extremely stable and hence cannot be decomposed by heat

(b) Carbonates of group II elements like calcium, magnesium, barium and transition elements like zinc, iron, lead, copper etc. are only decomposed to oxides and carbon dioxide when heated. No further decomposition is possible after this. The oxides are extremely stable and can only be reduced by electrolysis.

Examples

(I) $CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$

(II) $MgCO_{3(s)} \rightarrow MgO_{(s)} + CO_{2(g)}$

(III) $CuCO_{3(s)} \rightarrow CuO_{(s)} + CO_{2(g)}$

Effects of heat on nitrates

(a) Nitrates of Group I elements like potassium and sodium decomposes into nitrites and oxygen when heated.

Examples

(I) $2KNO_{3(s)} \rightarrow 2KNO_{2(s)} + O_{2(g)}$

(II) $2NaNO_{3(s)} \rightarrow 2NaNO_{2(s)} + O_{2(g)}$

(b) Nitrates of group II elements like calcium, magnesium, barium and transition elements like zinc, iron, lead, copper etc. are decomposed to metal oxides, nitrogen dioxide and oxygen when heated.

Examples

(I) $2Ca(NO_3)_2 \rightarrow 2CaO + 4NO_2 + O_2$

(II) $2Mg(NO_3)_2 \rightarrow 2MgO + 4NO_2 + O_2$

(III) $2Cu(NO_3)_2 \rightarrow 2CuO + 4NO_2 + O_2$

Aluminium

Aluminium is the most abundant metal in the earth's crust

Ores of aluminium

1. **Bauxite**

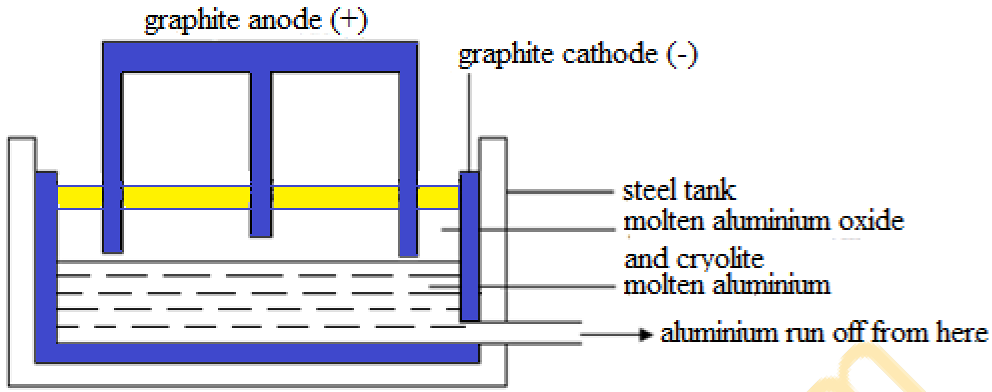
   Chemical name: Aluminium oxide

   Formula: Al_2O_3

The oxide is **very stable** and hence cannot be decomposed by heat or reduction with carbon

Extraction of aluminium

Aluminium is extracted by electrolysis of molten bauxite, Al_2O_3 in the electrolytic cell.
Extraction of aluminium from bauxite is carried out in cell graphite (carbon) electrodes
Bauxite is dissolved in cryolite (Na_3AlF_6) to lower its lower melting point
Aluminium is formed at the cathode and settles at the bottom in molten form
Aluminium is tapped out by opening an outlet when it has accumulated



Aluminium is formed at the cathode
$Al^{3+} + 3e \rightarrow Al$
Oxygen gas is formed at the anode
$2O^{2-} - 4e \rightarrow O_2$
The oxygen gas immediately reacts with the graphite electrodes to form carbon dioxide gas since they are made of carbon. For this reason, the anodes are eaten away and hence they are replaced at regular intervals
$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

Summary

Three main stages in the extraction of aluminium from bauxite

Bauxite is impure aluminium oxide, Al_2O_3

1. Bauxite is purified
2. Pure aluminium oxide is dissolved in molten cryolite, Na_3AlF_6
3. Electrolysis is performed in the cell

The apparent unreactivity of aluminium

Despite being high in the reactivity series, aluminium does not easily react with water and oxidizing acids like nitric acid and sulphuric acid.
Aluminium forms aluminium oxide in the presence of air. This is because shortly after being extracted out, a thin protective layer of aluminium oxide forms on its surface.
This oxide is insoluble and resistant to corrosion. So it forms a protective coating for aluminium.

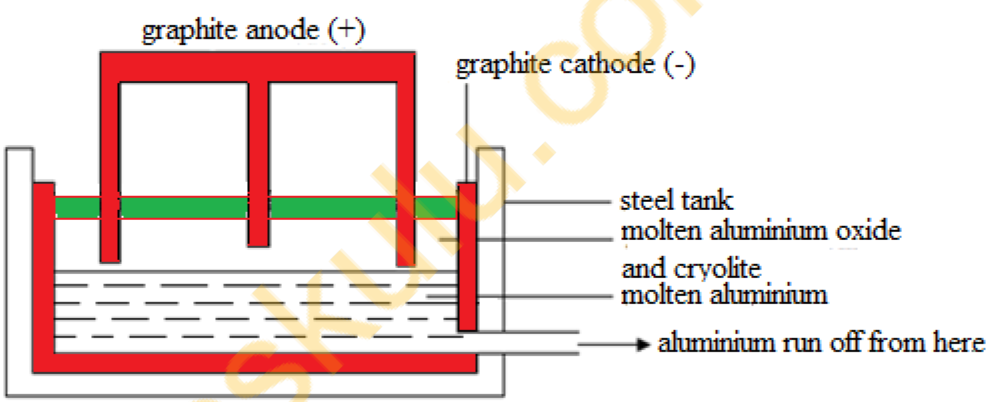
Uses of aluminium

1. It is used in overhead electrical cables because it is a good conductor of electricity

2. It is used in making cooking utensils like sauce pans and kettles because it is a good conductor of heat and it is resistant to heat
3. It is used in making food wrappers and drink cans due to its resistance to corrosion and is non-toxic
4. It used in the manufacture of aeroplanes because it is lighter (low density) and has high strength. It is cheaper and therefore preferred than copper
5. It is used in making light aluminium roofing sheets
6. It is used in making alloys e.g. duralumin
7. It is used in making aluminium paints. The powdered metal is used with oil

Example

1. The diagram below represents the electrolytic cell used for the production of aluminium. The electrolyte contains aluminium oxide and cryolite (sodium aluminium fluoride) and is molten at about 800^oC. The electrodes are made from graphite.



(a) Why is a mixture of cryolite and aluminium oxide, and not pure aluminium oxide, is used as the electrolyte?
(b) Write the equations for the reactions occurring at

(I) the positive electrode
(II) the negative electrode

(c) Explain why the graphite anodes need to be replaced at regular intervals?
(d) Calculate the maximum mass of aluminium that can be made from 408 tonnes of aluminium oxide.
(e) Aluminium foil is used to make food containers because it does not corrode easily.
(I) Explain why aluminium does not corrode easily

(II) Give a use, other than for food containers, together with the physical property that makes aluminium suitable for that use

(III) Give a further use of aluminium, other than for food containers, together with a different physical property from that given in part (II) that makes aluminium suitable for that use

Solution

(a) Aluminium oxide has a very high melting point. With the addition of cryolite, the melting point is greatly reduced. Hence it is economical

(b) (I) $2O^{2-}_{(l)} \rightarrow O_{2(g)} + 4e$

(II) $Al^{3+}_{(l)} + 3e \rightarrow Al_{(s)}$

(c) The oxygen produced at the graphite anode oxidizes the graphite to carbon dioxide.

(d) From the equation: $2Al_2O_3 \rightarrow 4Al + 3O_2$

$204g\ Al_2O_3 \rightarrow 108g\ Al$

$408\ \text{tonnes}\ Al_2O_3 \rightarrow x$

$x = \frac{408\ \text{tonnes x}\ 108g}{204g}$

x = 216 tonnes Al

(e) (I) Aluminium forms aluminium oxide in the presence of air. This oxide is insoluble and resistant to corrosion. So it forms a protective coating for aluminium

(II) It is used in making cooking utensils since it has very good conductivity, in addition to its good appearance and resistance to corrosion use

(III) It is a component in several alloys used in air craft construction. Its favorable use is due to its low density and high tensile strength

Exercise

1. Before experimenting with aluminium to place in the reactivity series, the surface of the aluminium must be scrapped. Why is this necessary?
2. Give two reasons why it is important to recycle metals

Copper

Ores of copper

1. **Malachite**

   Chemical name: Copper (II) carbonate

   Formula: $CuCO_3$
2. **Copper pyrites**

   Chemical name: Copper (I) sulphide

   Formula: Cu_2S

Extraction of copper

Copper is extracted from copper (I) sulphide by thermal decomposition. This is usually done in the presence of oxygen. The copper (I) sulphide is reduced to copper by heating in a controlled supply of air. The impure copper is called blister copper.

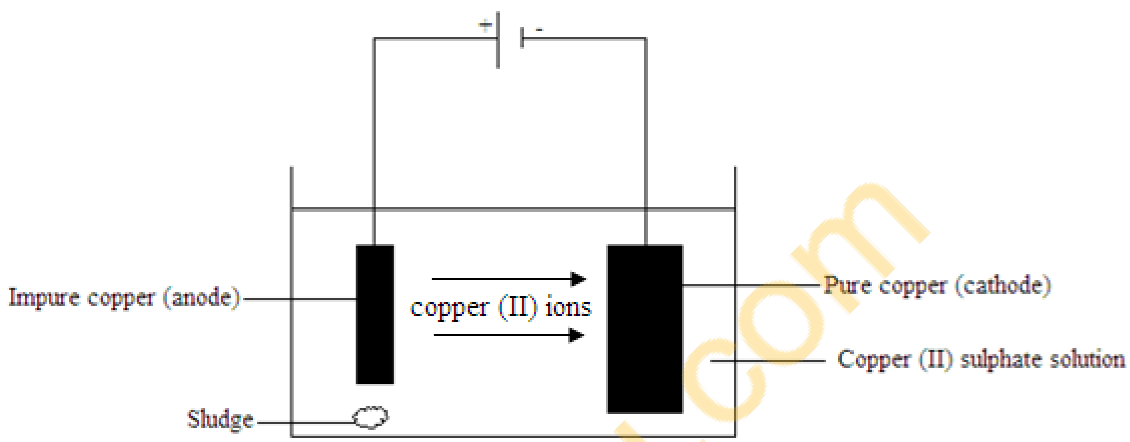
$Cu_2S_{(l)} \quad + \quad O_{2(g)} \rightarrow 2Cu_{(s)} \quad + \quad SO_{2(g)}$

Purification of copper

The copper formed is impure

Silver, gold and cobalt are usually present as impurities

It is purified by electrolysis



The electrolyte is an acidified solution of an electrolyte containing the metal ion

The impure copper anode loses mass as copper ions are formed.

At the anode pure copper from pure anode goes into solution, reducing the size of the anode

$Cu_{(s)} \rightarrow Cu^{2+}{}_{(aq)} + 2e$

Impurities called anode sludge fall to the bottom of the cell. The anode sludge may contain valuable metals such as silver and gold

At the anode copper ions accept two electrons each and are deposited as pure copper metal on the electrode which increases in size.

$Cu^{2+} + 2e \rightarrow Cu$

Uses of copper

1. it is used in making electric cables because it is a good conductor of electricity
2. it is used in making alloys e.g. bronze and brass
3. it is used in making coils
4. it is used in making ornaments in jewelry industries e.g. coins, necklaces and rings
5. it is used in making cooking utensils and boilers since it is a good conductor of heat

Zinc

Ores of zinc

1. **Zinc blend**

Chemical name: Zinc sulphide
Formula: ZnS

2. **Calamine**
Chemical name: Zinc carbonate
Formula: $ZnCO_3$

Extraction of zinc

Zinc blend is heated in air to form zinc oxide and sulphur dioxide

$2ZnS_{(s)} + 3O_{2(g)} \rightarrow 2ZnO_{(s)} + 2SO_{2(g)}$

Zinc oxide is then reduced to zinc metal by carbon monoxide

$ZnO_{(s)} + CO_{(g)} \rightarrow Zn_{(s)} + CO_{2(g)}$

Uses of Zinc

1. It is used in making alloys e.g. brass which is an alloy of zinc and copper
2. It is used to galvanize iron to prevent rusting
3. It is used in making roofing sheets
4. It is used in the preparation of dry cell batteries

Exercise

1. Zinc is extracted from its ore, zinc blend. The zinc blend is heated in air to form the oxide and an acidic gas. The oxide is then reduced to zinc.
   (a) Name the acidic gas
   (b) What is the chemical name of zinc blend
   (c) Write down the chemical equations for the two processes involved in the extraction of zinc from zinc blend
   (d) Why is not possible to reduce aluminium oxide to aluminium using carbon?
   (e) Important uses of zinc are galvanizing steel, making alloys and manufacturing dry cells.
      (i) Why is steel galvanized?
      (ii) Name an alloy containing zinc metal. State one physical property of the alloy

Iron

Ores of iron

1. **Heamatite**
Chemical name: Iron (III) oxide
Formula: Fe_2O_3
2. **Magnetite**
Chemical name: Tri iron tetra oxide
Formula: Fe_3O_4
3. **Siderite**
Chemical name: Iron (II) carbonate
Formula: $FeCO_3$

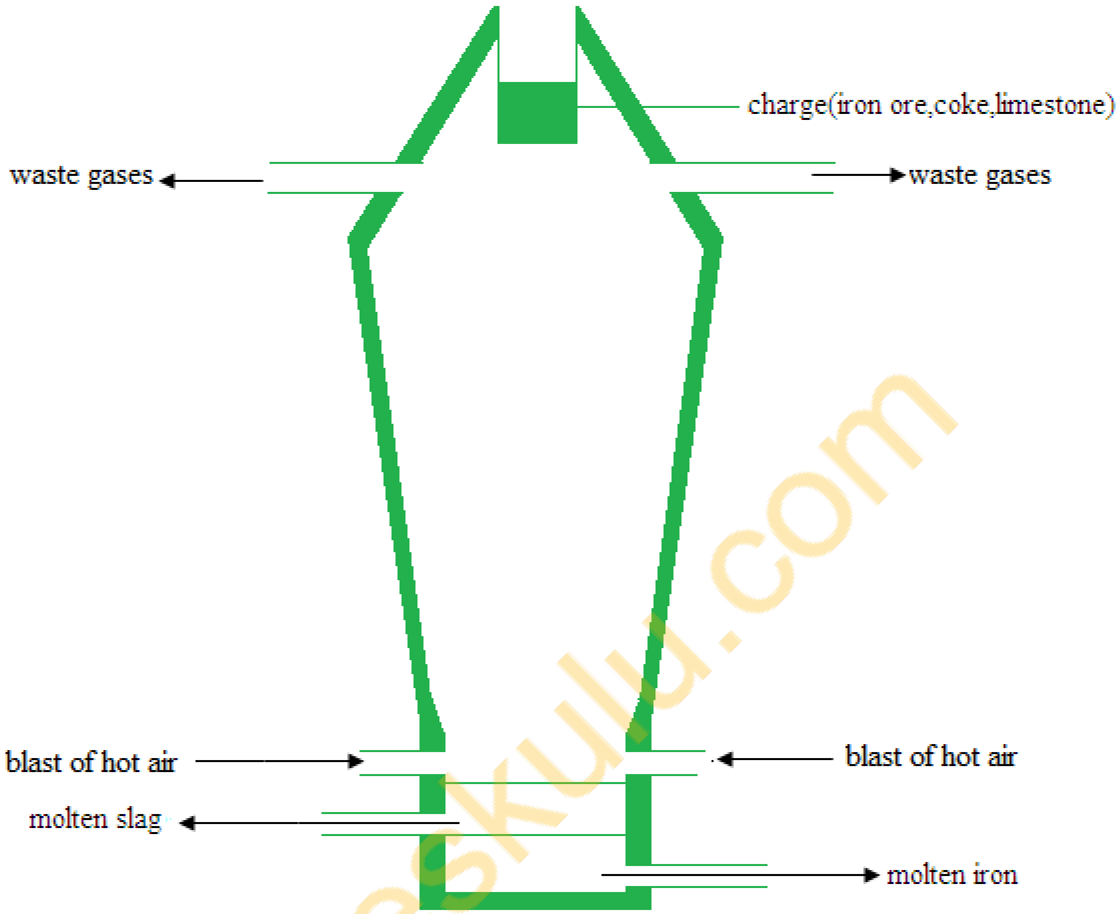
Extraction of iron

Iron is extracted from Heamatite by the reduction method in the blast furnace

The blast furnace

The blast furnace is tower of about 30 – 40 meters high

It is made of steel, and lined with fire proof bricks with a high melting point.



Charge

To extract iron, three substances called charge (raw materials) are mixed together.

These are:

1. **Iron ore**: The chief ore is heamatite, Fe_2O_3
2. **Lime stone**: This is calcium carbonate, $CaCO_3$
3. **Coke**: This is pure carbon, C

The charged is crushed and placed into the top of the blast furnace. It is then roasted in air.

Chemical reactions in the blast furnace

1. Coke reacts with oxygen in hot air to form carbon dioxide. This is an oxidation process. The reaction rises the temperature in the blast furnace.

   $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

   Hot air (oxygen) is forced into the bottom of the blast furnace

2. Carbon dioxide rising up reacts with more coke to form carbon monoxide
   Carbon (coke) is a reducing agent because it reduces carbon dioxide to carbon monoxide
   $C_{(s)} + CO_{2(g)} \rightarrow 2CO_{(g)}$
3. Carbon monoxide react with iron oxide to form liquid iron and carbon dioxide
   Carbon monoxide gas is a reducing agent because it reduces the iron oxide to iron
   $Fe_2O_{3(s)} + 3CO_{(g)} \rightarrow 2Fe_{(l)} + 3CO_{2(g)}$
   $Fe_3O_{4(s)} + 4CO_{(g)} \rightarrow 3Fe_{(l)} + 4CO_{2(g)}$
   Iron trickles to the bottom of the furnace where it is tapped off.
4. Limestone decomposes to calcium oxide and carbon dioxide
   $CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$
   The purpose of limestone is to act as a flux
5. Calcium oxide reacts with silicon dioxide (sand) to form calcium silicate or slag. The slag runs down the furnace and floats on the iron. This prevents the molten iron from being oxidized by the incoming oxygen. Slag is tapped off.
   $CaO_{(s)} + SiO_{2(s)} \rightarrow CaSiO_{3(l)}$
   Calcium oxide is used to remove impurities such as silicon dioxide(SiO_2)

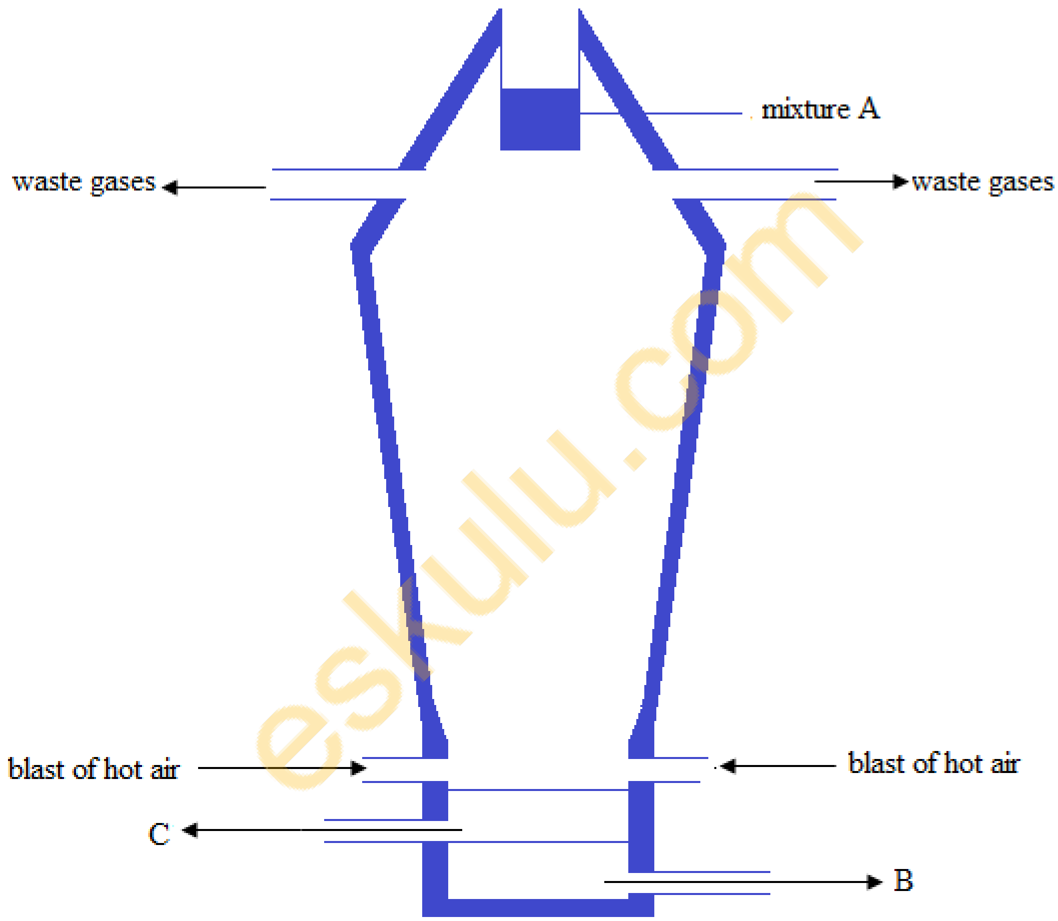
Uses of slag

1. It is used in the manufacture of cement
2. It is used for making roads

Note: The raw iron obtained in this process is called cast iron or pig iron. It contains impurities. The purest form of iron is called wrought iron.

