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ECZ GRADE 10 CHEMISTRY SUMMARISED NOTES (FOR 5070 & 5124) WITH QUESTIONS AND ANSWERS.

ESKULU ZM

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This document summarizes Chemistry (5070 & 5124)) notes according to the ECZ (Examinations Council Syllabus).

The questions and answers are adapted from actual international past exam papers.

Prepared by William N for eskulu.com

Contact +260978031524

Contents

1. NATURE OF MATTER

1.1 THE THREE STATES OF MATTER

Matter is anything that has mass and occupies space. It exists in three states;

- Solids.
- Liquids.
- Gases.

STATE	SHAPE	VOLUME
Solid	Has a definite shape.	Has a fixed volume.
Liquid	No definite shape. Takes shape of the containing vessel.	Has a fixed volume.
Gas	No definite shape. It easily spreads and fills up the space available in a container.	No fixed volume.

CHANGE OF STATES

MELTING

This is the changing of a substance from a solid into liquid. The temperature at which a substance melts is called its **melting point.**

When a liquid changes into a solid it is said to have **solidified.**

EVAPORATION (VAPORISATION)

This is the change of liquid to gas/vapour. Evaporation of water takes place even at room temperature. Heating speeds up the process of evaporation. The temperature at which a substance boils is called its **boiling point.**

Factors that affect the rate of evaporation;

- Surface area.
- Humidity.
- Wind current.

CONDENSATION

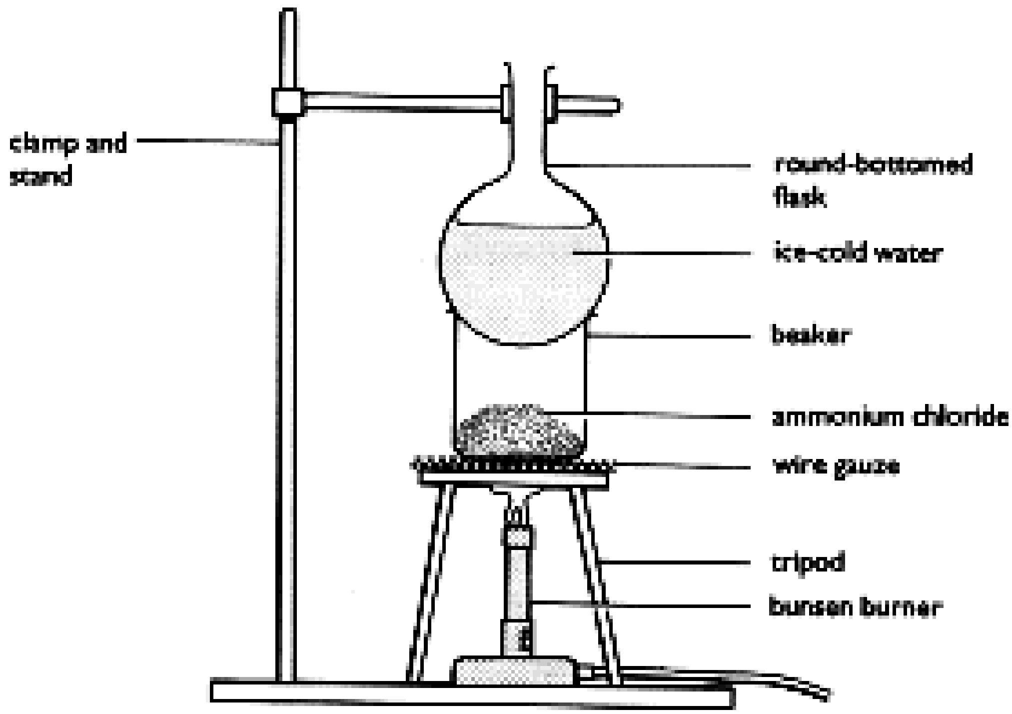
This is the process of vapour changing to a liquid.

SUBLIMATION

As ammonium chloride was heated, there were dense white fumes in the beaker. The solid ammonium chloride changed directly into gas without forming a liquid. On reaching the cold surface of the flask above the beaker the gaseous ammonium chloride turned directly into solid.

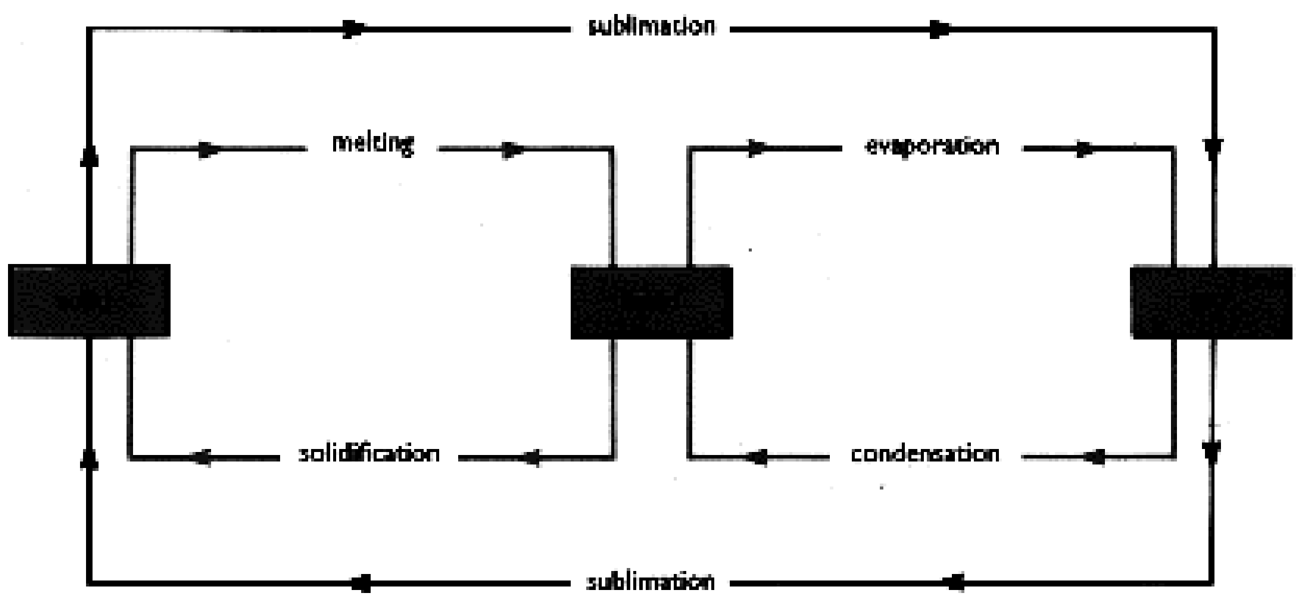
The ice-cold water provides the cold surface which cools the gaseous ammonium chloride.