Exercise

1. The diagram below shows the blast furnace



(a) Name the components of mixture A, B and C in the diagram above
(b) Explain how the furnace is heated to the high temperature needed for the reduction of the iron ore.
(c) Name two processes that lead to the production of carbon dioxide in the blast furnace
(d) Name the gaseous reducing agent in the furnace
(e) Which of the components of mixture A produces the gaseous reducing agent in (b).

(f) Construct an equation for a reaction in which the gaseous reducing is produced.
(g) Outline the changes that take place in a blast furnace during the manufacture of iron
(h) The iron produced by the blast furnace is converted into alloys. Explain what is meant by an alloy.
(i) How is mixture C removed from the furnace?
(j) What is the chemical name given to the change by which iron oxide becomes iron?

Air

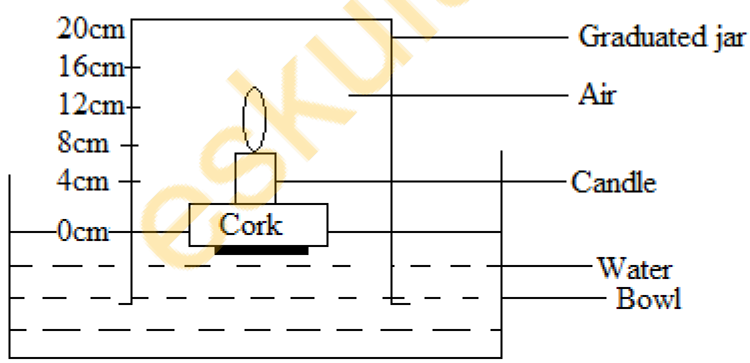
Clean, dry air is a mixture of gases

Composition of clean dry air

Component	**Percentage**
Nitrogen, N_2	79%
Oxygen, O_2	20%
Carbon dioxide, CO_2	0.04%
Noble gases (mainly argon)	0.93%

Example

1. A student wanted to find the percentage of water in air. A candle wax was fixed to a weighed piece of cork. This was floated on water. The student then lit the candle and lowered a graduated gas jar over it, as shown below



The candle flame slowly went out and the water level rose in the jar

(a) Why did the flame go out?
(b) What two substances are produced when a wax (hydrocarbon) candle burns?
(c) What did the water replace when it rose in the jar?
(d) Why did the water not fill the jar?
(e) The water rose to the 3.5cm mark on the jar. Calculate the percentage change of volume of the air.

Solution

(a) All the oxygen needed to support combustion in the jar is used up
(b) Carbon dioxide and water
(c) The water replaced the oxygen gas used up to burn the candle
(d) Oxygen occupies only 20% by volume of air. The remaining gas is mainly nitrogen gas
(e) Percentage change = $\frac{3.5\text{cm}}{20\text{cm}}$ x 100%
= 17.5%

Non metals

Hydrogen gas

Formula: H_2

At room temperature and pressure, hydrogen gas exists as a diatomic molecule, H_2

Occurrence

Hydrogen occurs in water, oils and natural gas

Preparation of hydrogen gas

1. **Laboratory preparation of hydrogen**

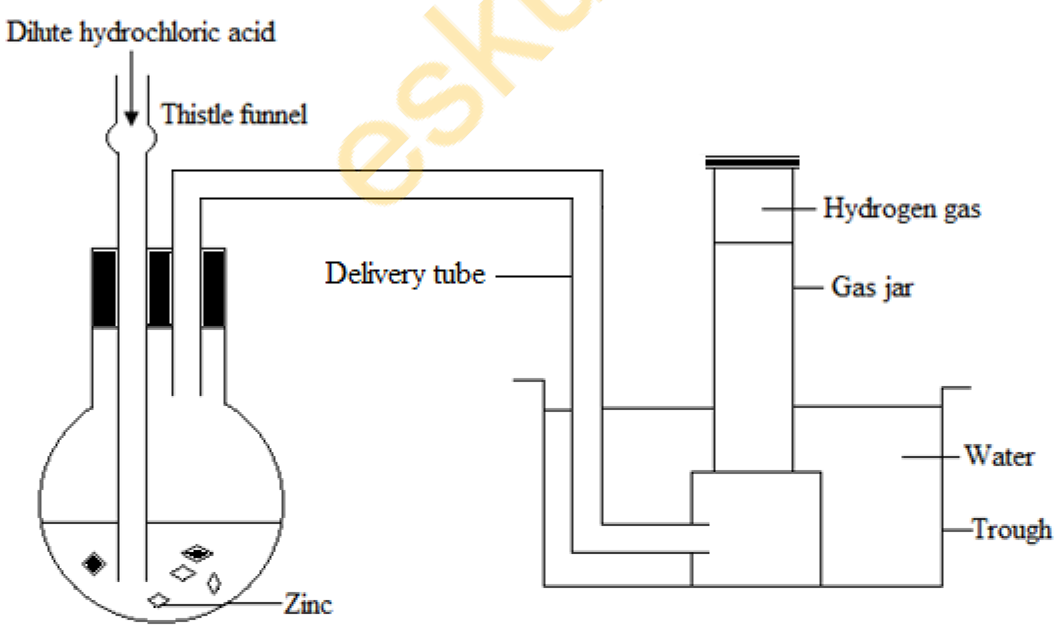
In the laboratory, hydrogen gas can be prepared by reacting:

- a reactive metal with a dilute acid
- a reactive metal with water

[A]Reaction of reactive metal with a dilute acid

Example

$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$



The gas is collected over water

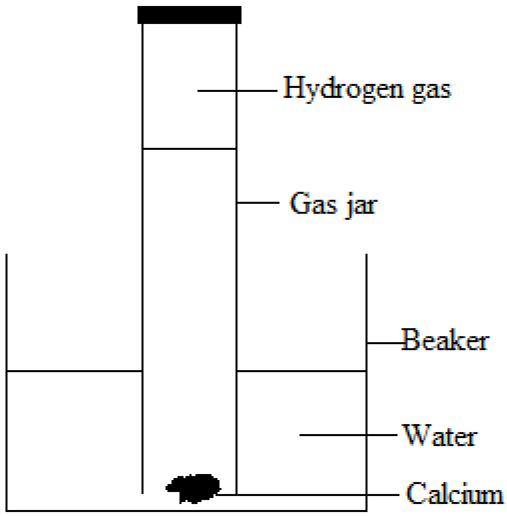
Drying agent: Concentrated Sulphuric acid

Method of collection: Down ward displacement of air or upward delivery since it is less dense

[B] Reaction of a reactive metal with water

Example

$Ca_{(s)} + 2H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$



Note

- Potassium and sodium float on water as they react with it, they must be wrapped in a wire mesh to make them sink.
- The reaction of potassium with water produces a lilac flame and the reaction of sodium with water produces a yellow flame

2. **Industrial preparation of hydrogen**

On the scale, hydrogen gas is manufactured by the reaction of methane and steam

$CH_{4(g)} + H_2O_{(g)} \rightarrow CO_{(g)} + 3H_{2(g)}$

Conditions

1. Catalyst: Nickel metal
2. Temperature: 1000°C
3. Pressure: 50 atmospheres

More steam is then added and the gases are passed over a catalyst iron (III) oxide to remove the carbon monoxide

$CO_{(g)} + H_2O_{(g)} \rightleftharpoons CO_{2(g)} + H_{2(g)}$

Carbon dioxide is removed by dissolving it in water under pressure

3. Hydrogen can also be produced by cracking of alkanes

Test for hydrogen

Hydrogen gas burns with a pop sound when a burning splint is introduced to it

Physical properties of hydrogen

1. It is colourless
2. It is odorless
3. It is less dense than air
4. It has a boiling point of -253^{o}C
5. It is not poisonous and does not support life

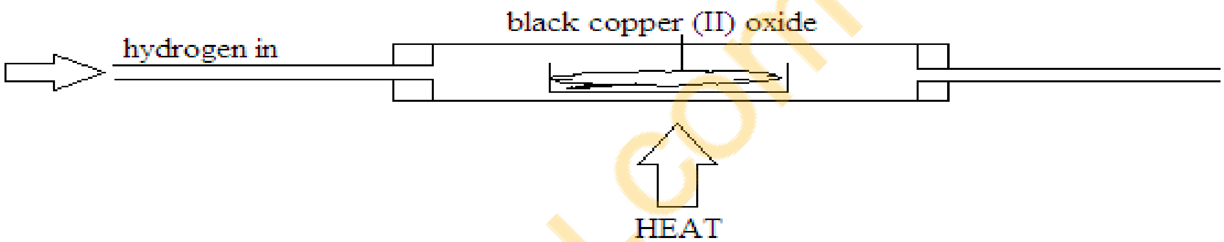
Chemical properties of hydrogen

1. It has no effect on litmus paper
2. It burns in oxygen with a blue flame producing a pop sound

   $2H_{2(g)}$ + $O_{2(g)}$ $2H_2O_{(g)}$
3. It is a reducing agent. It reduces the oxides of metals below it in the reactivity series

Example

$CuO_{(s)} + H_{2(g)} \rightarrow Cu_{(s)} + H_2O_{(g)}$



When hydrogen is passed over black copper (II) oxide in the apparatus above, the black powder turns pink

4. Hydrogen burns in chlorine with a white flame forming hydrogen chloride

Example

$H_{(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$

A mixture of hydrogen and chlorine is explosive in sun light

Uses of hydrogen gas

1. It is used in the manufacture of ammonia in haber process
2. It is used in the manufacture of margarine from vegetable oil in the process called catalytic hydrogenation
3. Liquid hydrogen is used as a fuel in rockets because it does not pollute the air.
4. It is used as a reducing agent

Exercise

1. A small piece of potassium is dropped on to called water and mixed with a little universal indicator.

(a) When potassium reacts with water, hydrogen gas is produced

(I) What name is given to reactions that produce heat?
(II) What colour does the universal indicator become?
(III) Describe a test for hydrogen

(b) (I) How do the observations differ when sodium is used instead of potassium?
(III) Write a balanced chemical equation for the reaction between sodium with water.

Oxygen gas

Formula: O_2

Oxygen gas exits as a diatomic molecule, O_2

It makes up about 21% of air by volume

Preparation of oxygen

1. Laboratory preparation of oxygen

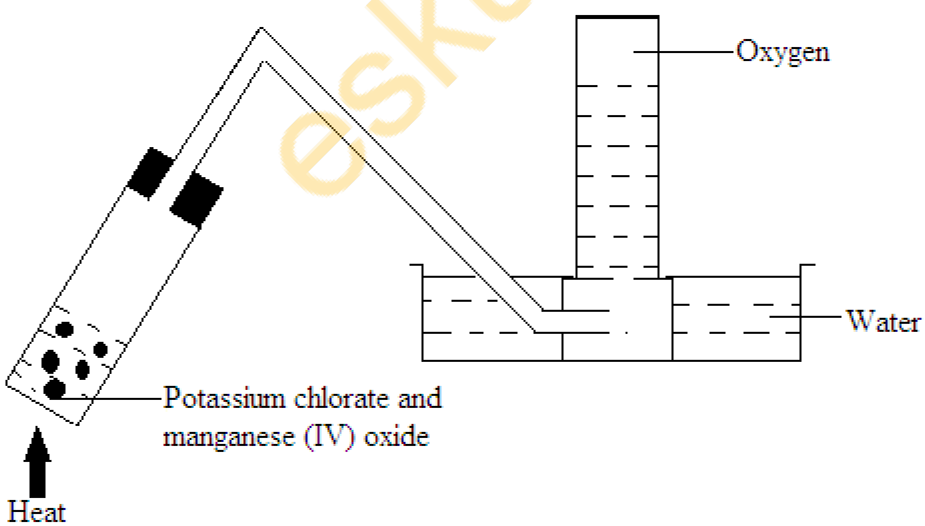
In the laboratory, oxygen gas can be prepared using the following chemicals:

(a) Potassium chlorate, $KClO_3$
(b) Hydrogen peroxide, H_2O_2
(c) Sodium nitrate, $NaNO_3$ and Potassium nitrate, KNO_3

[A] Potassium chlorate

Potassium chlorate is mixed with a catalyst manganese (IV) oxide upon heating and decomposes into potassium chloride and oxygen gas

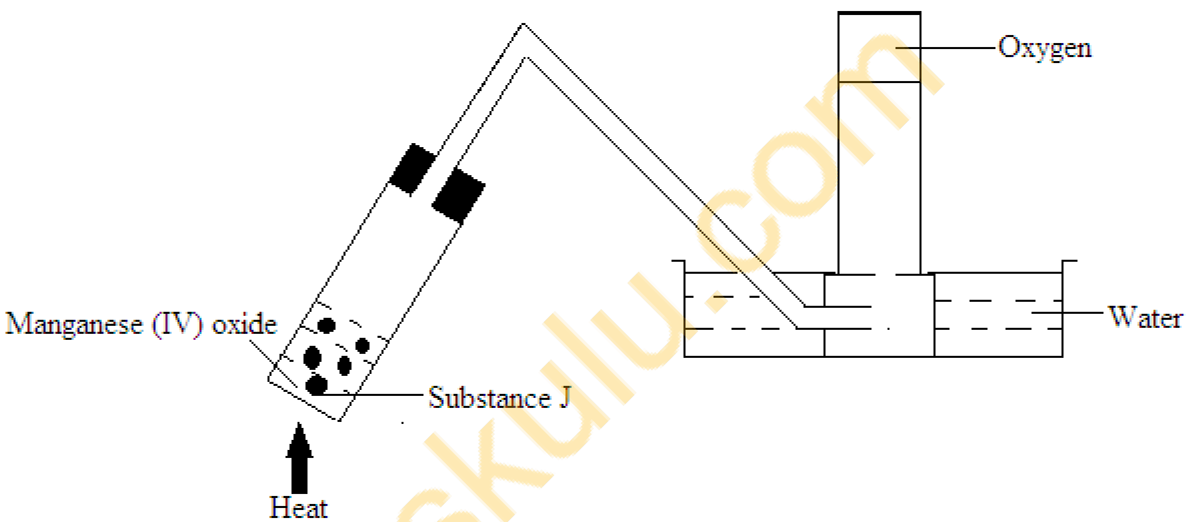
$2KClO_{3(s)} \rightarrow 2KCl_{(s)} + 3O_{2(g)}$



Note: This experiment is explosive

Example

1. Study the diagram below that shows preparation and collection of oxygen gas



(a) Name substance J
(b) Write a balanced chemical equation for the reaction in the experiment
(c) What is the method of collecting the gas used in the experiment?
(d) What is the function of the manganese (IV) oxide in the experiment?
(e) Explain why the first gas collected in the gas jar was not pure oxygen?
(f) State one industrial use of oxygen

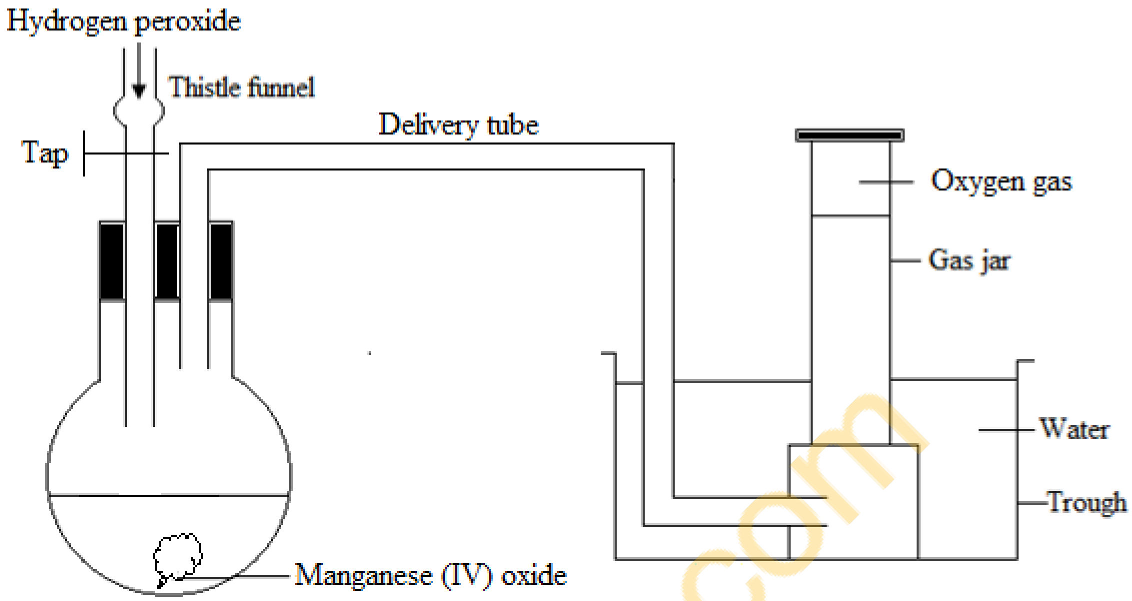
Solution

(a) Potassium chlorate
(b) $2KClO_{3(s)} \rightarrow 2KCl_{(s)} + 3O_{2(g)}$
(c) Down ward displacement of water
(d) Acts as a catalyst
(e) It is contaminated with air already in the apparatus
(f) Welding (as oxyacetylene flame) etc.

[B] Hydrogen peroxide

Oxygen gas is also prepared by the decomposition of hydrogen peroxide solution using manganese (IV) oxide as a catalyst. No heating is required

$2H_2O_{2(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$



Oxygen gas is collected over water

Method of collection: Down ward displacement of water

Drying agent: Concentrated sulphuric acid

[C] Sodium nitrate and potassium nitrate

Oxygen gas can be prepared by heating sodium nitrate or potassium nitrate

(I) $2NaNO_{3(s)} \rightarrow 2NaNO_{2(s)} + O_{2(g)}$

(II) $2KNO_{s(s)} \rightarrow 2KNO_{2(s)} + O_{2(g)}$

2. **Industrial preparation of oxygen gas**

On the large scale, oxygen gas is manufactured by fractional distillation of liquid air. Liquid oxygen boils at -183°C while liquid nitrogen boils at -196°C

Nitrogen gas which has a the lower boiling point distils off first

Test for oxygen gas

Oxygen gas relights a glowing splint introduced in its container

Uses of oxygen

1. It is used in the manufacture of steel in the blast furnace
2. It is used in oxygen tents in hospitals for patients in the intensive care unit
3. It used for welding in the oxy-acetylene flame
4. It used by deep sea divers and mountain climbers
5. It is used as liquid oxygen in rockets when in outer space to support burning of hydrogen

Physical properties of oxygen

1. It is colorless
2. It is odorless
3. It is slightly soluble in water
4. It is slightly denser than air
5. It supports burning
6. It boils at $-183^{o}C$
7. It does not burn

Chemical properties of oxygen

1. **Respiration**

Respiration is the process by living organisms oxidize glucose to produce carbon dioxide, water and energy

$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O + \text{energy}$

2. **Combustion**

Alternative term: Burning

Combustion is the process by which a substance reacts with oxygen to produce an oxide and heat. It is also defined as the burning of substances in oxygen.

Example

Carbon burns in oxygen to form carbon dioxide and heat energy.

$C + O_2 \rightarrow CO_2 + \text{energy}$

3. **Rusting**

Rusting is the corrosion of iron.

Corrosion is a process by which something, especially a metal, is destroyed progressively by chemical action, as iron when it rusts

It is an electrochemical process by which iron corrodes in the presence of oxygen, water and an electrolyte.

Rust

Rust is a reddish brown coating of iron oxide on the surface of iron or steel that forms when the metal is exposed to air and moisture

Chemical formula: $Fe_2O_3.xH_2O$

Chemical name: Hydrated iron (III) oxide.

Colour: Reddish – brown

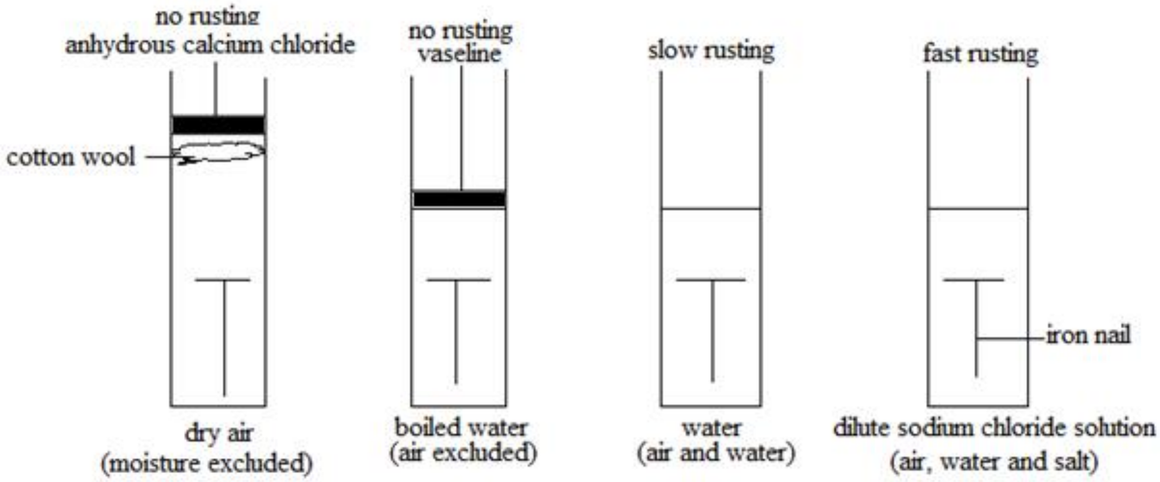
Formation of rust

Iron reacts with oxygen in air in the presence of water to form rust.

Overall reaction: $4Fe_{(s)} + 3O_{2(g)} + 2xH_2O_{(l)} \rightarrow 2Fe_2O_3.xH_2O_{(s)}$

Conditions necessary for iron to rust

1. Presence of oxygen
2. Presence of water or moisture
3. Presence of a strong electrolyte such as sodium chloride or sulphuric acid



Prevention of rusting

1. **Painting**

A paint coat excludes both air and water from contact with iron

2. **Oiling**

Oil or grease may be applied on the surface of iron metal to prevent water and oxygen from contact with iron

3. **Galvanizing**

Galvanizing is the coating of iron with zinc metal. Zinc is higher than iron in the reactivity series of metals; so if the surface is scratched, the zinc is oxidized in preference to iron. This is called sacrificial protection.

4. **Electroplating or alloying**

Iron can be electroplated or alloyed with non-corrosive elements like nickel, copper, chromium and carbon. The coating of iron with a metal which does not corrode easily protects iron from rusting. An electric current is used in electroplating.

Exercise

1. Burning and rusting are two chemical processes

(a) Give four ways by which rusting of iron metal can be prevented

(b) Give one in which rusting and burning :

(I) resemble one another

(II) differ from one another

(c) Explain how you could show the presence of iron metal in a sample of rust

2. Oxygen may be prepared by the decomposition of hydrogen peroxide

(a) Explain the term decomposition

(b) Name the catalyst used in this preparation

(c) Draw the apparatus which could be used to prepare and collect the oxygen produced

(d) On the industrial scale, oxygen is separated from nitrogen by fractional distillation.
   (I) Explain what is meant by fractional distillation and name one other commercial process in which it is used.
   (II) What is the chemical name for rust?
   (III) State any conditions which are needed for iron metal to rust
   (IV) Mention and describe any three methods of rust prevention

Sulphur dioxide

Formula: SO_2

Sulphur burns in oxygen with a blue flame to form sulphur dioxide

Properties of sulphur dioxide

1. It has a smell
2. It does not burn or support burning
3. It is fairly soluble in water forming sulphurous acid
4. It decolourizes the solution of acidified potassium permanganate since it is a reducing agent
5. It turns aqueous potassium dichromate (VI) from orange to green
6. It is denser than air.

Uses of sulphur dioxide

1. It is used as a bleaching agent in the manufacture of wood pulp for paper
2. It is used as a food preservative since it kills bacteria
3. It is used in the manufacture of sulphuric acid in contact process

Manufacture of sulphuric acid

Sulphuric acid is manufactured on large scale by contact process.

Raw material when making sulphuric acid

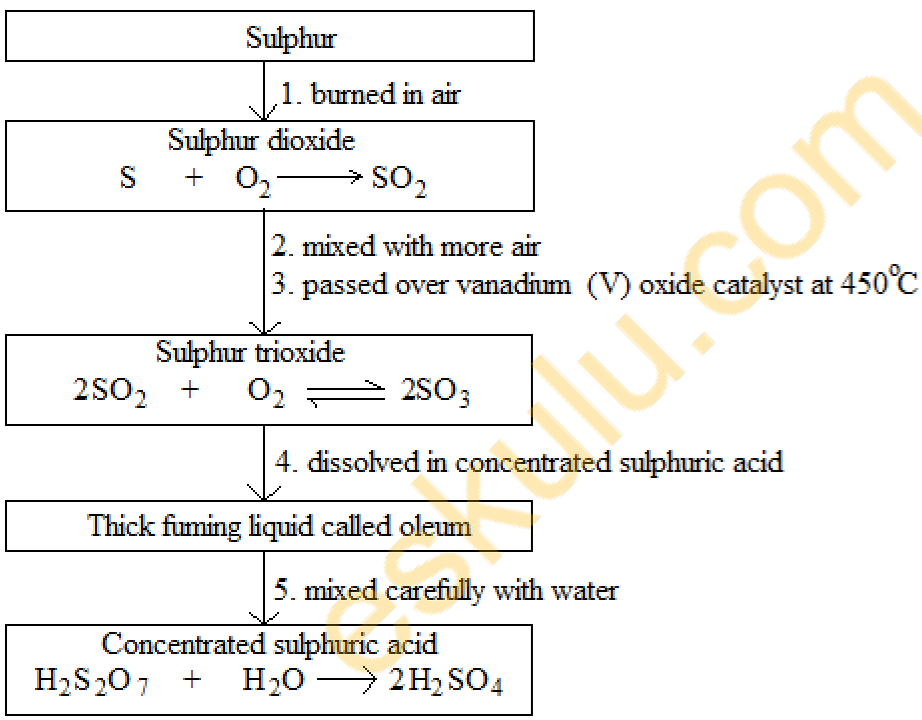
- Sulphur
- Air
- Water

The following steps are involved:

1. Sulphur reacts with oxygen in air to form sulphur dioxide.
   $S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$
2. Sulphur dioxide reacts with more oxygen in air to form sulphur trioxide.
   $2SO_2 + O_2 \rightarrow 2SO_3$

3. **Note:** Sulphur dioxide is passed over vanadium (V) oxide catalyst at a temperature of 450^oC. The catalyst will not work below 400^oC. So at lower temperatures, the reaction will be too slow. And at higher temperatures, the yield of sulphur trioxide drops. So the reaction is carried out at 450^oC as a compromise.
4. Sulphur trioxide is dissolved in cold concentrated sulphuric acid to form an oily liquid called oleum, $H_2S_2O_7$
   $SO_3 \quad + \quad H_2SO_4 \quad \rightarrow H_2S_2O_7$
5. Oleum is then converted to sulphuric acid by diluting it with water.
   $H_2S_2O_7 \ + \ H_2O \ \rightarrow \ 2H_2SO_4$

Summary



Diluting sulphuric acid

During dilution, sulphuric acid is added to water and not the other way round i.e.do not add water to acid because this results in an explosion since the specific heat capacity of sulphuric acid is small and hence the heat evolved will boil the solution.

Properties of sulphuric acid

1. It is a strong electrolyte

2. It reacts with carbonates and hydrogen carbonates to produce a salt, water and carbon dioxide
3. It neutralizes bases and alkaline solutions
4. It reacts with metals above hydrogen in the reactivity series to form a salt and hydrogen gas
5. It displaces less volatile acids like hydrochloric acid and nitric acid from their salts
6. It is a strong dehydrating agent since it removes water from other compounds e.g.
   $H_2SO_4 + C_{12}H_{22}O_{11} \rightarrow 12C + 11H_2O$
7. It is a powerful oxidizing agent

Uses of sulphuric acid

1. It is used in the manufacture of detergents and soap
2. It is used in the manufacture of fertilizers like ammonium sulphate
3. It is used as an electrolyte in car batteries
4. It is used in refining of petroleum
5. It is used as a drying agent when concentrated

Exercise

1. The following equation is for the contact process

$$2SO_2 + O_2 \rightleftharpoons 2SO_3$$

The table below shows the percentage conversion to sulphur trioxide by this process at a number of different temperatures at constant pressure

Temperature in ^{o}C	Percentage of SO_3
350	99
450	97
550	86
650	55

(a) The process is carried out at 450^{o}C. Suggest why this temperature has been chosen

(b) The contact process uses vanadium (V) oxide as a catalyst.

(I) Explain why the catalyst is used

(II) Explain why the catalyst used does not increase the percentage of sulphur trioxide produced

Chlorine gas

Formula: Cl_2

Chlorine exists as diatomic molecules, Cl_2

Properties of chlorine gas

1. It is a greenish yellow gas with a chocking smell.

2. It is poisonous
3. It is sparingly soluble in water
4. It bleaches damp litmus paper

Preparation of chlorine

Chlorine is prepared by the action of oxidizing agents on hot concentrated hydrochloric acid solution

$2KMnO_4 + 16HCl \rightarrow 2KCl + 2MnCl_2 + 5Cl_2 + 8H_2O$

Manufacture of chlorine

Chlorine is manufactured by the electrolysis of concentrated sodium chloride or brine in the mercury cathode cell. Sodium hydroxide is obtained as a by product

Sources of sodium chloride include sea water and rock salt

Uses of chlorine

1. It is used in the sterilization of drinking water and water in swimming pools
2. It is used in the manufacture of bleaching agents and pesticides
3. It is used to make polyvinylchloride (PVC)
4. It is used in the manufacture of hydrochloric acid

Silicon

Formula: Si

Silicon is a non - metal in group IV of the periodic table. It is a common element in most rocks as it is found in combination with oxygen. It is commonly found as silicon dioxide or silica e.g. quartz and sand which is an impure form of silicates e.g. mica and calcium silicate

Uses of silicon

1. It is used in the manufacture of semi – conductors dioxides for radios and televisions and other electronic devices
2. It is used in the manufacture of silicones. Silicones are macromolecules of polymers of silicon, oxygen and alkyl group of organic compounds

Different silicones are able to exist as oils, waxes or plastics. Silicone plastics are fire resistant since they have to form sand and carbon dioxide when burnt in oxygen

Uses of sand

1. It is used in making glass
2. It is used in manufacture of silicon by reduction
3. It is used in the construction of buildings when mixed with cement and water

Nitrogen gas

Formula: N_2

Nitrogen exists as a diatomic molecule, N_2

Physical properties of nitrogen

1. It is colourless
2. It is odorless
3. It does not burn and does not support burning except magnesium which burns with it.

Chemically, nitrogen gas is very unreactive due to the strong bonds between the nitrogen atoms in the molecule However, under forced conditions; nitrogen reacts with hydrogen gas to form ammonia in the presence of the iron catalyst

Manufacture of nitrogen gas

Nitrogen gas is manufactured by fractional distillation of liquid air

Uses of nitrogen

1. It is used in the manufacture of ammonia
2. It is used to provide an inert atmosphere during the arc welding of aircraft parts and in electric bulbs.

The need for nitrogen, phosphorus and potassium compounds in plant life

Nitrogen is essential in the manufacture of proteins while phosphorus is mainly needed for enzyme activity. Potassium is important in the controlling of rates of photosynthesis and respiration.

Ammonia

Formula: NH_3

Physical state: Gas

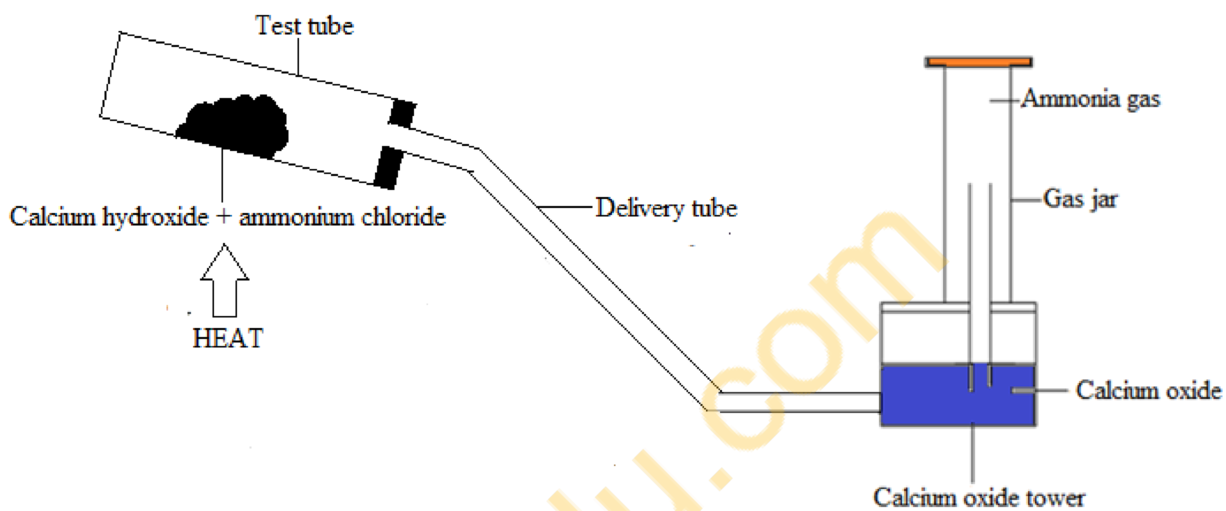
Preparation of ammonia

[A] Laboratory preparation of ammonia

In the laboratory, ammonia gas can be prepared by the action of heating an alkali with an ammonium compound.

Example

$Ca(OH)_{2(aq)} \quad + \quad 2NH_4Cl_{(aq)} \rightarrow \quad CaCl_{2(aq)} \quad + \quad 2H_2O_{(l)} \quad + \quad 2NH_{3(g)}$



The test tube must be tilted downwards to prevent the water formed from running back into the reaction mixture which may boil and cause an explosion.

Method of collection: Ammonia gas is collected by down ward displacement of air because it is less dense than air.

Drying agent used: Calcium oxide

Drying agents like calcium chloride and concentrated sulphuric acid should not be used because they easily react with ammonia.

Physical properties of ammonia

1. It is a colourless gas with a pungent choking smell
2. It is less dense than air
3. It is very soluble in water
4. It is easily liquefied, either by cooling to -33^{o}C or by compressing. This makes it easy to transport in tanks and cylinders.

Chemical properties of ammonia

1. It turns damp red litmus paper blue
2. It burns in pure oxygen with a yellow - brown flame
3. It dissolves in water to form ammonium hydroxide

   $NH_{3(g)} + H_2O_{(l)} \rightarrow NH_4OH_{(aq)}$

Note

$NH_{3(g)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$

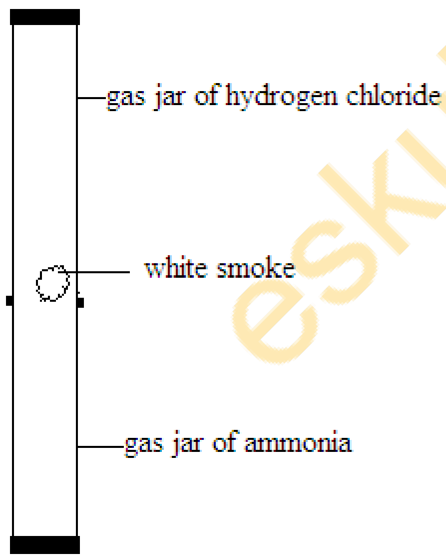
$NH_4OH_{(aq)} \rightarrow NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$

4. It is a reducing agent because it reduces oxides to metals low in the reactivity series of metals.

Example

$3CuO_{(s)} + 2NH_{3(g)} \rightarrow 3Cu_{(s)} + 3H_2O_{(l)} + N_{2(g)}$

5. It reacts with hydrogen chloride gas to form a white smoke.



The white smoke is made of tiny particles of solid ammonium chloride:

$NH_{3(g)} + HCl_{(g)} \rightarrow NH_4Cl_{(s)}$

Test for ammonia gas

Ammonia gas turns damp red litmus paper blue. It is alkaline in nature

[B] Industrial preparation of ammonia

The industrial preparation of ammonia is called Haber process

Ammonia gas produced industrially from nitrogen and hydrogen

Importance of Haber process

Haber process is important because it produces ammonia on a large scale

Raw materials for Haber process

- Nitrogen
- Hydrogen

Sources or raw for nitrogen and hydrogen

1. Nitrogen
   Source: Liquid air (obtained by fractional distillation)

Nitrogen is obtained by burning hydrogen in air. Air is mostly nitrogen and oxygen, with small amounts of other gases. Only the oxygen reacts with hydrogen, forming steam:

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

When the steam condenses, the gas that remains is mainly nitrogen.

2. Hydrogen
   Source: Methane **(**natural gas)
   Steam

Hydrogen is made from methane and steam.