Ammonium chloride is said to **sublime**.



Sublimation is the change from solid to gas, or gas to solid, without forming a liquid.

Other substances which sublime are **Iodine** and **Ammonium sulphate.**

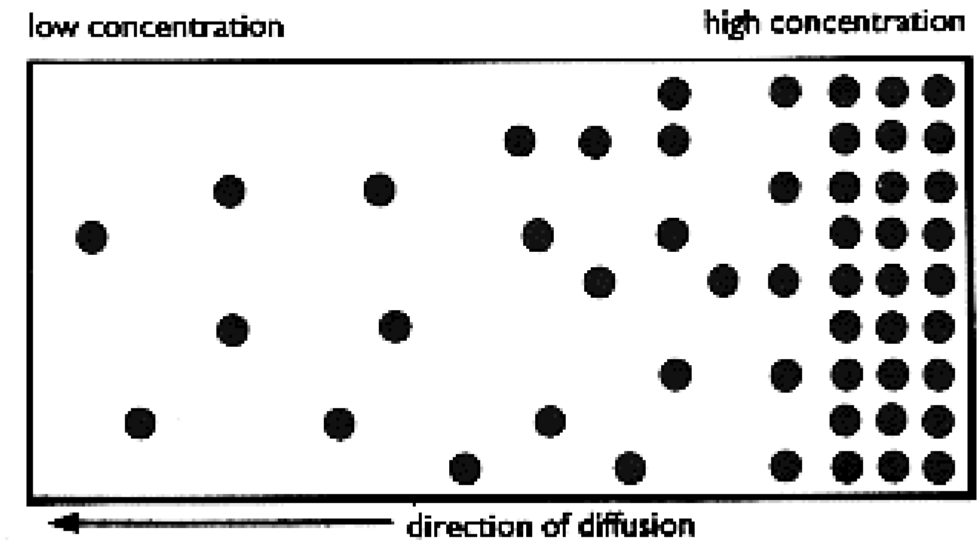


1.2 THE KINETIC THEORY

It states that matter is made up of small discrete particles.

DIFFUSION

This is the movement of particles from an area of high concentration to an area of low concentration, resulting in an even distribution of particles.



a.)Diffusion in gases

Requirements

Beakers, porous pot, hydrogen supply, glass tubing, coloured water, clamp and stand.

Explanation

As the hydrogen is generated into the inverted beaker the level of coloured water in the glass tubing steadily falls. This suggests that the pressure in the porous pot is greater than the pressure in the atmosphere.

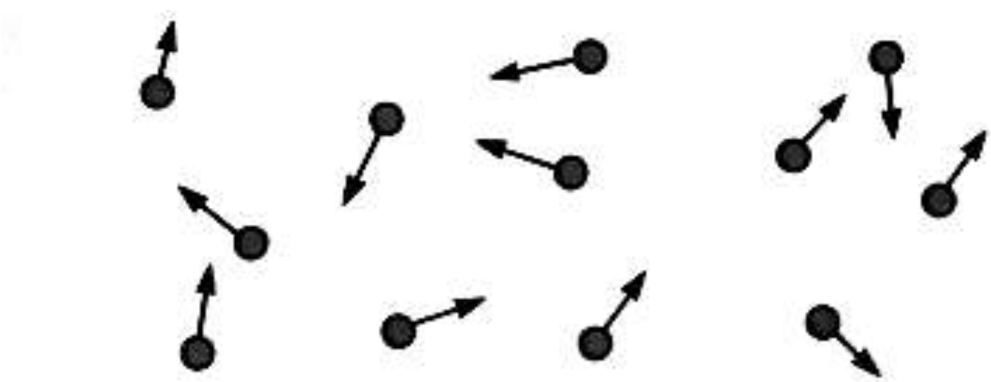
This excess pressure pushes down the column of water in the tube.

WHAT CAUSED THIS EXCESS PRESSURE?

The particles of the hydrogen gas must have crossed the walls of the porous pot, explaining the bubbling which is observed in the beaker. The porous pot contains tiny holes in its walls. It is through these tiny holes that the hydrogen particles passed.

The hydrogen particles in the porous pot created the excess pressure. They must have moved from outside the porous pot to the space inside it.

This means that the particles of hydrogen are not stationary. They are in motion. They moved from an area where they are in high concentration to an area where they are in low concentration.



Particles in a liquid

b.)Diffusion in Liquids

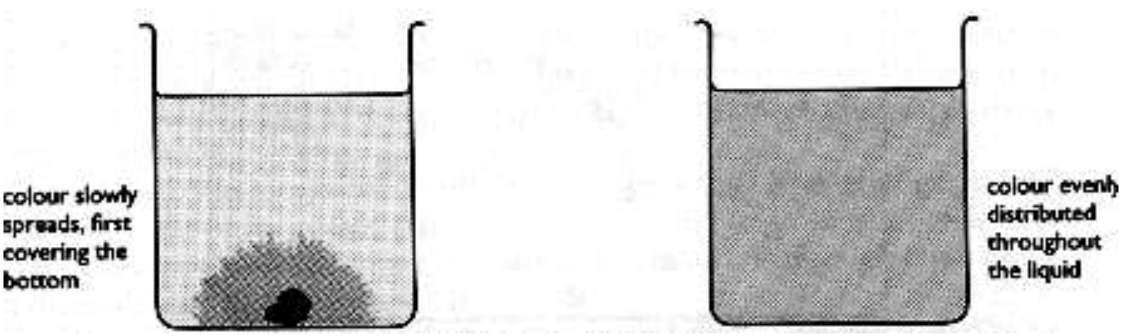
Requirements
Potassium dichromate (VI), Beaker, Water.

Explanation
As the crystal dissolves the colour slowly spreads through the liquid. Due to the effect of gravity the colour first covers the bottom.

It then slowly rises up through the liquid. Eventually the colour distributes itself evenly throughout the liquid.

What do you learn from this? The colour shows the presence of particles of potassium dichromate (VI). It is potassium dichromate (VI) particles which slowly move from where they are in high concentration to where they are in low concentration. This is diffusion in liquids. The particles first cover the bottom as they are pulled by the force of gravity. Then they slowly move upwards and reach all other areas.