$CH_{4(g)} + 2H_2O_{(g)} \rightarrow CO_{2(g)} + 4H_{2(g)}$

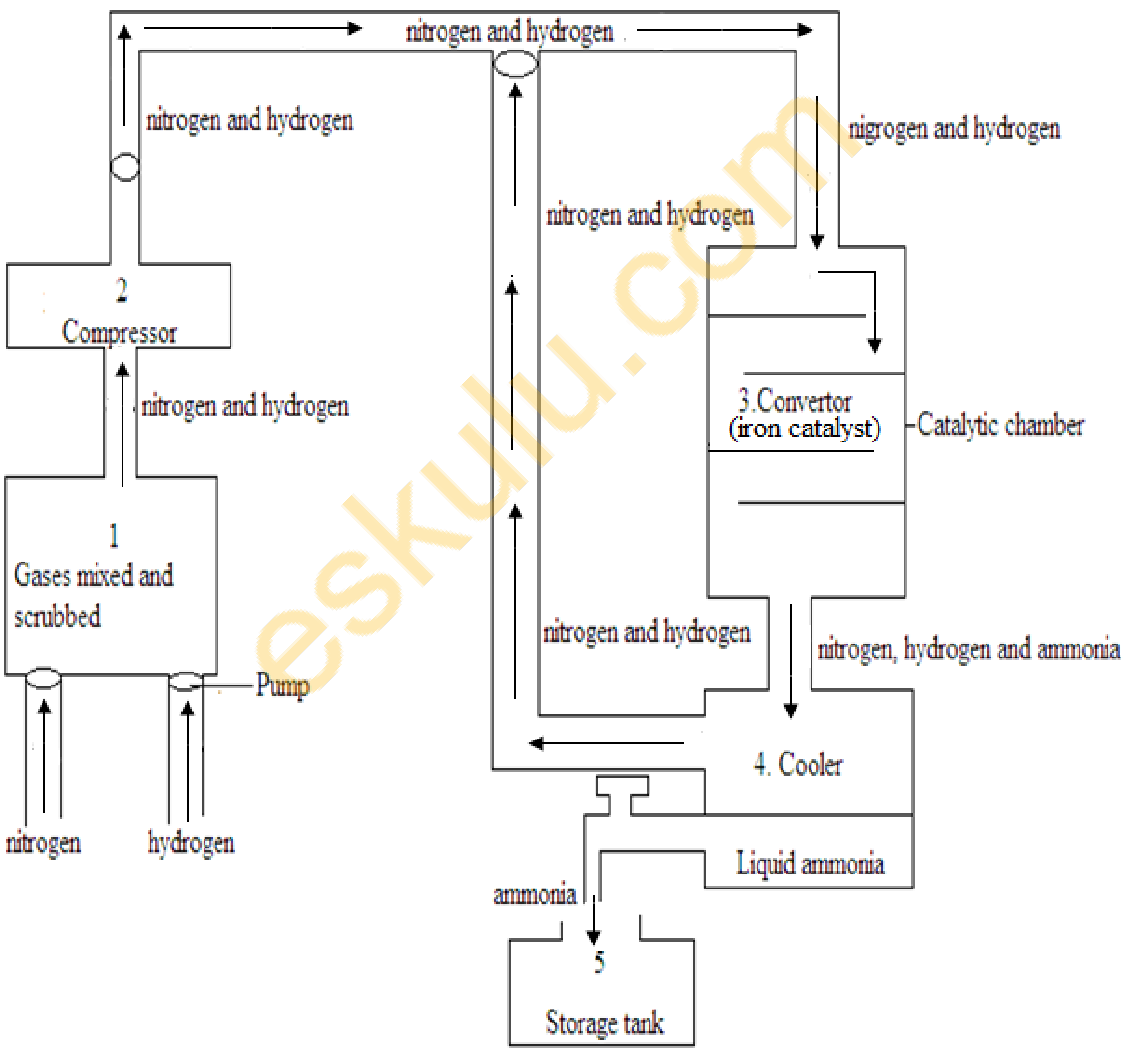
Reaction between nitrogen and hydrogen

$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

Essential conditions needed for the manufacture of ammonia in haber process

1. Catalyst: Iron
2. Temperature: 350^{o}C to 450^{o}C
3. Pressure: 350 atmospheres

The reaction scheme



Nitrogen is unreactive. To make it react with hydrogen, a process called Haber process is used:

1. The two gases are mixed. The mixture is cleaned or scrubbed, to get rid of any impurities
2. Next the mixture is compressed. This pushes the gas molecules closer together
3. Then the mixture goes to the convertor. This is a round tank containing beds of hot iron. The iron is a catalyst for this reaction:

   $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

   The double headed arrow in the equation means the reaction is reversible. So it does not go to completion. A mixture of nitrogen, hydrogen and ammonia leaves the convertor.
4. The mixture is cooled until the ammonia condenses. At the end of the reaction, about 10% of ammonia is produced. The unreacted gases of nitrogen and hydrogen are pumped back to the convertor (iron catalyst) for another chance to react.
5. The ammonia is run into tanks and stored as a liquid, under pressure.

Uses of ammonia

1. It is used in the refrigerating plants as a refrigerant
2. It is used in the manufacture of explosives
3. It is used in the manufacture of plastics and glue
4. It is used in the manufacture of nitric acid
5. It is used as a cleaning agent
6. It is used in the manufacture of fertilizers such as ammonium sulphate, ammonium nitrate and ammonium phosphate.

Examples

(a) $2NH_{3(g)} + H_2SO_{4(aq)} \rightarrow (NH_4)_2SO_{4(aq)}$

(b) $NH_{3(g)} + HNO_{3(aq)} \rightarrow NH_4NO_{3(aq)}$

(c) $3NH_{3(g)} + H_3PO_{4(aq)} \rightarrow (NH_4)_3PO_{4(aq)}$

Ionic equations

1. $(NH_4)_2SO_{4(aq)} \rightarrow 2NH_4^+{}_{(aq)} + SO_4^{2-}{}_{(aq)}$
2. $NH_4NO_{3(aq)} \rightarrow NH_4^+{}_{(aq)} + NO_3^-{}_{(aq)}$
3. $(NH_4)_3PO_{4(aq)} \rightarrow 3NH_4^+{}_{(aq)} + PO_4^{3-}{}_{(aq)}$

Note

1. Ammonium sulphate is used as a fertilizer because it contains nitrogen and sulphur needed by plants
2. Ammonium nitrate is used as a fertilizer because it contains nitrogen needed by plants
3. Ammonium phosphate is used as a fertilizer because it contains nitrogen and phosphorous needed by plants

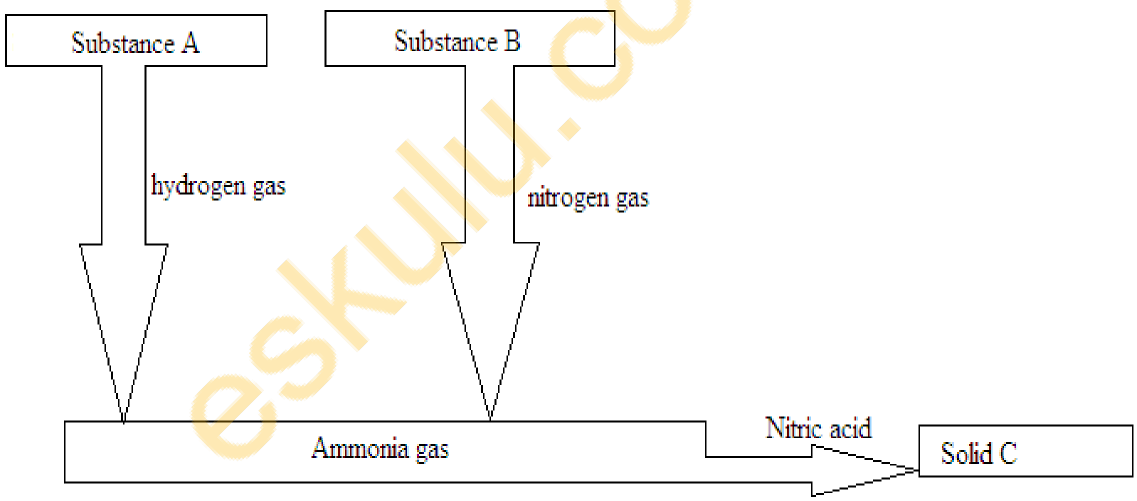
4. The three essential elements present in many fertilizers are nitrogen, phosphorous and potassium (N.P.K)

Problems with chemical fertilizers

- Chemical fertilizers tend to make the soil acidic if they are used for long periods of time
- Chemical fertilizers are easily leached away e.g. nitrates (NO_3^-), sulphate (SO_4^{2-}) and chlorides (Cl^-). They are leached because they easily dissolve in water and move with it. This is wasteful. Besides, when the rainwater drains into rivers, the nitrates cause pollution

Example

1. Study the reaction scheme below



(a) Identify substances A, B and C

(b) Ammonia is manufactured by direct synthesis in the Haber process. The reaction is reversible and releases 46kj of heat energy per mole of ammonia. Write a balanced chemical equation including all the information given.

(c) State three essential conditions which are needed for the manufacture of ammonia in Haber process

(d) What is the use for solid C

(e) Write down the chemical equation for the reaction between ammonia and nitric acid

(f) Write any two uses of ammonia besides the one shown above.

Solution

(a) A – Nitrogen
B – Hydrogen
C – Ammonium nitrate

(b) $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

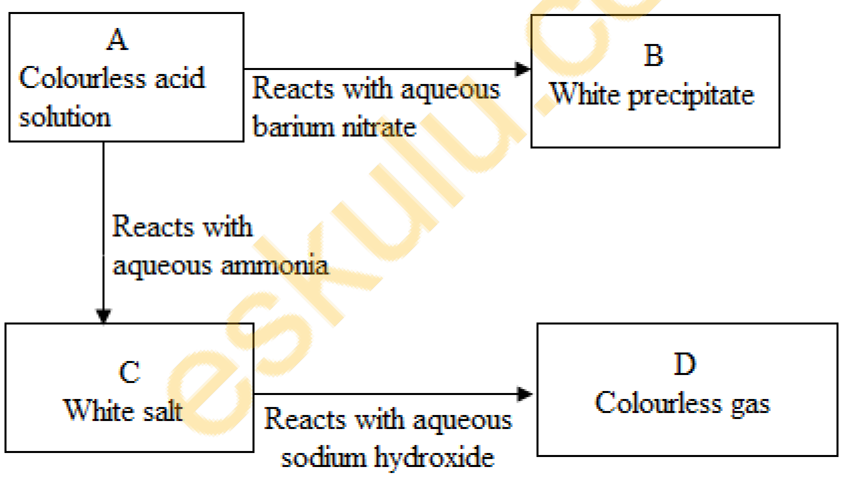
(c) Catalyst: Iron
Temperature: 350°C to 450°C
Pressure: 350 atmospheres

(d) It is used as a fertilizer because it contains nitrogen needed by plants

(e) $NH_{3(g)} + HNO_{3(aq)} \rightarrow NH_4NO_{3(aq)}$

(f) It is used in the refrigerating plants as a refrigerant.
It is used in the manufacture of explosives.
It is used in the manufacture of plastics and glue.
It is used as a cleaning agent.

2. The diagram below shows some of the properties and reactions of aqueous ammonia and some other substances



(a) Identify, by name, the substances A, B, C and D

(b) Write an ionic equation including symbols for the reaction of:

(I) A with aqueous barium nitrate to form B

(II) A with aqueous ammonia to form C

(III) C with aqueous sodium hydroxide

(c) Name a method you would use to collect gas D

Solution

(a) A – Sulphuric acid

B – Barium sulphate
C – Ammonium sulphate
D – Ammonia

(b) (I) $Ba^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)}$

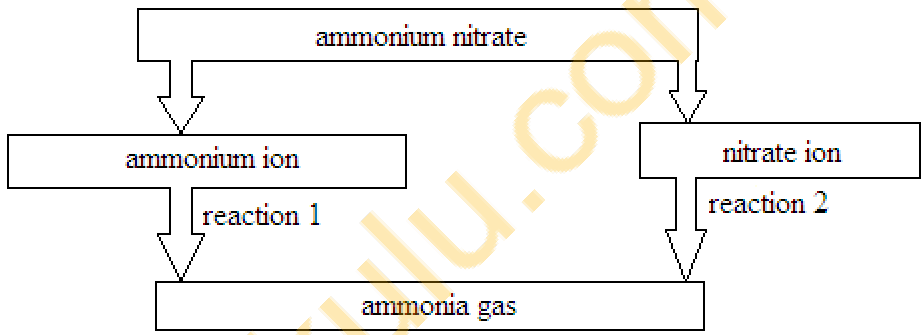
(II) $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

(III) $NH_4^+{}_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + NH_{3(g)}$

(c) Down ward displacement of air.

Exercise

1. Ammonium nitrate contains ammonium ions (NH_4^+) and nitrate ions (NO_3^-)
   (a) Name two substance which are used to manufacture ammonium nitrate
   (b) Explain why ammonium nitrate can act as a fertilizer
   (c) The diagram below shows ammonium ions (NH_4^+) and nitrate ions (NO_3^-) can be changed into ammonia.



(i) Describe how you would carry out reaction 1 and reaction 2
(ii) Write an ionic equation for reaction 1 including state symbols

2. In the industry, nitrogen is combined with hydrogen to form ammonia
   (a) From what raw materials is the hydrogen extracted?
   (b) Explain why haber process is important?
3. Name three essential elements present in many fertilizers
   (a) Name two compounds which can be used as fertilizers
   (b) Explain why after fertilizers have been added to the soil heavy rains can be a disadvantage.
4. Ammonia is an important compound used in making fertilizers. It is manufactured by direct synthesis by haber process. Ammonia is a base – it neutralize acids
   (a) State the raw materials for the haber process
   (b) Write a balanced chemical equation for the formation of ammonia by the haber process

(c) Ammonia combines with nitric acid to form ammonium nitrate, which is the most important nitrogenous fertilizer

$NH_{3(g)} + HNO_{3(aq)} \rightarrow NH_4NO_{3(aq)}$

(I) Calculate the percentage by mass of nitrogen in ammonium nitrate

(II) Why is ammonium nitrate used as a fertilizer?

(d) Apart from nitrogen, name two other elements which are essential for plant growth

Pollution

pollution is the contamination of the environmet with harmful substances. The harmful substances are called pollutants.

Types of pollution

There are three main types of pollution.

- Land pollution
- Air pollution
- Water pollution

[A] Land pollution

Land pollution occurs when pollutants are added to the land. The major causes of land pollution are mateials which do not rot or decay. These materials resist bacterial action and are said to be non biodagradable.

Examples of land pollutans

- Plastics
- Glass

[B] water pollution

Water polution occurs when pollutants are added to water.

Examples of water pollutants

1. Oil spillage: Oil spillage by oil tankers and leaking engines of speed boats pollutes the water. It disturbes marine life since oxygen supply is cut off.

Solution: You can scoop the oil from the surface of the water

2. Weeds: Weeds usually prevent smooth transport and make H.E.P genearation very difficult.

Solution: You can cut and remove all the weeds from the water

3. Industrial wastes: Acidic industrial effluents make the water slightly acidic. This disturbes the balance of living things .

Solution: Industrial wastes should be treated with slaked lime to neutralize them. You can also set up dumping sites

4. Acid rain: Acid rain reduces the PH of the water

5. Debris: Debris include solid particles and plant parts

[C] Air pollution

Air pollution occurs when pollutants are added to the atmosphere

Examples of air pollutants

- Carbon monoxide
- Sulphur dioxide
- Oxides of nitrogen
- Soot
- Lead compounds

Carbon monoxide

Formula: CO

Source: Incomplete combustion of fuel in car exhausts.

Carbon monoxide is produced by the incomplete combustion of carbon containing compounds in limited suuply of air such as in engines of vehicles and blast furnace

Effects: When carbon monoxide is inhaled in the body, it enters the blood stream; it combines with haemoglobin making it difficult for red blood cells to tansnport oxygen. This may lead to suffocation, unconsciousness or death.

Solution: Use alternative source of fuel like hydrogen which does not pollute the air. Avoid incomplete combustion of fuels

Sulphur dioxide

Formula: SO_2

Source: Combustion of fossil fuels.

Sulphur dioxide is produced as a result of the combination of fossil fuels containing sulphur compounds. Sulphur dissolve in rain water which come down as acid rain

$$SO_{2(g)} + H_2O_{(l)} \rightarrow H_2SO_{3(aq)}$$

Adverse effects of acid rain

Acid rain has a lot of adverse effects on buildings, roofing sheets, monuments, vegetation and on the lungs

1. Acid rain make the soil to become acidic and hence destry vegetation
2. Acid rain dissolve monuments (statues) made of calcium carbonate
3. Acid rain erodes buildings and iron roofing sheets. Iron roofing sheets rust rapidly because of the presence of an electrolyte
4. Sulphur dioxide combines with water vapour in the lungs to form an acid. The acid can cause temperary or permanently damaged to the lungs

Carbon dioxide

Formula: CO_2

Source: Complete combustion of fuels like petrol and charcoal

Effects: Depletes the ozone layer leading to global warming. It also causes causes acid rain.

Solution: Massive afforestation

Oxides of nitrogen

Source: Combustion of fuels in car exhaust.

Oxides of nitrogen are found in car exhaust gases.

Effects: Compounds of nirogen form cumulative poison in tissues of living things. Nitrogen combines with oxygen at high temperature in the cylinder of car engines to form nitrogen monoxide which is very poisonous. Nitrogen monoxide damage lungs and irritate eyes.

Lead compounds

Source: Combustion of fuels in car exhausts.

Lead compounds are found in car exhaust gases.

Effcects: When lead is taken into the body, it may settle and be stored in the bones or membranes of the brain. This lead to damaging of the bones, brain and the nervous system.

Quantities of lead in the blood as low as 100 micrograms per cubic decimeter can cause children to have low intelligence, poorer memories and less mascular coordination than children who do not carry that burden of lead.

Petrol contains about $0.5g/dm^3$ to $1g/dm^3$ of lead. Lead poisoning is an example of global poisoning.

Solution: Lead tetraethyl is added to petro as an inhibitor to prevent knocking

Ozone

Formula: O_3

Source: Photochemical smog

Effects: At ground level, ozone is an eye irritant, causes breathing problems and damages plants.

Example

1. Ozone, O_3, is an atmospheric pollutant in the lower atmosphere but is beneficial higher up in the atmosphere.

(a) How is ozone formed in the lower atmosphere?

(b) Ozone in the upper atmosphere is being depleted. Describe briefly how this is happening and some of the health problems caused by ozone depletion.

(c) At room temperature ozone decomposes slowly to form oxygen, O_2.
The decomposition can be represented by the equation below. The reaction is exothermic. One mole of ozone will release 143 kJ when it is fully decomposed.
$2O_3 \rightarrow 3O_2$

(I) In terms of the energy changes that take place during bond breaking and bond making, explain why this reaction is exothermic.

(II) Explain why the **rate** of this decomposition increases as the **temperature** increases.

(III) Calculate the energy released when 16 g of ozone is decomposed.

Solution

(a) Ozone is formed by photochemical reactions (or sparks in air, ultraviolet on O_2)
(b) Ozone removed by reaction with chlorine (atoms) and is derived from CFC's
Ozone loss causes skin cancers or cataracts or crop damage or skin diseases or eye damage
(c) (I) Bond breaking is endothermic/absorbs energy and bond forming is exothermic/releases energy more energy released than absorbed
(II) As temperature increases molecules move faster or increased K.E hence more frequent collisions or more molecules energy exceeds the activation energy.

(III) $2O_3 \rightarrow 3O_2$

$n = \frac{m}{MM}$

$n = \frac{16g}{48g/mol}$

n = 0.33mol

1mol O_3 → 143 kJ
0.33mol O_3 → x

$$\frac{0.33 \text{ mol } O_3 \text{ x } 143kJ}{1 \text{ mol } O_3}$$

x = 47.7 kJ

General solutions to the problems of pollution

1. Recycling of waste products
2. Setting up protection standards. This involves limiting or banning of the manufacture and handling of certain products that are a source of pollution
3. Developing engines that use fuels which pollutes less and using catalytic convertors.
4. Setting up advisory boards to monitor pollution. Such boards would advise and give assistance on the disporsal of pollutants

Note: Fossil fuels must be conserved because they are limited and non-renewable energy sources.

Exrcise

1. The increase in industrial activity in Zambia has resulted in high level of air pollution.
   (a) What is meant by air pollution?
   (b) Name two common air pollutants in Zambia and state what effect each pollutant has on either plant or animal life

(c) State one way in which the emission of one of the pollutants you have named can be minimised.
(d) When limestone is heated strongly, quiclime, CaO, is formed according to the following chemical equation:
$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$
What would be observed if the gas produced above was passed through lime water until there was no further change?
(e) Name two gases, other than sulphur dioxide, that pollute the atmosphere. Give one source of each of each pollutant. Sulphur dioxide is shown as an example of a pollutant
Pollutant: Sulphur
Source: Burning coal
(f) Give one reason for conserving fossil fuels
(g) What is damaged by the sulphur dioxide released when some fossil fuels are burnt. . Give two examples

Water

Chemical formula: H_2O

Test for water

If a liquid is water, it will:

(a) turn blue colbat chloride paper pink
(b) turn white anhydrous copper (II) sulphate blue
(c) boils at 100^oC and freezes at 0^oC at natutral pressure if it is pure water

Uses of water

(a) It is essential for all living things
(b) In homes, water is used for cooking, washing and flushing
(c) It is used as a solvent
(d) It is used for making beer, cement
(e) It is used for cooling machines in the industry

Water from the source (rivers and underground) is never completely pure because it contains the following impurities:

1. Bacteria and tiny organisms: Most bacteria are harmles but some can cause diseases

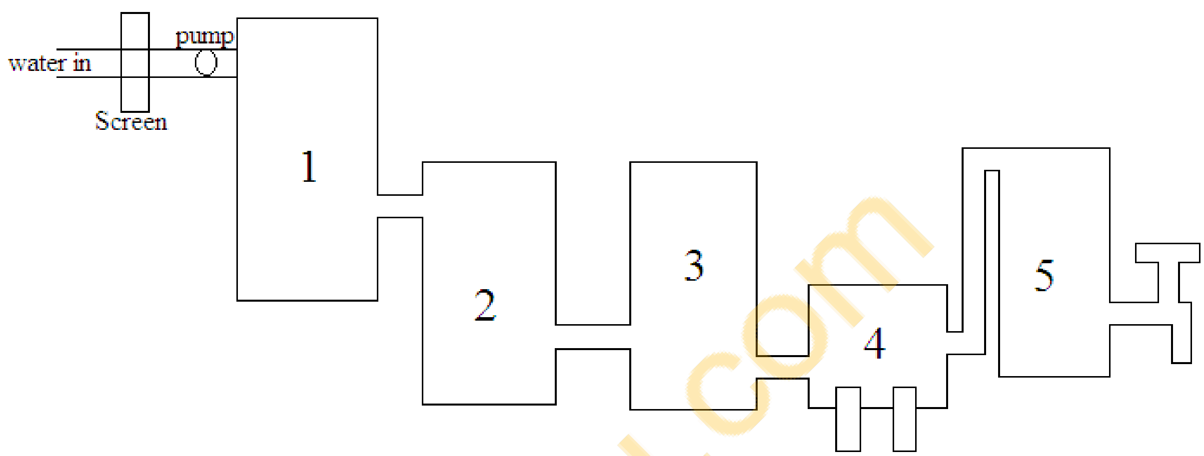
2. Dissolved substances: Nitrates and sulphates from the soil, gases from air and some calcium and magnesium compounds from the rocks

3. Solid particles: Solid particles include particles of mud, sand, grit, twigs, dead plants, dead animals and perhaps tins and rags that people have dumped.

Before the water is safe to drink, the bacteria and solid substances must be removed. This is done at the water works.

The water works - Purification of water at the water works

The diagram below shows what happens at the water works



The water is pumped through the screen which gets rid of large bits of rubbish. Then the water gets through following stages:

1. **Course sand filter**
   The course sand filter tank traps the larger particles of solid
2. **Sedimentation tank**
   Here chemicals are added to water, to make the smaller particles stick together. These particles then settle to the bottom of the tank.
3. **Fine sand filter**
   The fine sand filter traps any remaining particles.
4. **Chlorination tank**
   In this tank, a little chlorine gas is added. Chlorine gas dissolves, and kills any remaining bacteria. This is called disinfecting or sterilising the water. In some places, a flouride compound is also added to the water, to help prevent tooth decay. Water is now fit to drink.
5. **Storage tank**
   Water is now pumped into high storage tanks and from there piped to homes and facories

The water that flows from the taps is clean but not quite pure. It still contains dissolved substances which were not removed at the water works. Pure water can be made by distillation of tap or rain water

Two main stages in the purification of water

(a) **Filtration**: Solid paticles are removed from the water

(b) **Chlorination**: Chlorine is added to water to kill the bactera

Exercise

1. Water is life and it is an important raw material in the home and industry
   (a) Describe the identty test of pure water
   (b) Outline the two main stages in the purification of water
   (c) State two uses of water in the industry

Types of water

There are two types; soft water and hard water

1. Soft water

Soft water forms lather easily with soap

2. Hard water

Hard water does not form lather easily with soap. Soap gives a scum with hard water.

Hardness in water is caused by dissolved calcium and magnesium compounds e.g calcium hydrogen carbonate, magnesium hydrogen carbonate, calcium sulphate and magnesium sulphate

The scum form because the compounds react with soap and give an insoluble product that floats on water

Calcium sulphate + sodium stearate → calcium stearate + sodium sulphate

(Soap)   (Scum)

Temperory hardness

Temperory hardness is caused by the presence of calcium hydrogen carbonate in water. It occurs in limestone areas when rain water containing a small amount of dissolved carbon dioxide from the air passes over the limestone.

$$CaCO_{3(s)} + H_2O_{(l)} + CO_{2(g)} \rightarrow Ca(HCO_3)_{2(aq)}$$

Temperory hardness can be removed by boiling the water. The white diposite (fur) which is produced by boiling hard water is calcium carbonate

$Ca(HCO_3)_{2(aq)} \rightarrow CaCO_{3(s)} + H_2O_{(l)} + CO_{2(g)}$

Permanent hardness

Permanent hardness is caused by small amounts of dissolved calcium and magnesium compounds.

Permanent hardness cannot be removed by boiling

Removal of permanent hardness

[A]Distillation

This method is too expensive to produce large quatities of soft water

[B]Addition of sodium carbonate (washing soda)

Sodium carbonate precipitates the calcium and magnesium ions which cause hardness

$Na_2CO_{3(aq)} + CaSO_{4(aq)} \rightarrow CaCO_{3(s)} + Na_2SO_{4(aq)}$

$Na_2CO_{3(aq)} + MgSO_{4(aq)} \rightarrow MgCO_{3(s)} + Na_2SO_{4(aq)}$

This method removes both temperory and permanent hardness but it can damage wollen materials since water becomes alkaline

Advantages of hard water

1. It has a better or pleasant taste due to dissolved compounds
2. It contains calcium cmpounds which are good for bones and teeth
3. Doctors think it helps to prevent heart disease

Disadvantages of hard water

1. Leads to wastage of soap and production of scum
2. Leaves fur in kettles, pipes, boilers and radiators. This makes them less effective and may also cause blockages.

Carborn

Allotropes of carbon

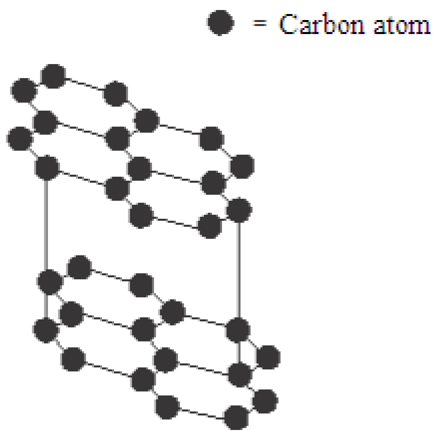
Definition: Allotopes are elements in different physical forms but in the same state
Allotropy is the existence of an element in different physical forms but in the same state
Carbon exists in the form of graphite, diamond and amorphous
Grapphite and diamond are allotropes of carbon.

[A] Graphite

Structure of graphite

Graphite is a soft, black, crystalline form of carbon that is a fair conductor of electricity.

It is made up of flat sheets of carbon atoms



Graphite

Each carbon atom is boneded to three others while the fourth electron is delocalised. Since each carbon atom forms covalent bonds to three other, this gives rings of six atoms that join to make flat sheets

The sheet of atoms lie on top of each other, held together by weak forces

Physical properties of graphite

1. It is a good conductor of electricity because of free moving electrons in between the layers of carbon atoms
2. It is soft and slippery. This is because the sheets of atoms can slide over each other easily.
3. It writes well on paper
4. It is has a density of 2.22g/cm^3
5. It has a high melting point. This is because the strong bonds between the carbon atoms within a layer make graphit difficult to pull apart in the direction of the layer.

Uses of graphite

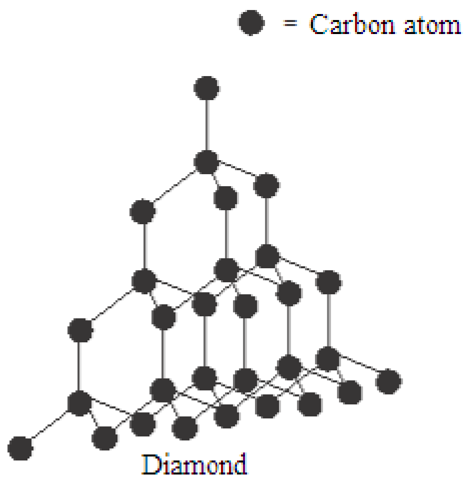
1. It is used as a lubricant because the layer of carbon atoms slide over each other easily
2. It is used as an electrode in electrolysis
3. It is used in making the "lead" for pencils. This is because it leaves a grey streak when it is drawn across a sheet of paper

[B] Diamond

Structure of diamond

Diamond is a colourless, crystalline solid with an extremely high density

It is a giant structure of carbon atoms



Each carbon atom shares electrons with each of its four nearest neighbours, thus forming four covalent bonds

In addition, each carbon atom is imagined to be at the centre of the tetrahedron surrounded by four other carbon atoms whose centres are at the corners of the tetrahedron

Physical properties of diamond

1. It is very hard - the hardest substance known. It has a very high melting point of about 3700°C because each atom is held in place by four strong bonds.
2. It is colourless and transparent with a dazzling brilliant lustre
3. It has a density of $3.5g/cm^3$
4. It does not conduct electricity because there are no ions or free electrons in it to carry charge

Uses of diamond

1. It is uesd for cutting tools and drilling devices
2. It is used for cutting glass
3.

[C] Amorphous carbon

Amorphous carbon such as coal and charcoal is porous and easily absorbs pigments from solutions e.g in the refining of white spoon sugar

Charcoal is used to absorb the brown colour from brown sugar which is then turned white

Chemical properties of carbon

1. All the forms of carbon react with oxygen to form carbon dioxide

   $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

Example

1. Diamond and graphite are allotropes of carbon. Diamond has a melting point of about 3700°C where as graphite has a melting point of about 3300°C.

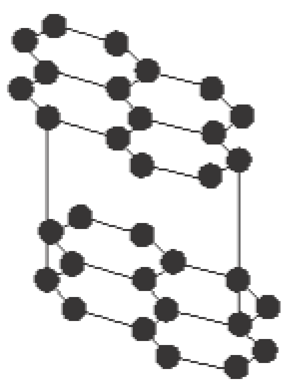
(a) Explain what is meant by the term allotropes?
(b) Explain why diamond does not conduct elecrticity where as graphite does

Solution

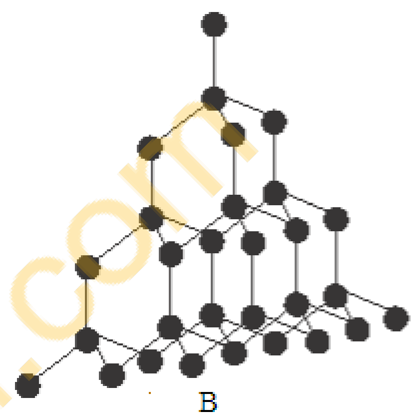
1. (a) Allotopes are elements in different physical forms but in the same state.
   (b) Diamond does not conduct electricity because there are no ions or free electrons in it to carry charge where as graphite conduct conduct electricity because of free moving electrons in between the layers of carbon atoms.

Exercise

1. Carbon exists in two different forms. They have the following sructures



A



B

(a) Name the form of carbon having:
   (I) Structure A
   (II) Structure B
(b) Give the use of the form of carbon having:
   (I) Structure A
   (II) Structure B
(c) Explain why both diamond and graphite have very high melting points
(d) Why is graphite used in pencils?

Carbonates, oxides and hydroxides

[A] Calcium carbonate

Chemical formula: $CaCO_3$

Special name: Limestone / marble

Calcium carbonate is a white solid which is insoluble in water

When strongly heated, calcium carbonate decomposes to form calcium oxide and carbon dioxide

$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$

Uses of calcium carbonate

1. It is used in the manufacture of cement
2. It is used for making glass
3. It is used to remove impurities like silica as slag in the blast furnace

[B] Calcium oxide

Chemical formula: CaO

Special name: Quick lime

Calcium oxide reacts vigorously with water to form calcium hydroxide and a lot of heat energy

$CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} +$ Heat energy

[C] Calcium hydroxide

Chemical formula: $Ca(HO)_2$

Special name: Slaked lime or lime water

Calcium hydroxide glows brightly at high temperatures

Calcium hydroxide turns milky or cloudy when carbon dioxide is passed through it and a white precipitate (suspension) of calcium carbonate is formed

$Ca(OH)_{2(aq)} + CO_{2(g)} \rightarrow CaCO_{3(s)} + H_2O_{(l)}$

When excess carbon dioxide gas is passed through for some time, the precipitate disappears and a clear solution of calcium hydrogen carbonate is formed

$CaCO_{3(s)} + H_2O_{(l)} + CO_{2(g)} \rightarrow Ca(HCO_3)_{2(aq)}$

Uses of calcium oxide and calcium hydroxide

1. They are used in treating acidic soils on the farm
2. They are used in the neutralization of acidic industrial waste products before discharging them into rivers and lakes
3. Calcium hydroxide is used as plaster of paris for broken arms and legs
4. Calcium oxide is used as lining in the blast furnace.
5. Calcium oxide is used to remove silica impurities in the extraction of iron.
6. Calcium oxide is used as a drying agent especially for ammonia gas.

Example

1. A white solid V is heated strongly. At a high temperature, V gives off carbon dioxide and changes to solid W which begins to glow. When water is added to W, it begins to fizz, produces heat and crumbles to a powder X. X dissolves in water to form an alkaline solution Y. When carbon dioxide is passed into solution Y, the solution turns cloudy. The cloudiness is caused by a white suspension Z and its formation is used as a test for carbon dioxide.
   (a) Deduce the identity of V, W, X, Y and Z giving either their chemical or common names.

(b) Write chemical equations for:
   (I) the action of heat on V
   (II) the action of carbon dioxide on solution Y
(c) State the use for either W or X

Solution

(a) V - Calcium carbonate, $CaCO_3$
   W – Calcium oxide, CaO
   X – Calcium hydroxide, $Ca(OH)_2$
   Y – Calcium carbonate, $CaCO_3$
   Z – Calcium hydrogen carbonate, $Ca(HCO_3)_2$
(b) (I) $CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$
   (II) $CaCO_{3(s)} + H_2O_{(l)} + CO_{2(g)} \rightarrow Ca(HCO_3)_{2(aq)}$
(c) W and X are used in treatment of acidic soils on the farms and also in the neutralization of acidic waste products before discharging them into rivers and lakes
   X is used as plaster of Paris for broken arms or legs.

Exercise

1. When carbon dioxide gas is bubbled through lime water for a short period, a white precipitate **G** is formed. When **excess** carbon dioxide is passed through, **G** dissolves and a colourless solution containing the soluble compound **H** is produced.
   (a) Identify
      (I) Solid **G**
      (II) Compound **H**
   (b) Write down the chemical equation including state symbols to show the formation of the white precipitate **G**.

Organic chemistry

Definition: Organic chemistry is the study of carbon compounds except carbonates, carbon monoxide, carbon dioxide, carbon disulphide etc.

Common terms used

1. Organic compounds

Organic compounds are compounds which contain carbon except carbonates, carbon monoxide, carbon dioxide, carbon disulphide etc.

2. Hydrocarbon

A hydrocarbon is a compound which consists of the elements carbon and hydrogen only,

Example

Hydrocarbon	**Formula**
Methane	CH_4
Ethene	C_2H_4
Ethyne	C_2H_2

3. Homologous series

A homologous series is a family of similar organic compounds

Each member in a family is called homologue

Characteristics of the homologous series

(a) Members in the family show a gradual change in physical properties i.e. change from gas to liquid and solid as the carbon chain increases.

(b) Members in the family have similar methods of preparation

(c) Members in the family have similar chemical properties because they have the same functional group.

(d) Members in the family can be represented by the general molecular formula:

Example

Organic family	**General molecular formula**
Alkanes	C_nH_{2n+2}
Alkenes	C_nH_{2n}
Alkynes	C_nH_{2n-2}
Alcohols (Alkanols)	$C_nH_{2n+1}OH$
Carboxylic acids (Alkanoics)	$C_nH_{2n+1}COOH$
Esters	$C_nH_{2n+1}COOC_mH_{2m+1}$

(e) Each member in the family differs to the next by CH_2 or by 14.

4. Functional group

A functional group is a group is a group which determines the chemical properties of organic compounds

Organic family	**Functional group**	**Name of functional group**
Alkanes	$-C-C-$	Single carbon – carbon covalent bond
Alkenes	$C=C$	Double carbon – carbon covalent bond
Alkynes	$C\equiv C$	Triple carbon – carbon covalent bond
Alkanols (alcohols)	$-OH$	Hydroxyl group

Alkanoics (carboxylic acids)	$-COOH$ or $\overset{\displaystyle O}{\overset{\|}{C}}-O-H$	Carboxyl group
Esters	$-COO-$	Ester functional group

Nomenclature of organic compounds
Nomenclature is a standardized way of naming organic compounds.
Nomenclature takes into account the following:

- Prefix
- Suffix

The prefix shows the number of carbon atoms in an organic compound.

Prefix	Number of carbon atoms
Meth	1
Eth	2
Prop	3
But	4
Pent	5
Hex	6
Hept	7
Oct	8
Non	9
Dec	10

The suffix shows the organic family of the organic compound.

Suffix	Organic family
ane	Alkanes
ene	Alkenes
yne	Alkynes
anol	Alcohols
anoic	Carboxylic acids

anoate	Esters

Alkanes

Alternative term: Saturated hydrocarbons

General molecular formula: C_nH_{2n+2} where n = 1, 2, 3, 4.......

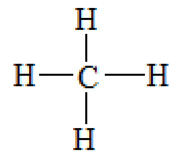
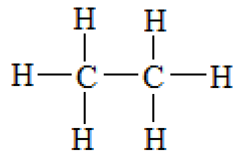
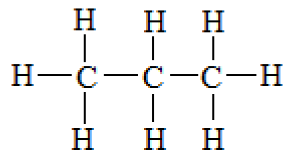
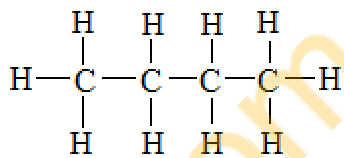
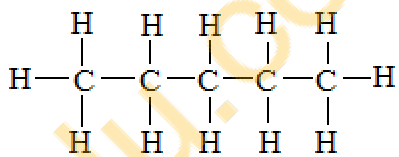
Alkanes have single carbon – carbon covalent bonds between carbon atoms.

They are called saturated hydrocarbons because they have the required (maximum) number of hydrogen atoms.

They end with ane

Examples of alkanes

Alkane	Molecular formula	Structural formula	Condensed formula	M_r

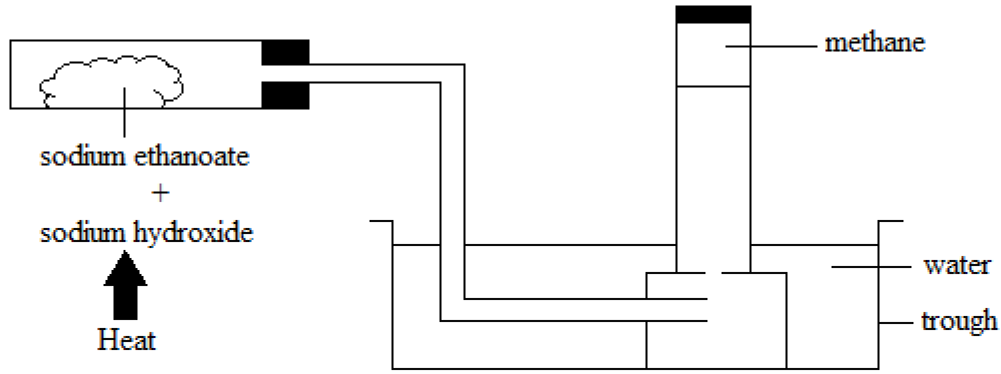
Methane	CH_4		CH_4	16
Ethane	C_2H_6		CH_3CH_3	30
Propane	C_3H_8		$CH_3\,CH_2CH_3$	44
Butane	C_4H_{10}		$CH_3\,(CH_2)_2CH_3$	58
Pentane	C_5H_{12}		$CH_3\,(CH_2)_3CH_3$	72

Preparation of methane

Methane can be prepared by heating a mixture of anhydrous sodium ethanoate, CH_3COONa, and soda lime (sodium hydroxide), NaOH.