Particles in liquids and gases are in continuous motion. The continuous motion of one set of particles within another produces uniform mixing of gases or liquids. Diffusion in liquids is much slower than that in gases.



Osmosis is the diffusion of water particles across a membrane from an area where they are in high concentration to an area where they are in low concentration.

FACTORS AFFECTING THE RATE OF DIFFUSION

The **rate of diffusion** is the amount of gas or liquid diffusing in a unit of time. It is affected by several factors. These factors include temperature, size of the particles and differences in the concentration of particles.

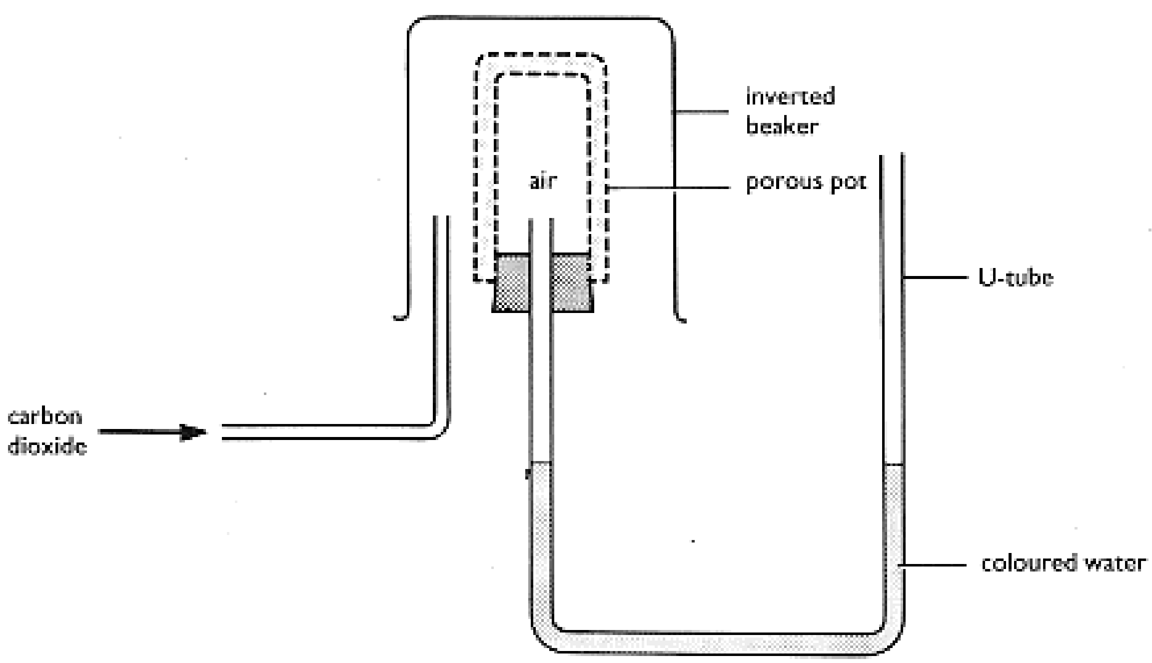
TEMPERATURE

The rate of diffusion increases as the temperature increases. Particles of liquids and gases are in constant motion. This means they possess **kinetic energy**. Heating increases the kinetic energy of the particles. With increased kinetic energy the particles move at higher speeds.

PARTICLE SIZE

Requirements

Beaker, porous pot, carbon dioxide supply, glass U-tube, coloured, clamp and stand.



Explanation

You should have observed that at the beginning the coloured water was at the same level in both arms of the U-tube. This shows that the pressure in the porous pot is equal to the atmospheric pressure.

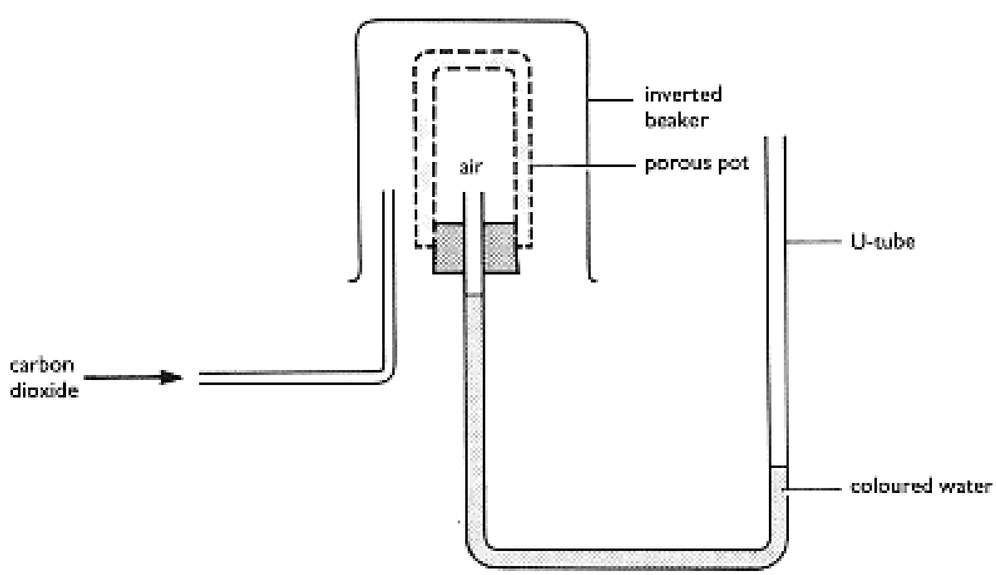
After some time, as the carbon dioxide is introduced into the beaker inverted over the porous pot, the level of coloured water in the arm leading to the porous pot rises and it falls in the other arm.

These results suggest that after about 20 minutes the pressure in the porous pot falls to become less than the atmospheric pressure. Hence the rise in the level of coloured water in the U-tube in the arm leading to the porous pot.
What causes this fall of pressure in the porous pot? The porous pot has very tiny holes. Through these tiny holes carbon dioxide diffuses into the pot and air diffuses out. The fall in pressure in the porous pot suggests that more air diffused out than carbon dioxide diffused in during 20 minutes.

This means that the rate of diffusion of air is greater, the mass is also greater and so is the weight.

Air particles are smaller than those of carbon dioxide. Generally when the size of the particles is greater, the mass is also greater and so is the weight.

<u>Note</u>: Particles which are heavier move more slowly. Those which are lighter move more quickly. Accordingly, those particles with higher speeds diffuse faster and vice versa. Thus, the smaller the particles the faster the rate of diffusion.



If carbon dioxide were substituted with hydrogen in the experiment the results would be reversed. This is because hydrogen particles are lighter than those of air and diffuse faster.

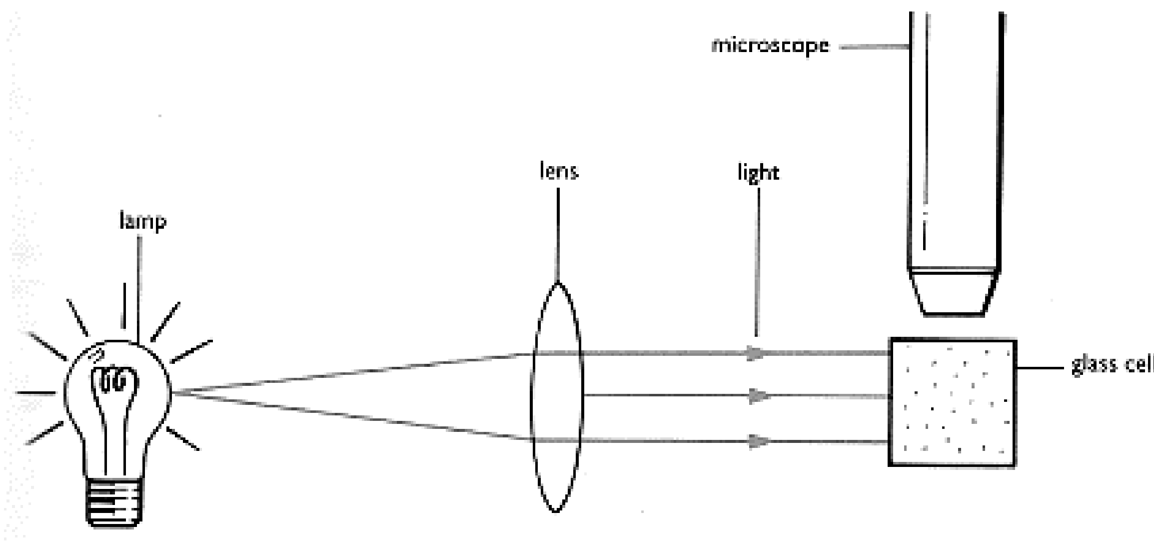
CONCENTRATION

Diffusion is faster if there is a large difference in the concentration of particles between two points

BROWNIAN MOTION

Requirements

Glass cell, source of smoke, lamp, microscope, lens.



Explanation

When you look through the microscope you can actually see bright specks undergoing random motion at high speed. The specks, seen under reflected light, do not maintain any particular direction. The random motion is due to collisions with air particles. The air particles knock the specks of smoke one way, then another and so on.

This is called **Brownian motion.**
Brownian motion is a term used for the continuous random motion of particles, particularly particles of gases and liquids.

PARTICLES AND THE KINETIC THEORY

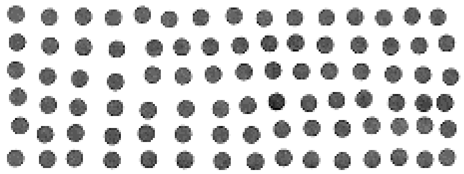
These are the main points of the kinetic theory;

- The particles of which matter is made up of are in constant motion.
- The particles exert strong forces upon each other when they are in contact; attractive forces hold the particles together and repulsive forces cause the particles to move apart.
- The physical state assumed by a substance at ordinary conditions, that is, whether it is solid, liquid or gas, depends mainly on the strength of the forces holding the particles together.
- Particles of different substances have different sizes and different masses; in general larger particles have higher mass and smaller particles lower mass.
- The speed of motion of the particles increases as the temperature increases and decreases as the mass increases.

1.3 THE KINETIC THEORY AND THE STATES OF MATTER

SOLIDS

In solids the particles are closely packed together. The attractive and repulsive forces counterbalance and the particles vibrate about their fixed positions. So solids have a definite shape and fixed volume.

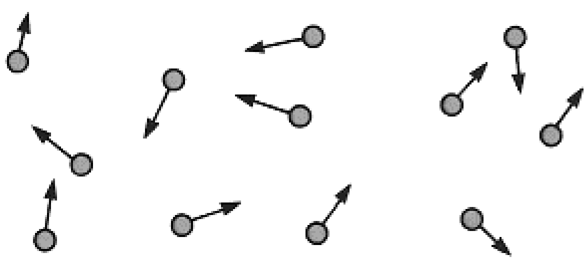


LIQUIDS

In liquids, particles are generally slightly further apart than in solids. They can move freely over short distances. The attractive forces are not strong enough to hold the particles together.

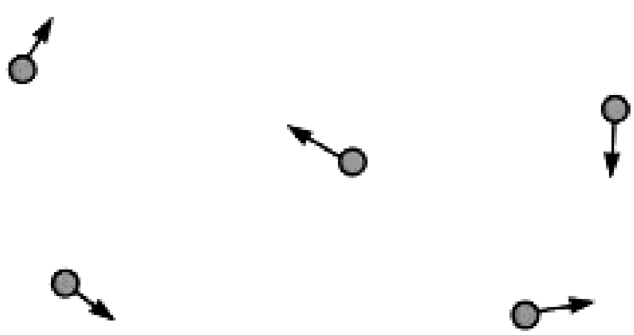
A liquid can therefore flow and take the shape of the containing vessel.

A liquid has a fixed volume but no definite shape.



GASES

In gases, particles are much further apart than in solids or in liquids. They move at very high speeds in the available space. The forces between them are negligible. Gases have no definite shape and no definite volume.



Note: Both gases and liquids have the ability to flow which means they can both be described as **fluids.**

Property	Solids	Liquids	Gases
Shape	Have a definite shape.	Have no definite shape. Take the shape of the container.	Have no definite shape. Spread to fill space available.
Density	Have higher densities than those of gases.	Have higher densities than those of gases.	Have densities less than solids and liquids.
Volume	Have a definite volume.	Have a definite volume.	Volume depends on size of container.

MELTING AND SOLIDIFICATION

Heating a solid substance causes its particles to gain kinetic energy. As a result, the particles vibrate faster and faster. They eventually break free from their average fixed positions and start moving randomly just like the particles in a liquid. Hence, the distance between the particles increases and the solid **melts** to form a liquid.

The temperature at which this occurs is the **melting point** of the solid.

When a liquid is cooled it loses heat to the surroundings. This means, in effect, that the kinetic energy of its particles decreases. This leads to a decrease in the random motion of the particles and a corresponding decrease in the distance between the particles. Eventually the attractive forces between the particles become strong enough to hold them in a regular pattern.