Reaction: $CH_3COONa_{(s)} + NaOH_{(s)} \rightarrow CH_{4(g)} + Na_2CO_{3(s)}$

Diagram:



Physical properties of alkanes

1. Their physical states vary at room temperature and pressure. Lower members are gases; others are liquids while higher members are solids.
2. Their melting and boiling points increases as the carbon chain (relative molecular masses) increases.
3. Their densities increases as the carbon chain increases.

Chemical properties of alkanes

1. **Combustion**

(a) In a plentiful supply of air (complete combustion), alkanes reacts with oxygen to form carbon dioxide and water.

Example

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

(b) In a limited supply of air (incomplete combustion), alkanes react with oxygen to form carbon monoxide and water

Example

$2C_2H_6 + 5O_2 \rightarrow 4CO + 6H_2O$

(c) In a very limited supply of air (incomplete combustion), alkanes react with oxygen to form carbon and water

Example

$C_3H_8 + 2O_2 \rightarrow 3C + 4H_2O$

2. **Substitution reaction**

A substitution reaction is a reaction in which one atom or group of atoms in a molecule is replaced by another.

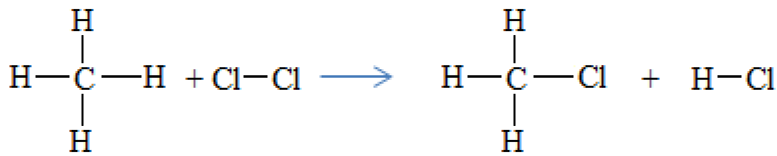
Alkanes undergo substitution reactions with halogens in the presence of ultraviolet light.

Example

Halogenation of methane

Methane can react with chlorine gas in the presence of sunlight to form chloromethane and hydrogen chloride.

$CH_{4(g)}$ + $Cl_{2(g)} \rightarrow CH_3Cl_{(g)}$ + $HCl_{(g)}$



Conditions

1. Temperature: 200^{o}C
2. Catalyst: Sunlight (ultraviolet light)

Isomerism of alkanes

Isomerism is the existence of compounds with the same same molecular formula but different structural formula

Isomers

Isomers are compounds which have the same molecula formula but different structural formula.

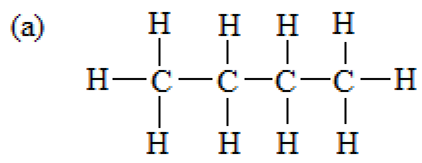
Alkanes with more than three carbon atoms exhibit isomerism

Nomenclature of alkane isomers

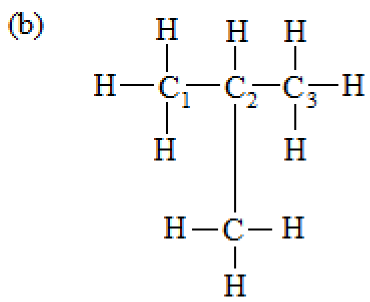
The International Union for Pure and Applied Chemistry (IUPAC) rules for naming alkanes

- Choose the longest continous carbon chain in the molecule.
- Number the carbon atoms in the longest chain chosen starting from the end that gives the smaller set of numbers nearer the branched methly group (CH_3)
- Write the number of carbon atom at which the branched methyl group is attached to the main chain in alphabetical order.
- Put a comma after each group number and a hyphen between the group number and the methyl group.
- Name the compound as a derivative of the normal alkane having the same number of carbon atoms as in the longest chain chosen.

Example 1: Butane has two isomers

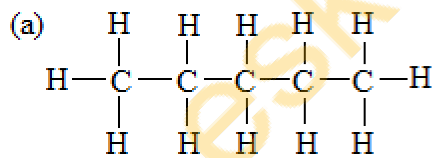


Molecular formula: C_4H_{10}
IUPAC name: Normal butane

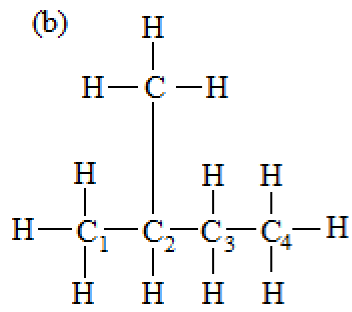


Molecular formula: C_4H_{10}
IUPAC name: 2 – methyl propane

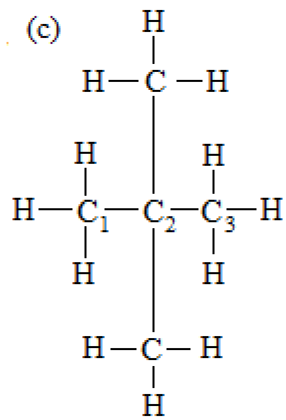
Example 2: Pentane has three isomers



Molecular formula: C_5H_{12}
IUPAC name: Normal pentane



Molecular formula: C_5H_{12}
IUPAC name: 2 – methyl butane



Molecular formula: C_5H_{12}
IUPAC name: 2, 2 – dimethyl propane

Exercise

1. Draw the structure of the compound with the IUPAC name; 2, 3, 4, 4 – tetra methylpentane

Cracking of alkanes

Cracking is the breaking down of large hydrocarbon molecules into simpler and smaller molecules.

Alkanes break down to produce alkenes, short chain alkanes and in some cases hydrogen gas.

Types of cracking

1. **Thermal cracking**
   This is where heat is used to break down large molecules into smaller ones.

Example

$C_4H_{10} \rightarrow C_2H_6 + C_2H_4$

$C_2H_6 \rightarrow C_2H_4 + H_2$

2. **Catalytic cracking**
   This is where both heat and a catalyst are usedto break down large molecules into smaller ones.
   Common catalysts used include aluminium oxide (Al_2O_3) and silica (SiO_2).

Example

$C_{17}H_{36(l)} \rightarrow 3C_2H_{4(g)} + C_8H_{18} + C_3H_{6(g)}$

Conditions

1. Catalyst: Aluminium oxide, Al_2O_3
2. Temperature: 600^oC

Importance of cracking

1. It is used in the production of raw materials for petrol chemicals like ethene and hydrogen. Hydrogen is used in making plasitics and margarine.
2. It is used in the production of petrol.

Exercise

1. One of the hydrocarbons found in crude oil is undecane, $C_{11}H_{24}$
   Under suitable conditions, undecane under goes the reaction below
   $C_{11}H_{24} \rightarrow C_7H_{16} + C_4H_8$
   (a) What term is used to describe this type of reaction in the petrochemical industry?
   (b) State two conditions necessary for this reaction to take place

Fuels

A fuel is a substance which burns in air to produce useful energy.

Sources of fuels

(a) **Natural gas**
   The main component of natural gas is methane.
   Natural gas is usually mined and purified before use.

(b) **Crude oil**

Alternative term: Petroleum

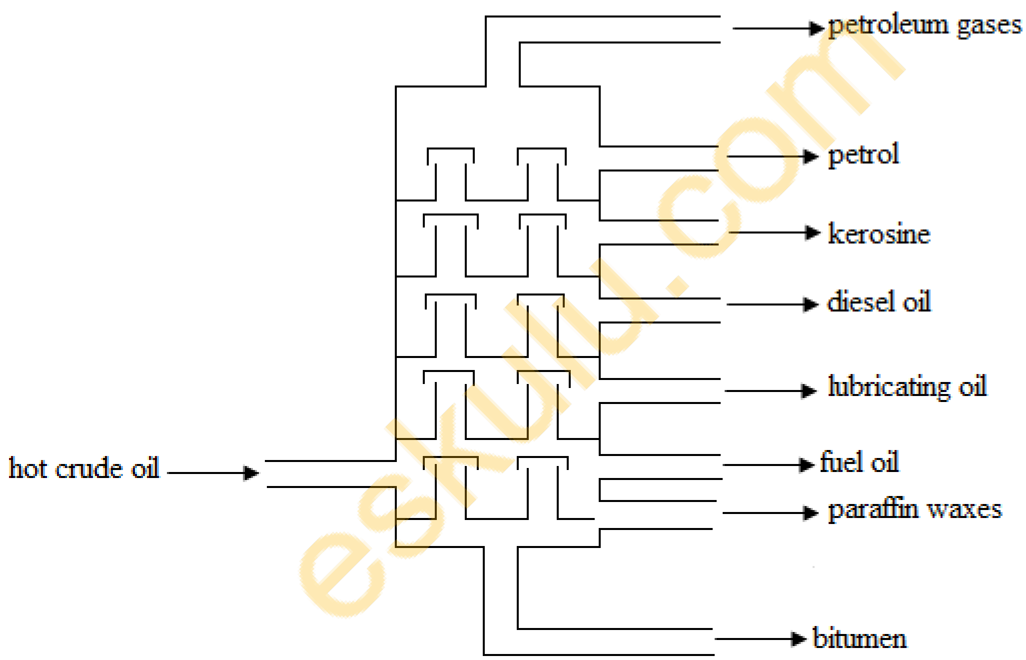
Crude oil is a mixture of long chain hydrocarbons and some elements like sulphur and nitrogen

Crude oil was formed millions of years ago by the decomposition of animals and plants under pressure.

Crude oil can be refined by fractional distillation.

Fractional distillation is a process of separating components of a mixture based on their boiling points.

The diagram below shows a fractionating tower and different fractions obtained from crude oil.



Fractions of crude oil and their uses

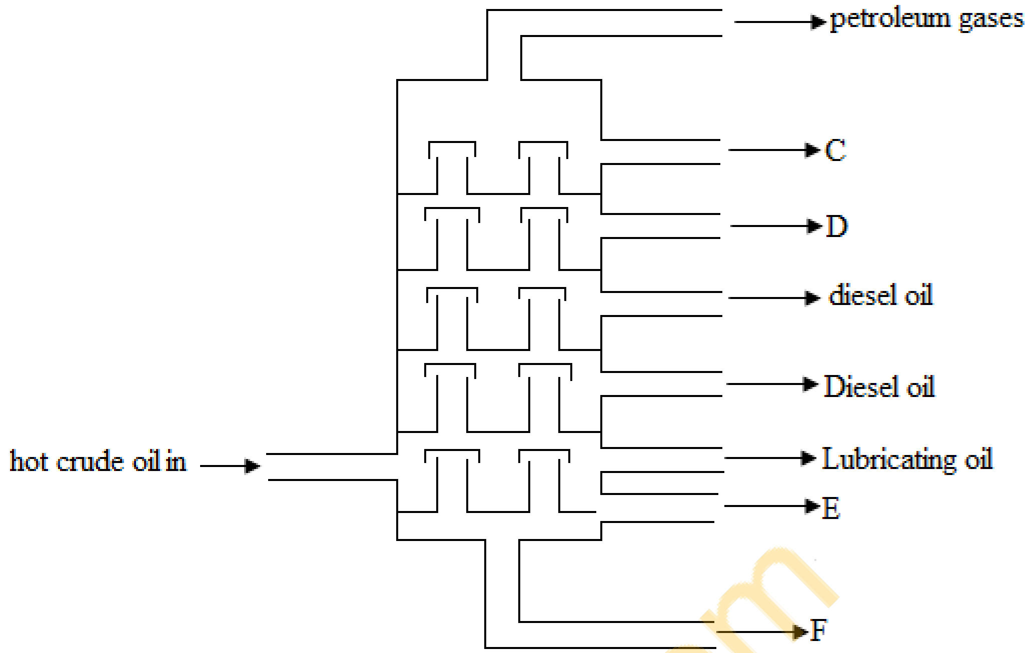
Fraction	Number of carbon atoms	Use
Petroleum gases	$C_1 - C_4$	Used in camping stoves and cookers as bottled gas
Petrol (gasoline)	$C_4 - C_{10}$	Used as a fuel in internal combustion engines in vehicles
Kerosine (paraffin)	$C_{10} - C_{17}$	Used as a fuel in jets Used in paraffin lamps for lighting Used in paraffin lamps for cooking
Diesel oil (gas oil)	$C_{17} - C_{20}$	Used as a fuel in diesel engines like trains, hammer mills, tractors, lorries
Lubricating oil	$C_{20} - C_{30}$	Used as a lubricant in machines to reduce friction
Fuel oil	$C_{30} - C_{40}$	Used as a fuel in ships, power stations and industrial machinary
Paraffin waxes	$C_{40} - C_{50}$	Used in making waxes, polishes and wax papers
Bitumen (residue)	C_{50} upwards	Used to tar roads Used as a roofing material

Note

1. Petroleum gases have the smallest boiling point and have the smallest relative molecular mass.
2. Bitumen has the highest boiling point and have the largest relative molecular mass. In addition, bitumen has the longest chain.
3. Octane is the main component of petrol

Exercise

1. Crude oil, a mixture of hydrocarbons, can be separated by fractional distillation. The diagram below shows a fractionating tower and different fractions obtained from crude oil



(a) What is meant by the term hydrocarbon?
(b) Which fraction contains the longest chain?
(c) State one major use of fraction:
    (I) C
    (II) D
    (III) E
(d) State the process involved when ethene is manufactured from one of the fractions obtained by the fractional distillation of crude oil.

Production of hydrogen from methane

Methane reacts with steam to form carbon monoxide and hydrogen gas.

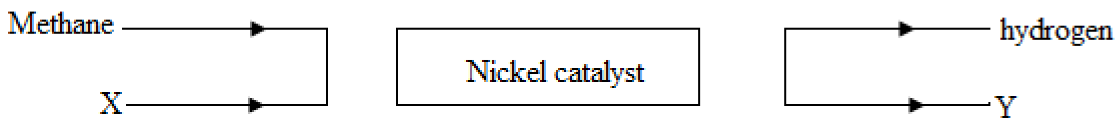
$$CH_{4(g)} + H_2O_{(g)} \rightarrow CO_{(g)} + 2H_{2(g)}$$

Conditions

1. Temperature: 200^oC
2. Catalyst: Nickel metal

Exercise

1. Hydrogen can be manufactured from methane using a nickel catalyst as shown in the diagram below



(a) Identify the substances X and Y
(b) Write a balanced chemical equation for the reaction in the manufacure of hydrogen
(c) State one chemical property of hydrogen
(d) State one industrial use of hydrogen

Alkenes

Alternative term: Unsaturated hydrocarbons

General molecular formula: C_nH_{2n} where n = 2, 3, 4, 5…..

Functional group: Double carbon – carbon covalent bond, C = C

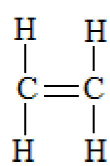
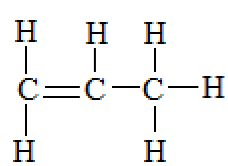
They have one or more double bonds between carbon atoms.

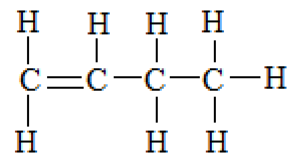
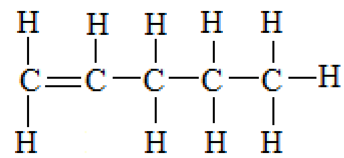
They end with ene.

They are called unsaturated hydrocarbons because they have a double bond between carbon atoms.

They have two hydrogen atoms less than their corresponding alkanes

Examples of alkenes

Alkene	Molecular formula	Structural formula	Condensed formula	M_r
Ethene	C_2H_4		$CH_2{=}CH_2$	28
Propene	C_3H_6		$CH_2{=}CHCH_3$	42

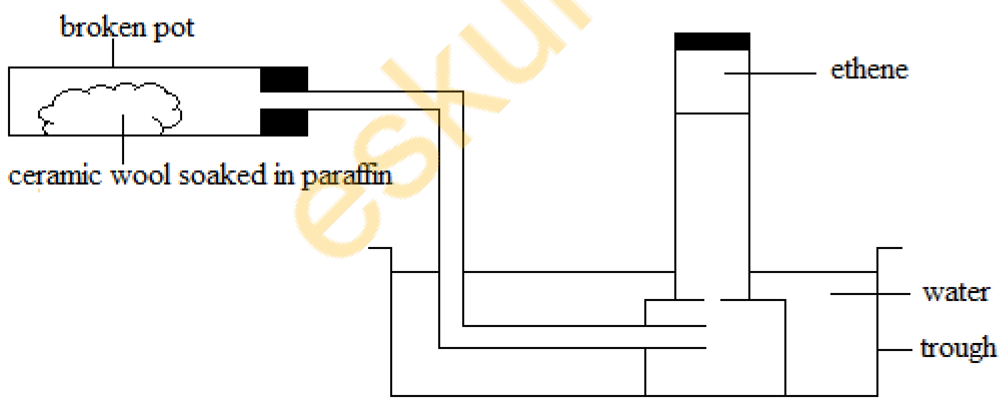
Butane	C_4H_8		$CH_2=CHCH_2CH_3$	56
Pentene	C_5H_{10}		$CH_2=CH(CH_2)_2CH_3$	70

Preparation of ethene gas

There are two methods used to prepare ethene.

1. **Catalytic cracking of alkanes**

This is where a catalyst is used to break down large hydrocarbon molecules into smaller ones.



Reaction: $C_{17}H_{36(l)} \rightarrow 3C_3H_{4(g)} + C_3H_{6(g)} + C_8H_{18(l)}$

Conditions

1. Temperature: 600^oC
2. Catalyst: Aluminium oxide, Al_2O_3

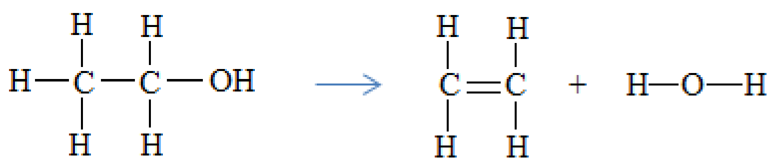
Note: The broken pot can act as a catalyst.

2. **Dehydration of ethanol**

Ethene can be prepared by the dehydration of ethanol using concentrated sulphuric acid as a dehydrating agent.

Sulphuric acid removes water.

$C_2H_5OH_{(s)} \rightarrow C_2H_{4(g)} + H_2O_{(l)}$



Conditions

1. Temperature: 180^{o}C
2. Dehydrating agent: Concentrated sulphuric acid.

Test for unsaturation

Alternative term: Test for alkenes

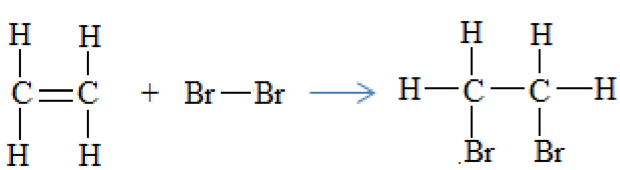
There are two ways you could test a hydrocarbon, to see whether it is an alkane or alkene

Test 1: Shake an alkene with bromine water.

Result: The brown colour of bromine disappears immediately. In other words alkenes decolorize bromine water.

Explanation: When ethene reacts with bromine water, the colourless compound (liquid) called 1, 2 – dibromoethane is formed.

$C_2H_{4(g)} + Br_{2(l)} \rightarrow C_2H_4Br_{2\,(l)}$



Note

When an alkane is shaken with bromine water, there is no reaction. Alkanes do not undergo addition reactions because they are saturated.

Test 2: Add acidified potassium per manganate (VII) solution to the hydrocarbon.

Result: The purple colour of potassium per manganate (VII) turns colourless if an alkene is present.

Chemical properties of alkenes e.g. ethene

Alkenes are more reactive than alkanes because the double bond opens up to allow chemical reactions.

Alkenes undergo two notable chemical reactions.

1. **Combustion**

Alkenes react with oxygen to form carbon dioxide and water.

Example

$$C_2H_{4(g)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 2H_2O_{(l)}$$

2. **Addition reactions**

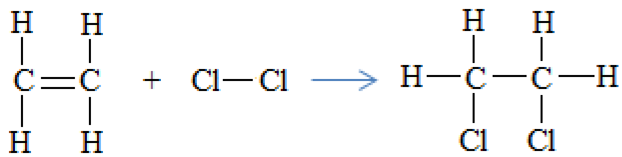
An addition reaction is a reaction where a molecule is added to an unsaturated molecule by breaking a double bond.

Examples of addition reaction involving ethene

(a) Halogenation – reaction with halogens

Ethene reacts with chlorine to form 1, 2 – dichloroethane.