Thus the substance **solidifies.**

EVAPORATION AND BOILING

When a liquid is heated its particles gain kinetic energy and so they move faster and faster. Some particles near the surface gain enough energy to break away from the attractive forces of the other particles to the extent that they escape into the air. This is **evaporation.** Eventually a temperature is reached at which so many particles try to escape from the liquid so rapidly that bubbles of gas begin to form inside the liquid. This is called **boiling.** The temperature at which this occurs is the **boiling point** of the liquid.

Note: Liquids with high boiling and melting points have strong attractive forces between the particles. Liquids with low melting and boiling points have weak forces between the particles. These liquids easily evaporate and are said to be **volatile.**

2. SOLUTIONS AND CRYSTALS

2.1 SOLUTIONS

SOLUTIONS IN WATER

Requirements

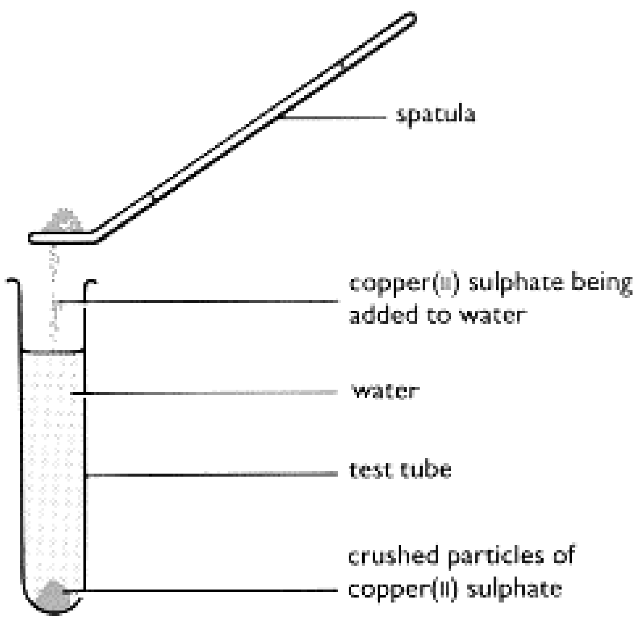
Finely crushed blue copper (II) sulphate crystals, test tube, water, and spatula.

Method

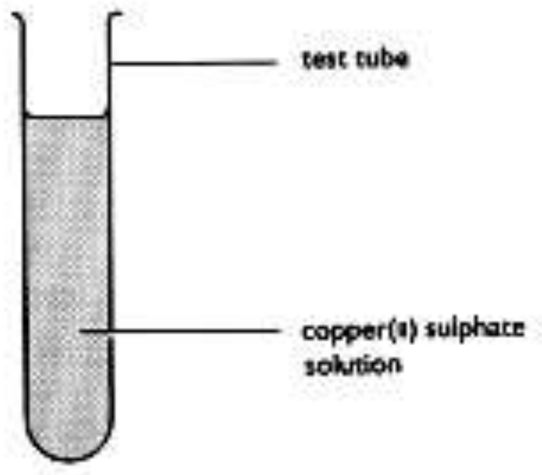
Shake up a little finely crushed copper (II) sulphate with some water in a test tube.

Explanation

When blue copper (II) sulphate is shaken up with water in a test tube, the water turns blue and the crushed copper (II) sulphate crystals are no longer visible.



The copper (II) sulphate **dissolved** in the water. It is said to be **soluble** in water. The substance which dissolves, in this case copper (II) sulphate, is known as a **solute.** The liquid in which it dissolves is known as a **solvent**. A solute is therefore, a substance which dissolves in a given solvent. A solvent is a substance in which substances dissolve. A **solution** is formed when a solution dissolves in a solvent. A solution is a **homogeneous** (uniform) mixture of two or more substances.



Solutions in which water is the solvent are known as **aqueous** solutions.
Chalk does not dissolve in water so it is said to be **insoluble.** The cloudy mixture of chalk powder and water is called a suspension.

A **suspension** is a liquid containing small particles of solid which are spread throughout the liquid. The particles settle down to the bottom of the container on standing, due to the force of gravity.

Fine sand and starch powder also behave like chalk powder when shaken up with water. These mixtures of sand and starch powder with water are **non-uniform mixtures.**

Some solvents and some of their uses

Solute	Solvent	Use of solution
Iodine	Ethanol	As antiseptic on cuts and wounds
Rubber	Petrol/benzene	Mending tyre tubes
Paint	Turpentine/kerosene	Thinning paint
Pigment	Water	Making dyes

Caution: THESE ARE FLAMMABLE.

MISCIBLE AND IMMISCIBLE LIQUIDS

Miscible liquids are liquids that mix with each other completely. Water and ethanol are completely miscible; therefore, they form a uniform mixture.

Liquids which do not mix with each other are known as **immiscible** liquids. Examples of immiscible liquids are water and cooking oil and water and kerosene.

FACTORS AFFECTING THE SOLUBILITY OF GASES IN WATER

The solubility of a gas in water depends on the:

- Pressure
- Temperature
- Nature of the gas

PRESSURE

An increase in pressure makes the gas more soluble.

TEMPERATURE

Gases, unlike solids, are generally more soluble in cold water. Therefore, their solubility increases with a decrease in the temperature of the water. In fact, most gases can be completely driven out of solution by boiling.

NATURE OF THE GAS

Most gases are soluble in water to some extent. Some, however, are more soluble than others. As a general rule, gases that are acidic or alkaline are very soluble in water. Examples of such gases are ammonia (alkaline), hydrogen chloride (acidic) and sulphur dioxide (acidic). Neutral gases are only slightly soluble in water. Examples of such gases include oxygen, nitrogen, hydrogen and helium.

SATURATED AND UNSATURATED SOLUTIONS

An **unsaturated** solution is one which is able to dissolve more solute at that particular temperature.

A **saturated** solution is one which cannot dissolve any more solute at that particular temperature. Excess solute is present.

In most cases when a saturated solution i warmed, it becomes unsaturated and is able to dissolve more solute. For example common salt is more soluble in hot water than in cold water. This is true for many solid solutes. They are more soluble in hot than in cold solvents. Examples of such substances include potassium nitrate and potassium chloride. Some substances, however, are less soluble in hot water. Examples include calcium sulphate and all gases.