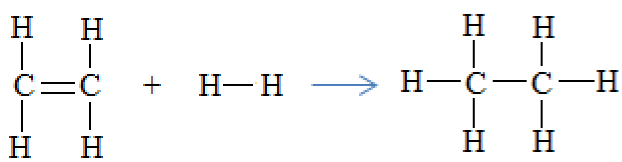
$$C_2H_{4(g)} + Cl_{2(g)} \rightarrow C_2H_4Cl_{2\,(l)}$$



(b) Hydrogenation – reaction with hydrogen

Alkenes react with hydrogen to form corresponding alkanes. Hydrogen makes the unsaturated compounds to become saturated. Ethene reacts with hydrogen to form ethane

$C_2H_{4(g)} + H_{2(g)} \rightarrow C_2H_{6\,(g)}$



Conditions

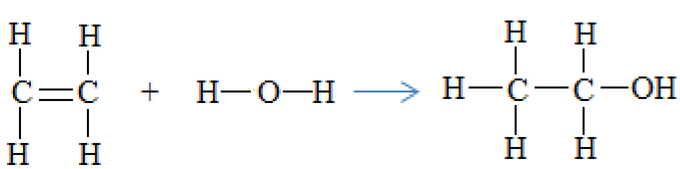
1. Temperature: 200^oC
2. Catalyst: Nickel metal

The reaction is important in the industry because it is used in the production of margarine from unsaturated vegetable oil. Vegetable oils are changed into fats which are saturated by hydrogenation.

(c) Hydration – reaction with water

Ethene reacts with steam to in the presence of phosphoric acid to form ethanol.

$C_2H_{4(g)} + H_2O_{\,(g)} \rightarrow C_2H_5OH_{(l)}$



Conditions

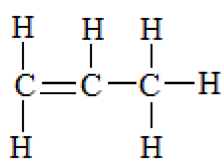
1. Temperature: 300^oC
2. Catalyst: Phosphoric acid, H_3PO_4
3. Pressure: 65 atmospheres

Uses of ethene

1. Used in the manufacture of plastics
2. Used in the manufacture of ethanol
3. Used in the process of ripening fruits

Exercise

1. Ethene, C_2H_4, is an unsaturated compound. It reacts with liquid bromine to form 1, 2 – dibromoethane.
   Ethane,C_2H_6, is a saturated compound. It does not readily react with bromine
   (a) Draw the full structural formula for ethene and 1, 2 – dibromoethane
   (b) Write the equation for the reaction between ethene and bromine and name the type of reaction taking place
   (c) Ethene like many other unsaturated compounds can be polymerized. Polyethene is formed when ethene is polymerized.



   Name the compound which has the above stucture

   (d) Margarine is manufactured using the addition reaction between hydrogen and a vegetable oil
      (I) State the conditions used for this reaction
      (II) What type of bond must be present in the vegetable oil for this reaction to take place.

Alcohols

Alternative term: Alkanols

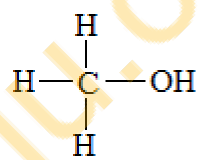
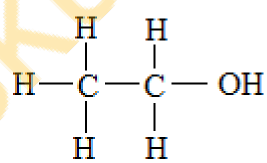
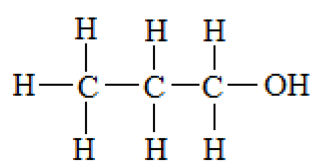
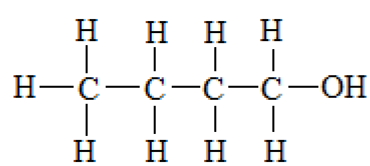
General molecular formula: $C_nH_{2n+1}OH$ where n = 1, 2, 3, 4…

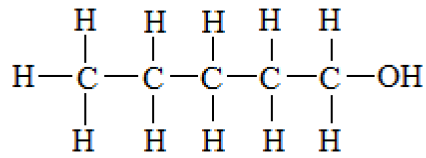
Functional group: hydroxyl group, $-$OH

They end with anol

They are not hydrocarbons because they contain oxygen

Examples of alcohols

Alcohol	**Molecular formula**	**Structural formula**	**Condensed formula**	**M_r**
Methanol	CH_3OH		CH_3OH	32
Ethanol	C_2H_5OH		CH_3CH_2OH	46
Propanol	C_3H_7OH		$CH_3(CH_2)_2OH$	60
Butanol	C_4H_9OH		$CH_3(CH_2)_3OH$	74

Pentanol	$C_5H_{11}OH$		$CH_3(CH_2)_4OH$	88

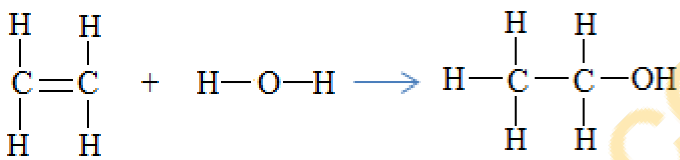
Preparation of ethanol

There are two methods of preparing ethanol.

(a) Hydration of ethene

Ethanol can be prepared when ethene reacts with steam

$$C_2H_{4(g)} + H_2O_{(g)} \rightarrow C_2H_5OH_{(l)}$$



Ethene Steam Ethanol

Conditions

1. Temperature: 300^oC
2. Catalyst: Phosphoric acid, H_3PO_4
3. Pressure: 65 atmospheres

(b) Fermentation of sugars

Fermentation is the decomposition of sugars using enzymes in yeast to produce ethanol and carbon dioxide. Sugars (glucose) is mixed with water and yeast and allowed to react for a few days in the absence of air.

Glucose → Ethanol + Carbon dioxide

$$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$$

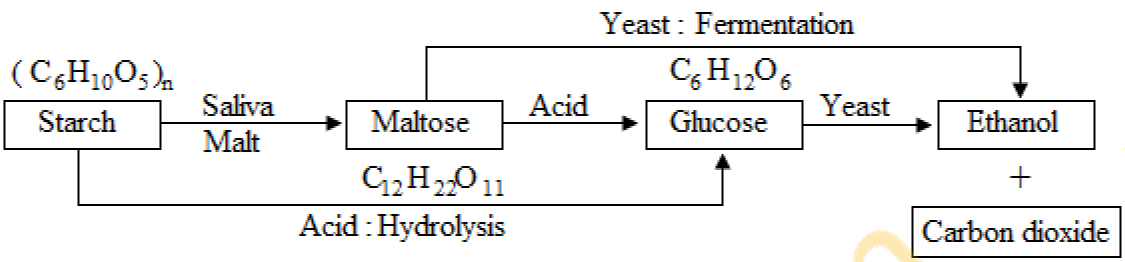
Starch hydrolysis

Hydrolysis is the chemical reaction of a compound with water which causes it to break down.

Starch is broken down to form **sugar** by:

1. **Acid hydrolysis** (heated with dilute acid)
   Acid hydrolysis is slow but eventually the starch is broken down into glucose, which is the monomer and will not undergo further hydrolysis
2. **Enzyme hydrolysis** (by enzyme amylase)
   The enzyme breaks down the starch into the disaccharide maltose which contains two glucose units minus a water molecule.

Summary of the breakdown of starch to maltose and glucose and then to ethanol



Conditions for fermentation

1. Optimum temperature: 37°C
2. Catalyst (enzyme): Zymase

Note

Fermentation should take place in the absence of air (oxygen) to prevent oxidation of ethanol to ethanoic acid.

If the temperature goes above 40°C , the enzymes in yeast which catalyze the reaction becomes denatured.

Physical properties of alcohols

1. They are colourless inflammable liquids
2. Their boiling points increases as the carbon chain increases
3. Their solubilities in water decreases as the number of carbon chain increases

Chemical properties of alcohols

1. Combustion

Alcohols burn in air (react with oxygen) to form carbon dioxide and water.

Example

Ethanol + oxygen → carbon dioxide + water

$$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$$

The reaction is exothermic. It gives out a lot of heat energy

2. Oxidation

Oxidation is the addition of oxygen to a substance

(a) Ethanol can be oxidized to ethanoic acid by bacteria in the air
$C_2H_5OH + O_2 \rightarrow CH_3COOH + H_2O$

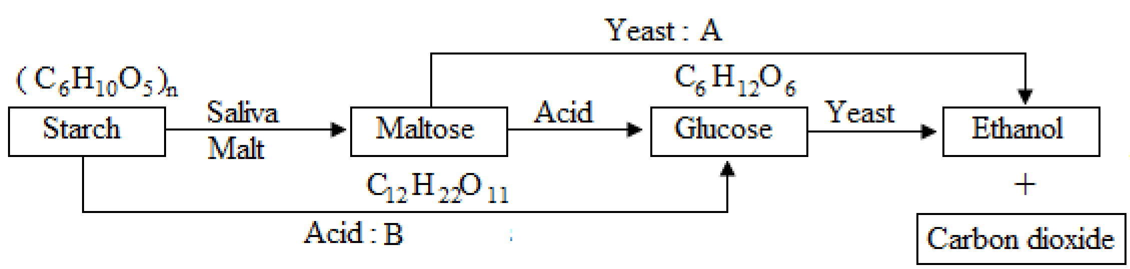
(b) Ethanoic acid can also be formed by using an oxidizing agent such as acidified potassium per manganate $[KMnO_4]$ and potassium dichromate (VI) $[K_2Cr_2O_7]$
$C_2H_5OH + 2[O] \rightarrow CH_3COOH + H_2O$
The oxygen is from the oxidizing agent.
When ethanol reduces potassium permanganate, the reaction is indicated by the colour change from purple to colourless on mixing.
When ethanol reduces potassium dichromate, the reaction is indicated by the colour change from orange to green on mixing.

Uses of ethanol

1. Used as a component in beer and wines
2. Used as a solvent
3. Used in making methylated spirit
4. Used as a fuel
5. Used in the preservation and sterilization of food

Example

1. The diagram below gives a summary of the breakdown of starch to maltose and glucose



(a) Name the processes represented by the letters A and B
(b) What is the purpose of the yeast in process A?

(c) Apart from water, which other product is produced when a mixture of propanol and concentrated potassium dichromate is heated?

(d) If ethanol is added to the product in part (c) in the presence of concentrated sulphuric acid,

(I) Name the compounds to which the products belongs

(II) Give two everyday uses of the product in (I) above.

Solution

(a) A = Fermentation

B = Hydrolysis

(b) To speed up the reaction (it acts as a catalyst)

(c) Propanoic acid

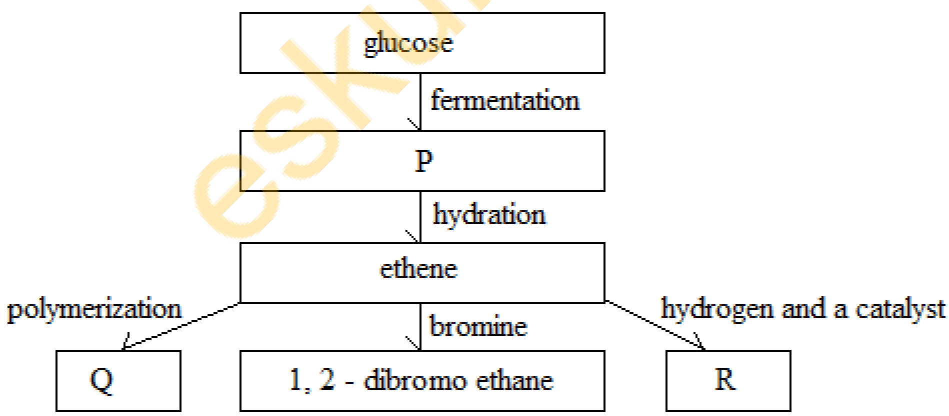
(d) (I) Esters

(II) They are used in perfumes due to sweet fruit smells

They are used in food and drink flavouring and preservation

Exercise

1. Carefully study the reaction below and use it to answer the questions that follow



(a) Give the names of the substances and the conditions needed to ferment glucose

(b) Give the name of the gas that is also produced during fermentation of glucose. Describe the simple test for this gas

(c) Give the name and molecular formula for substances P

(d) Give the names of substance Q and R

(e) Write an equation for the addition reaction between ethene and bromine.

Carboxylic acids

Alternative term: Alkanoics

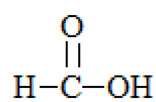
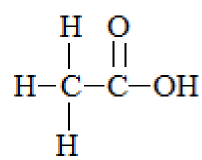
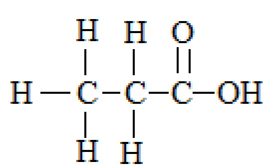
General molecular formula: $C_nH_{2n+1}COOH$ where n = 0, 1, 2, 3….

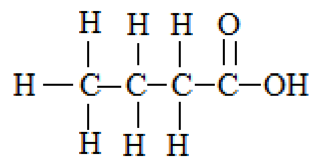
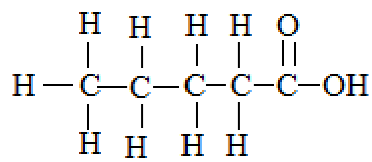
Functional group: Carboxyl group, ─ COOH.

They end with anoic acid

They are not hydrocarbons because they contain oxygen

Examples of carboxylic acids

Carboxylic acid	Molecular formula	Structural formula	Condensed formula
Methanoic acid	HCOOH		HCOOH
Ethanoic acid	CH_3COOH		CH_3COOH
Propanoic acid	C_2H_5COOH		CH_3CH_2COOH

Butanoic acid	C_3H_7COOH		$CH_3(CH_2)_2COOH$
Pentanoic acid	C_4H_9COOH		$CH_3\,(CH_2)_3COOH$

Preparation of ethanoic acid

(a) Oxidation of ethanol

(I) Ethanoic acid can be prepared by the oxidation of ethanol by bacteria in atmospheric air.

$$C_2H_5OH + O_2 \rightarrow CH_3COOH + H_2O$$

(II) Ethanoic acid can also be prepared by the oxidation of ethanol using an oxidizing agent e.g. Acidified potassium dichromate(VI)

$$C_2H_5OH + 2[O] \rightarrow CH_3COOH + H_2O$$

The oxygen is from the oxidizing agent.

The orange acidified potassium dichromate (VI) solution turns green in this reaction.

Physical properties of carboxylic acids

1. They turn blue litmus paper red
2. They have PH values less than 7
3. They have a sour taste

Chemical properties of carboxylic acids

1. They react with reactive metals to form a salt and hydrogen gas

Example

$$2Na + 2CH_3COOH \rightarrow 2CH_3COONa + H_2$$

2. They react with alkalis and bases to form a salt and water only

Example

$NaOH + CH_3COOH \rightarrow CH_3COONa + H_2O$
$CaO + 2CH_3COOH \rightarrow (CH_3COO)_2Ca + H_2O$

3. They react with carbonates and hydrogen carbonates to form a salt, water and carbon dioxide

Example

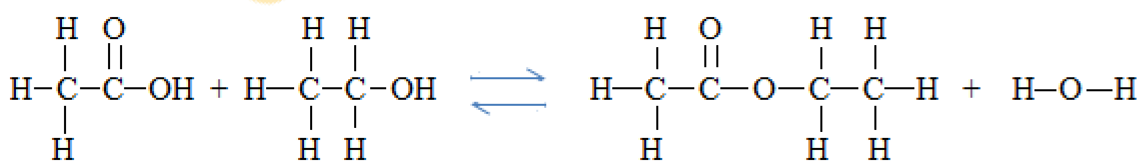
$NaHCO_3 + CH_3COOH \rightarrow CH_3COONa + H_2O + CO_2$
$CaCO_3 + 2CH_3COOH \rightarrow (CH_3COO)_2Ca + H_2O + CO_2$

4. They react with alcohols to form esters and the process is called esterification. Esterification is a reaction between a carboxylic acid and an alcohol to form an ester in the presence of sulphuric acid. Esters are sweet smelling compounds.

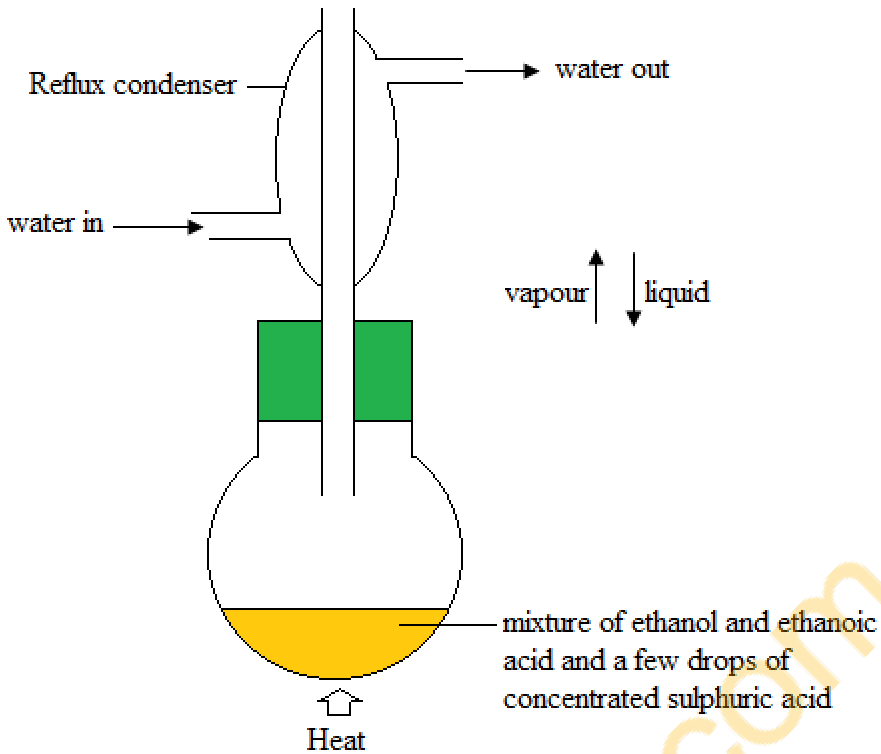
Example

Ethanoic acid can react with ethanol to form an ester called ethyl ethanoate and water. In this reaction, ethanoic acid loses the – OH group while ethanol loses the – H group to form water. The remaining sections of the molecules join together to form the ester.

$$CH_3COOH + C_2H_5OH \rightleftharpoons CH_3COOC_2H_5 + H_2O$$



Ethanoic acid ethanol ethyl ethanoate water



Conditions for esterification

1. Temperature: 180^oC
2. Catalyst: Sulphuric acid

Note

The name of the ester follows the order: alcohol, then acid. For example, if methanol reacts with propanoic acid, the ester formed will be called methyl propanoate.

Reflux condenser: It is held vertically to prevent the escape of any unchanged ethanol. Ethanol has a low boiling point and vaporizes easily. When the ethanol vapour comes into contact with the cold surface of the condenser, it will liquefy and return to the flask.

Esterification is reversible $\rightleftharpoons$

The back ward reaction is called hydrolysis

To prevent hydrolysis, sulphuric acid is added to remove (absorb) water. Water can react with ethyl ethanoate to form ethanoic acid and ethanol.

Special property of esters

1. Esters have sweet smells

Uses of esters

1. They are used in perfumes due to sweet fruit smells
2. They are used in food and drink flavouring and preservation

Similarities between esterification and neutralization

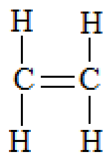
1. Both reactions produce water
2. Both reactions are exothermic

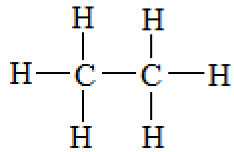
Differences between esterification and neutralization

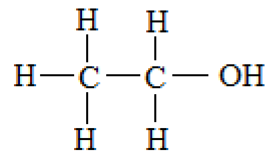
1. Esterification is reversible while neutralization is not reversible
2. Esterification produces an ester while neutralization produces a salt
3. Esterification is slower while neutralization is faster
4. Esterification involves a carboxylic acid (organic acid) and an alcohol while neutralization involves any acid (organic acid or mineral acid) and a base.