2.2 CRYSTALS

A **crystal** is a solid that consists of particles arranged in an orderly and repetitive manner resulting in a regular shape. Crystals are formed from saturated solutions. The process of crystal formation is known as **crystallisation.**

Crystallisation takes place when a saturated solution is cooled, when it is evaporated to dryness by heating or when it is allowed to evaporate slowly at room temperature. When a saturated solution evaporates, the remaining solution cannot hold all the solute in solution. The excess solid, therefore, settles out as crystals.

When a saturated solution is heated to cooled fast, fairly good-sized crystals with well-defined shapes are formed.

When a saturated solution is heated to dryness, evaporation is very fast. The crystals formed are small and their shape is not well-defined. Gentle heating should be used because hot liquid spurts out of the evaporating dish if it is heated strongly. This is also called spitting.

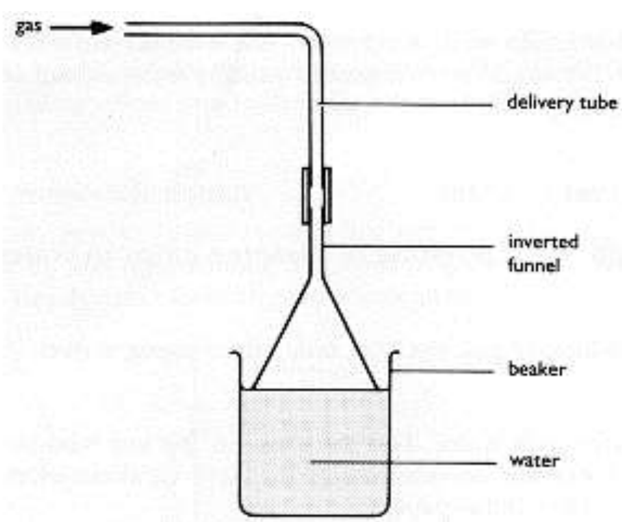
When a saturated solution is allowed to evaporate slowly at room temperature, the crystals formed are large and have a well-defined shape. A loose paper cover is necessary to slow down evaporation and to prevent dust particles from entering the solution.

Dust particles would interfere with the process of crystallisation as the solute would be deposited on the dust particles resulting in imperfect crystals being formed.

HOW TO GROW CRYSTALS OF COPPER (II) SULPHATE

Requirements

Copper (II) sulphate, a small crystal of copper (II) sulphate, water, string, glass rod, beaker.



Explanation

After a number of days, the small crystal becomes bigger. This is because as the solvent evaporates, the solute gets deposited on all sides of the crystal. The shape of the crystal, however, does not change. It just becomes more clearly defined. A large crystal has, therefore, grown from a small crystal hanging in a saturated solution. The small crystal is known as a **seed crystal.**

Once again, dust must be kept out of the solution to prevent unwanted crystallisation that is crystals forming in solution other than on the seed crystal. Crystals of many other substances can be grown in the same way.

EXAM TYPE QUESTIONS WITH ANSWERS I

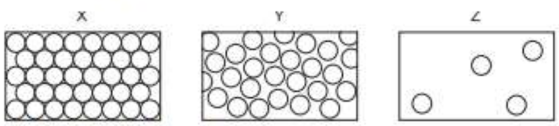
Attempt the following questions before looking at the answers

1 In which changes do the particles move further apart?

gas $\underset{Y}{\overset{W}{\rightleftharpoons}}$ liquid $\underset{Z}{\overset{X}{\rightleftharpoons}}$ solid

A W and X **B** W and Z **C** X and Y **D** Y and Z

2 Diagrams X, Y and Z represent the three states of matter.



Which change occurs during boiling?

A X to Y **B** Y to Z **C** Z to X **D** Z to Y

3 Which change of state takes place during evaporation?

A gas to liquid
B liquid to gas
C liquid to solid
D solid to gas

4 In which process do particles move closer together but remain in motion?

A condensation
B diffusion
C evaporation
D freezing

5 The changes that occur when a substance changes state are shown below.

solid $\underset{Z}{\overset{W}{\rightleftharpoons}}$ liquid $\underset{Y}{\overset{X}{\rightleftharpoons}}$ gas

Which process, W, X, Y or Z, is occurring in the following four situations?

1 Butter melts on a warm day.
2 Water condenses on a cold surface.
3 The volume of liquid ethanol in an open beaker reduces.
4 Ice forms inside a freezer.

	1	2	3	4
A	W	X	Y	Z
B	W	Y	X	Z
C	X	Y	Z	W
D	X	Z	Y	W

6 The diagram shows a sugar lump in a cup of tea.

Which two processes must happen to spread the sugar evenly in the tea?

	first process	second process
A	diffusion	dissolving
B	dissolving	diffusion
C	dissolving	melting
D	melting	diffusion

ANSWERS

1. D
2. B
3. B
4. A
5. B
6. b

3. SEPARATION TECHNIQUES

Certain methods are used to determine the purity of the products and to purify those which are found to be impure. These methods are called **separation techniques**.

3.1 CRITERIA OF PURITY

This refers to those physical properties which are characteristic of a pure substance. Impure substances show variations in physical properties.

MELTING POINT

Pure substances have sharp **melting points.** During melting the temperature remains constant. If there is any change in temperature during melting, it will be a small change of between 0.5-2°C. Within this range the substance should melt completely.

Usually a temperature – time graph is plotted. This is done by obtaining temperature readings at different time intervals until the temperature rises some 5-10°C above the constant temperature. Such a graph will clearly show the melting point of a substance. The figure below shows the melting point graph for the pure substance, naphthalene.

MELTING POINT OF A PURE SUBSTANCE

The graph suddenly levels out and then sharply rises. This shows that the melting point is **sharp** and indicates that the substance whose melting point is being determined is pure.
Pure naphthalene melts at 80°C.

MELTING POINT OF A MIXTURE

This shows that impurities affect the melting point of substances. In general, impurities **lower** the melting point of substances. An impurity may also prolong the period of melting so that, instead of a rapid change, melting is slow and over a range of temperature.

BOILING POINT

This is the temperature at which the saturated vapour pressure of a liquid equals the external atmospheric temperature. As a result, bubbles form in the liquid and the temperature remains constant until all the liquid has evaporated.
As the boiling point of a liquid depends on the external atmospheric pressure, boiling points are usually quoted for standard atmospheric pressure (760 mmHg).

Boiling points of some liquids

Substance	Boiling point (°C)
Acetone	56
Ethanol	78
Benzene	80
Water	100
Toluene	111
Acetic acid	118
Tartaric acid	170
Glycerol	290

If the temperature of a boiling liquid is the same as the known temperature of the vapour, then the liquid is pure. In general impurities **raise** the boiling point of a liquid.

DENSITY

The **density** of a substance is its mass per unit volume. Density is measured in grams per cubic centimetre (g/cm^3).

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

The density of a pure substance (solid, liquid or gas) is always constant. If the density of the substance being determined is of the same value as that of the pure substance, then the substance whose density is being determined is pure. However, if the value obtained is different from that of the pure substance, then the substance is impure. Impurities in substances **lower or raise** their densities. This mainly depends on the nature and quantity of the impurities.

MELTING AND BOILING POINTS OF AN UNKNOWN SUBSTANCE

Densities Of Some Substances

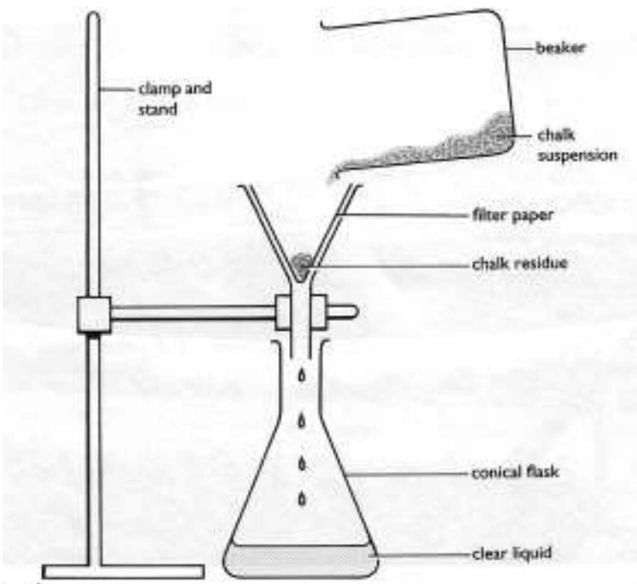
Substance	Density (g/cm^3)
Ethanol	0.789
Water	1.0
Silicon	2.33
Zinc	7.14
Iron	7.86
Copper	8.92
Silver	10.5
Mercury	13.6

3.2 METHODS OF PURIFICATION

SEPARATING SOLIDS FROM LIQUIDS

Filtration

Separating A Solid From A Liquid By Filtration



The method used to separate chalk from its suspension is called **filtration.** **Filtration** is a method of separating solid particles from a fluid using a filter. The substance which remains on the filter paper is called a **residue** and the clear liquid collected after filtration is called a **filtrate.**

Filtration is also used in the manufacture of sulphuric acid by the contact process. Sulphur dioxide, mixed with an excess of air is purified by filtration to remove particles of dust and other impurities. Another important use of filtration is the purification of water at the water works.

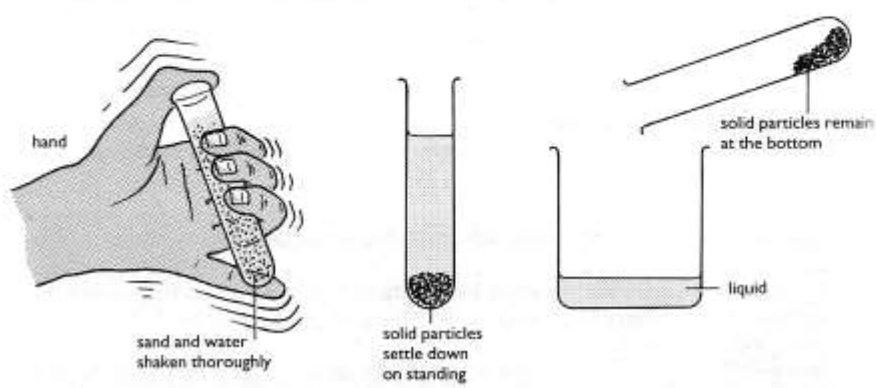
PURIFICATION OF WATER AT THE WATER WORKS

Water which is for human consumption is mainly obtained from rivers and lakes. The water from these sources contains impurities. The mineral matter dissolved in it, bacteria and suspended impurities are the main types of impurity found in the water.

Suspended impurities are removed by filtration. This is achieved by passing the water through large filter beds made from layers of gravel and sand. The figure below shows a model of a sand filter.

Filtration does not remove all the suspended impurities. The excess impurities which seep through the layers of gravel and sand are removed by adding aluminium sulphate. This causes coagulation (sticking together) of the suspended matter and the formation of a **precipitate** of aluminium hydroxide which settles out. This is a physical process called **sedimentation.** In order to kill bacteria, water is sterilised by means of chlorine. Where dissolved mineral matter is injurious to health, it is also removed. In cases where the water is acidic, lime is added to neutralise the acid.

DECANTATION



When sand and water are shaken thoroughly, a cloudy suspension is produced. On standing, the solid particles of sand settle down leaving a clear liquid on top.

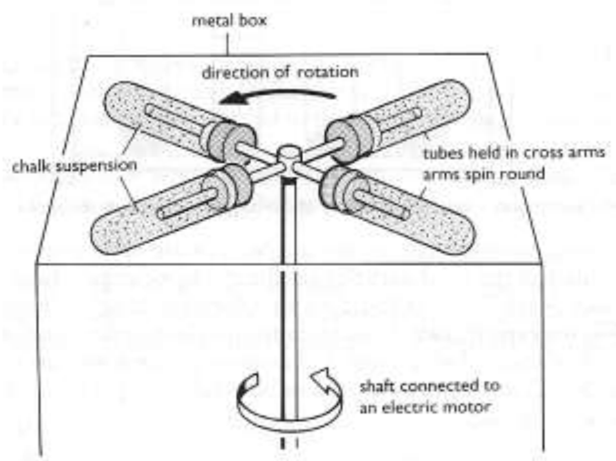
When solid impurities in liquids settle down, the liquid can, with care, be poured off, leaving the solids at the bottom of the container. This method is called **decantation**.

One application of decantation is the separation of wine from its sediments.