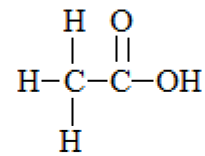
Exercise

1. Substance A is an organic compound which contains an – OH group. On boiling A with acidified potassium dichromate (VII), substance B is formed. An aqueous solution of B has PH 4 and reacts with sodium carbonate to give carbon dioxide.
   (a) Analysis of A gave the following results
   C 52.17%, H 13.04%, O 34.78%
   The relative molecular mass of A is 46. Use the information to identify A and B and draw the full structural formula of A
   (b) Describe what would be seen when the acidified solution of potassium dichromate (VII) was added to A and name the type of reaction in which A is converted to B
   (c) What type of reaction occurs when concentrated sulphuric acid reacts with an alcohol?
   (d) Suggest the full structural formula for the substance produced when concentrated sulphuric acid reacts with the compound of the formula, C_3H_7OH.
2. The structures of four organic compounds belonging to different homologous series are shown below


J


K


L


M

(a) State the name of each of the compounds shown above

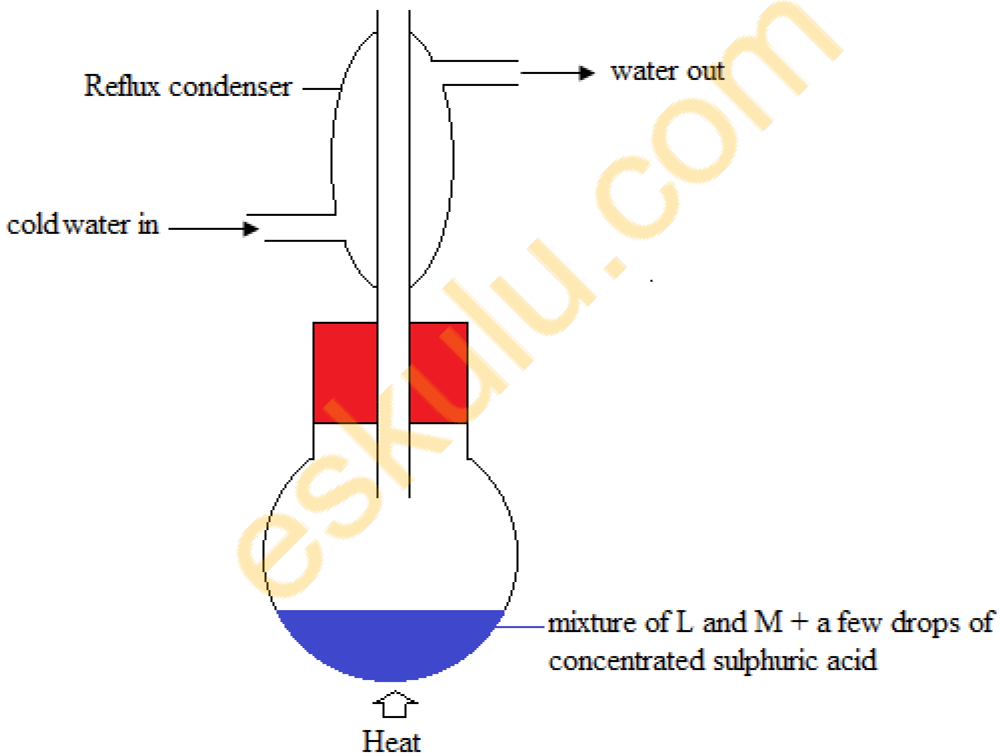
(b) State the type of reaction occurring when;

(I) J is converted to K

(II) J is converted to L

(III) L is converted to M

(c) Compound L reacts with compound M in the presence of a few drops of concentrated sulphuric acid using the apparatus below:



(I) What is the purpose of the concentrated sulphuric acid?

(II) State the name of the organic compound formed in the reaction of L and M.

(III) To which homologous series does the organic compound formed in (II) belong?

Macromolecules

Alternative term: Polymers

Definition: Macromolecules are giant molecules formed by joining smaller units called monomers.

Macromolecules are produced by the process called polymerization.

Polymerization

Definition: Polymerization is the joining up of smaller units called monomers to form larger molecules called polymers.

Types of macromolecules

There are two types of macromolecules; synthetic and natural macromolecules

1. **Synthetic macromolecules**

Alternative term: Synthetic or artificial polymers

Synthetic polymers are man-made structures

They are divided into two categories; addition polymers and condensation polymers

(a) Addition polymers

Addition polymers are polymers formed from smaller identical unsaturated monomers

No other product is formed apart from the polymer

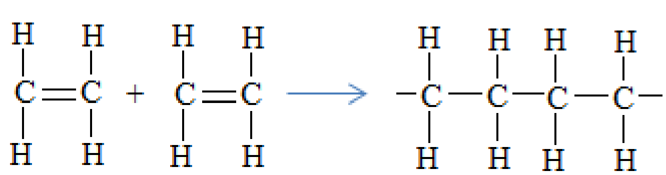
Addition polymers are formed by the process called addition polymerization.

Addition polymerization is polymerization where the polymer has the same empirical formula as the monomer.

Examples of addition polymers

(I) Polyethene

Polyethene is formed when ethene molecules combine



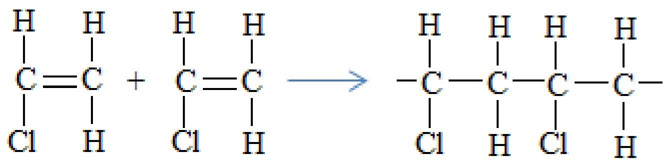
Polyethene

Uses of Polyethene

1. Used in making plastics bags
2. Used in making squeezing bottles

(II) Polyvinyl chloride

Polyvinyl chloride is formed when vinyl chloride molecules combine



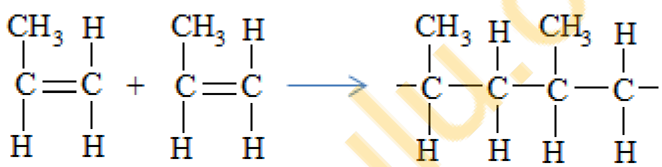
Polyvinyl chloride

Uses of polyvinylchloride

1. Used in making PVC paints
2. Used in making electrical insulators, records, seat covers, rain coats

(III) Polypropene

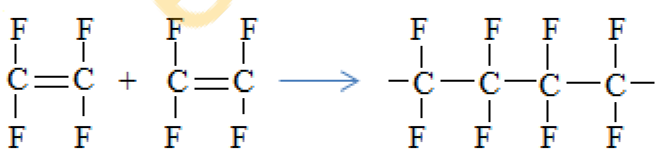
Polypropene is formed when propene molecules combine



Ethane Ethene Polypropene

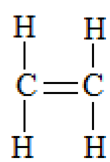
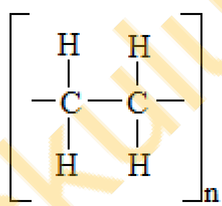
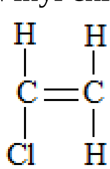
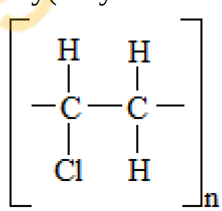
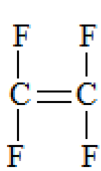
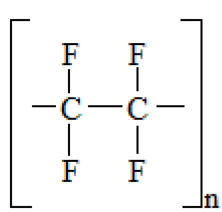
(IV) Polytetrafluoroethene

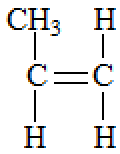
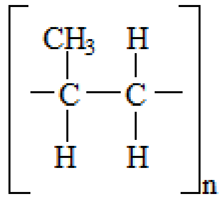
Polytetrafluoroethene is formed when tetrafluoroethene molecules combine



Polytetrafluoroethene

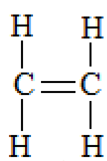
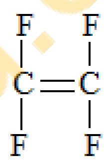
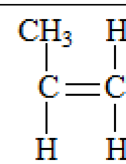
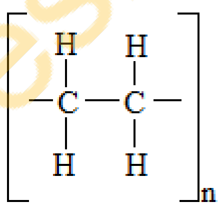
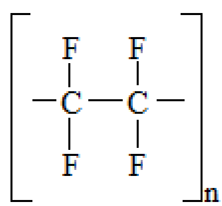
Summary

Monomer	Polymer
Ethene 	Poly(ethene) 
Vinyl chloride 	Poly(vinyl chloride) 
Tetrafluoroethene 	Poly(tetrafluoroethene) 
Propene	Poly(propene)

Exercise

1. The table below shows the structural formulae of various monomers and the polymers that can be made from them.

Monomer	 Ethene	 Tetrafluoroethene	 W
Polymer	 Poly(ethene)	 Poly(tetrafluoroethene)	Y

(I) Identify the monomers which are hydrocarbons.
(II) Name the monomer W.
(III) Draw the displaced structure for polymer Y and name it.
(IV) Identify the common feature among the monomers in the table above.
(V) Give the name of the compounds with the feature in (IV) above.

(b) **Condensation polymers**

Some man-made polymers are formed by condensation polymerization.

Condensation polymerization involves two smaller units which combine to form a larger molecule with the elimination of the water molecule.

Condensation polymers do not have the same empirical formula as the monomers.

Examples of condensation polymers

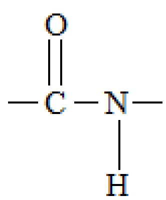
- Nylon
- Terylene

(I) **Nylon**

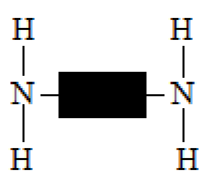
Nylon is a typical polyamide with amide linkages.

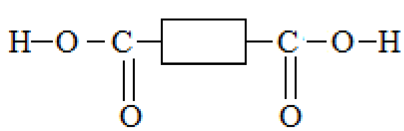
A polyamide is a polymer containing many amide linkages.

The amide linkage



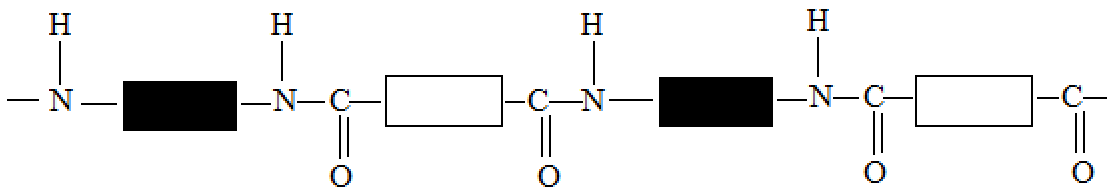
Monomers: Diamine and Dicarboxylic acid

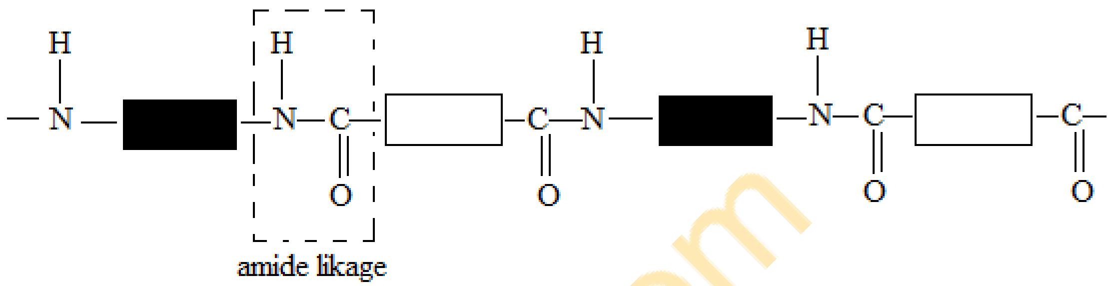
Diamine : 

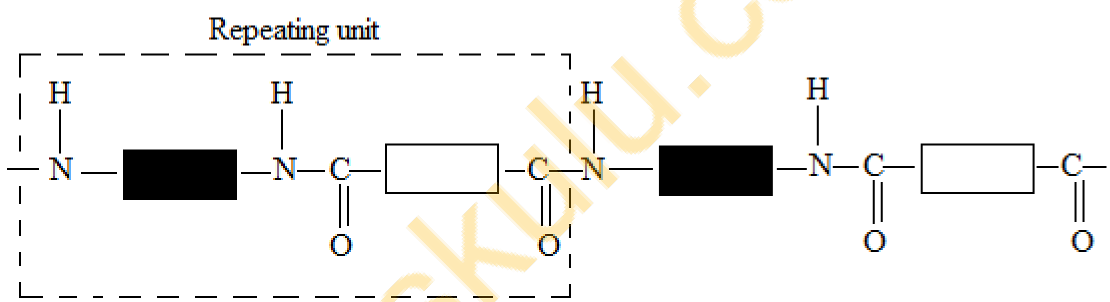
Dicarboxylic acid : 

The structure of nylon is similar to that of protein

Part of the structure of nylon







Uses of nylon

1. Making tough bearing
2. Making clothing
3. Making ropes
4. Making bristle for brushes

Exercise

1. Nylon is a synthetic polymer which has an amide linkage similar to that of proteins
   (a) What is meant by the term synthetic polymer?
   (b) Draw the structure of nylon, showing at least three monomer units.
   Circle the amide linkage in the structure drawn

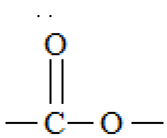
(c) Draw the structures of the two monomers from which nylon is made
(d) By what type of polymerization is nylon formed from its monomers?
(e) What other product is formed?

(II) **Terylene**

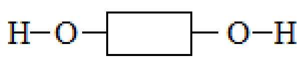
Terylene is typical polyester with ester linkages.

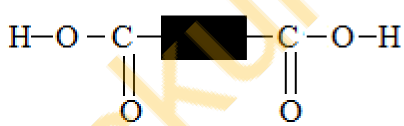
A Polyester is a polymer containing many ester linkages

The ester linkage



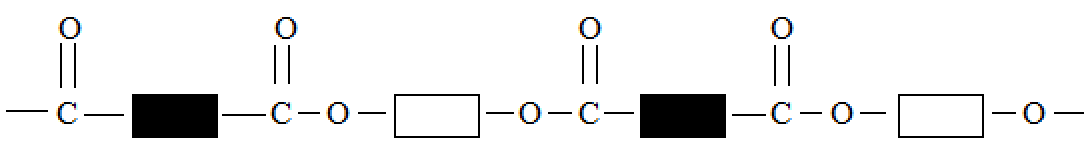
Monomers: Diol and Dicarboxylic acid

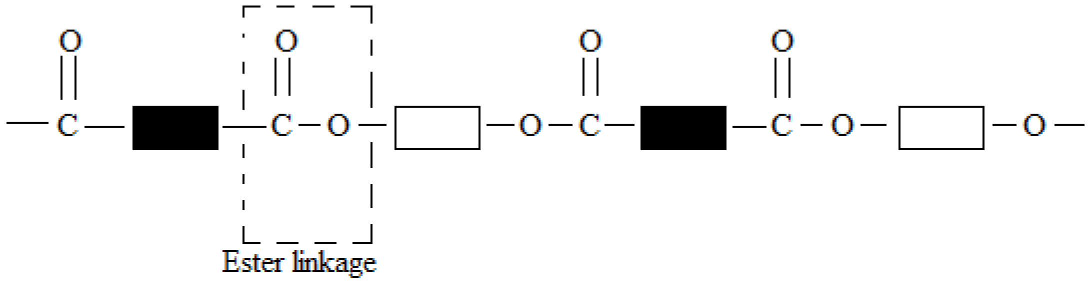
Diol : 

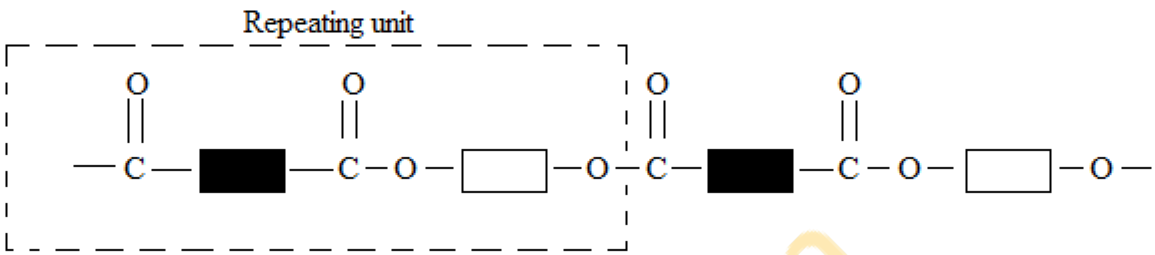
Dicarboxylic acid : 

The structure of Terylene is similar to that of fats.

Part of the structure of Terylene





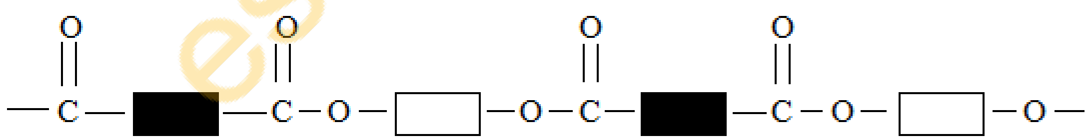


Uses of Terylene

1. Making tents and sails.
2. Making clothing

Exercise

1. Terylene is synthetic polyester which has an ester linkage similar to fats. The structure of a Terylene molecule is shown below



(a) What is a polyester?
(b) Name the elements that make up an ester link
(c) State one use of Terylene
(d) Draw a box around a repeating unit in this Terylene structure
(e) Name the plastic commonly used for making plastic bag
(f) State two reasons why there is need to recycle plastics
(g) Explain why Terylene can cause pollution

Advantages of synthetic polymers

(a) They are durable. They do not rust, corrode or decay
(b) They are lighter than steel, wood or stone
(c) They are thermal and electrical insulators
(d) They are not expensive. They are produced as by-products of oil refining
(e) They are flexible in use

Disadvantages of synthetic polymers

(a) They are non-biodegradable. This means they cannot be decomposed by bacteria.
(b) Non-biodegrability makes the disposal of plastics difficult and this result in pollution problems.
(c) Plastics burn easily and may produce poisonous gases on combustion. They need to be coated with fire retardants to reduce the risk of fire.

Reasons for recycling plastics

- They are difficult to dispose of: Plastic bags do not rot when they are thrown away, so they pollute the environment
- When some plastics burn, they produce harmful gases: For example polyvinylchloride (PVC) gives off fumes of hydrogen chloride when it burns. This would form hydrochloric acid in the eyes and throat.

2. Natural macromolecules

Alternative term: Natural polymers

Natural macromolecules occur in living organisms

Examples of natural macromolecules

- Proteins
- Fats
- Carbohydrates

(a) Proteins

Proteins are made by condensation polymerization

They are condensation polymers

They are poly amides like nylon because they contain the amide linkages

Protein hydrolysis results in amino acids

(b) Fats

Fats are complex esters formed from fatty acids and glycerol

Fats have the structure similar to Terylene and poly carbonates

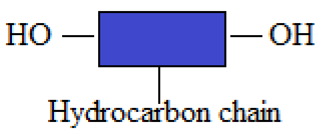
Fat hydrolysis results in fatty acids and glycerol

(c) Carbohydrates

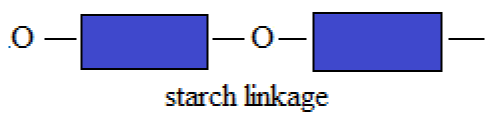
Carbohydrates are sugars which include starch and cellulose

Carbohydrates are formed from simple sugars like glucose

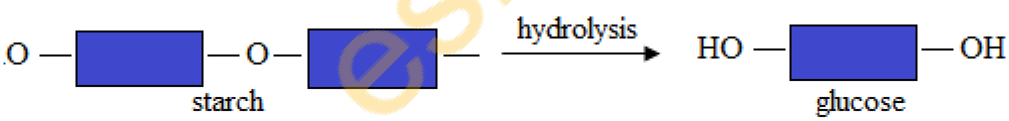
Part of a simple sugar (glucose)



Part of the structure of starch



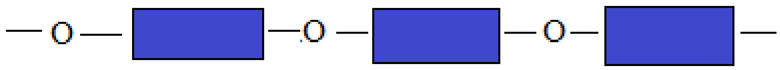
Starch hydrolysis results in glucose



Exercise

1. Some large molecules may be broken down to smaller molecules by hydrolysis.
   (a) What is formed in the hydrolysis of?
      (I) Protein
      (II) Fat
      (III) Starch

(b) If starch is represented as



Complete the following to represent its hydrolysis