SEDIMENTATION

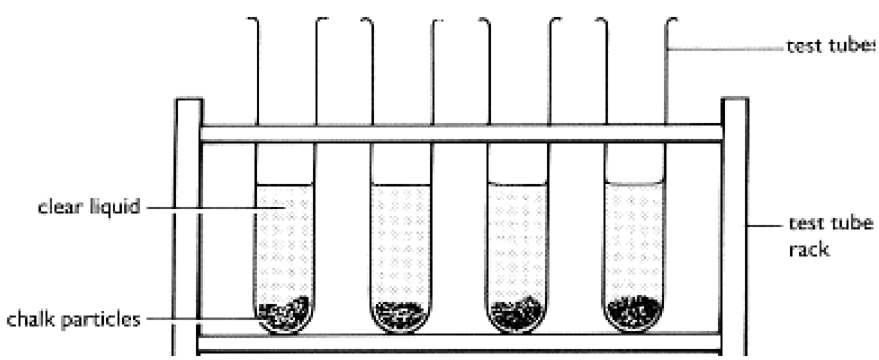
Sedimentation is the term used to describe the settling down of the solid particles in a liquid to produce a flocculent leaving a clear liquid on the top which can be tapped off.

SEPARATING A MIXTURE BY SEDIMENTATION



Centrifuge in use

The contents of the first four test tubes do not settle down after standing for five minutes. However the contents of the test tubes after rotating for five minutes in a centrifuge settle down leaving a clear liquid on top. This process is called **centrifugation.**



Sedimentation – test tubes on standing for five minutes

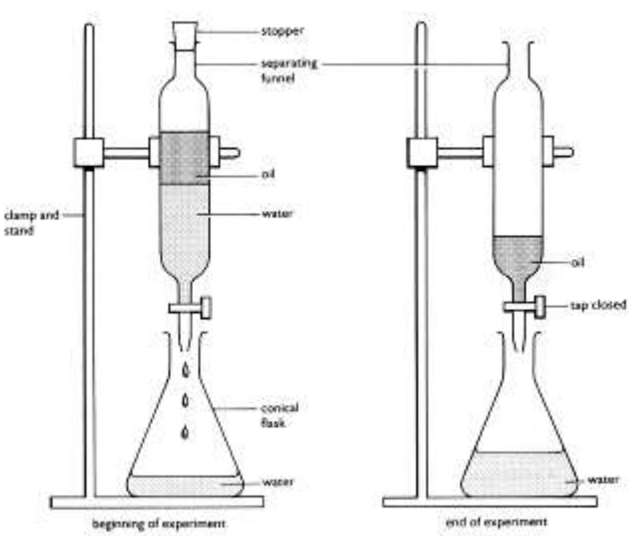
In a centrifuge, tubes containing the suspension are spun round very fast in a metal box, so that all the solid particles are flung into a compact heap at the bottom of the test tubes, leaving a clear liquid on top which can then be decanted.

The clear liquid is called the **centrifuge**. Centrifugation quickens sedimentation. Centrifugation has the advantage of being faster and easier to perform and the residue is more readily available.

Centrifugation is used in hospitals to separate the red corpuscles from the plasma in blood.

SEPARATING LIQUIDS FROM LIQUIDS

Separating Immiscible Liquids Using A Separating Funnel



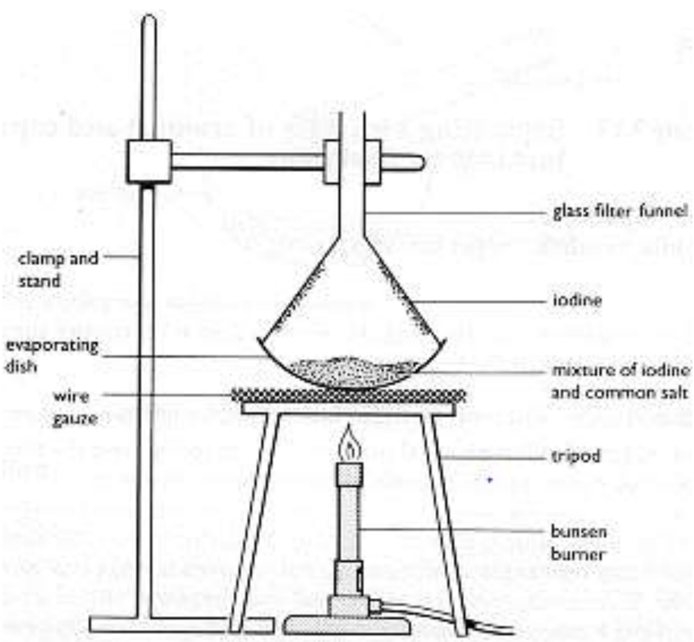
When the separating funnel is shaken vigorously the oil seems to mix with the water. On standing, the oil and water separate forming an upper layer of oil and a lower layer of water. With a mixture of two liquids which do not mix completely, such as oil and water, decantation could be used to separate the mixture but the more accurate method of separating such a mixture is by using a separating funnel.

A **separating funnel** is a piece of apparatus which is used to separate mixtures of liquids which do not mix completely (immiscible liquids). The denser liquid forms the lower layer and is drained off first. The remaining liquid can then be transferred into another container.

METHODS OF SEPARATING MIXTURES OF SOLIDS

Sublimation

Separating A Mixture Of Iodine And Common Salt By Sublimation



Separating two solids by sublimation

When a mixture of iodine and common salt is heated, the iodine changes from solid to vapour. The vapour sublimes on the cold surface of the glass filter funnel and the solid reappears.

It is carried out in a fume cupboard to avoid vapours of iodine spreading into the air. Iodine vapour is poisonous.

Iodine is separated from common salt. Heating for a long time completely separates the two substances.

This separation is possible because one of the substances in the mixture **sublimes**. This process is normally used to separate substances which sublime from those which do not.

Ammonium chloride, naphthalene, ice below 0°C and solid carbon dioxide (dry ice) are other examples of substances which sublime.

FLOATATION

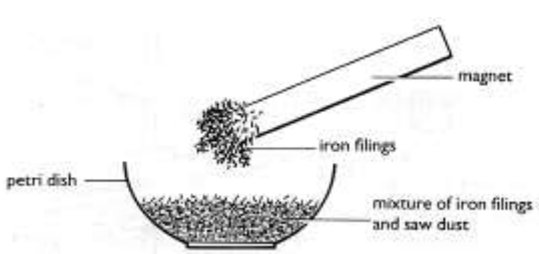
Separating A Mixture Of Sawdust And Copper Turnings By Floatation

The sawdust floats while the copper turnings sink to the bottom of the beaker. The sawdust can be recovered and the water decanted in order to obtain the copper turnings. In this way sawdust is separated from copper turnings.

This method of separating mixtures is used when one component of a mixture is denser than the liquid and the other is less dense than the liquid. Sawdust being less dense will float and the copper turnings being denser will sink. This method of separating mixtures is called **floatation**.

MAGNETISM

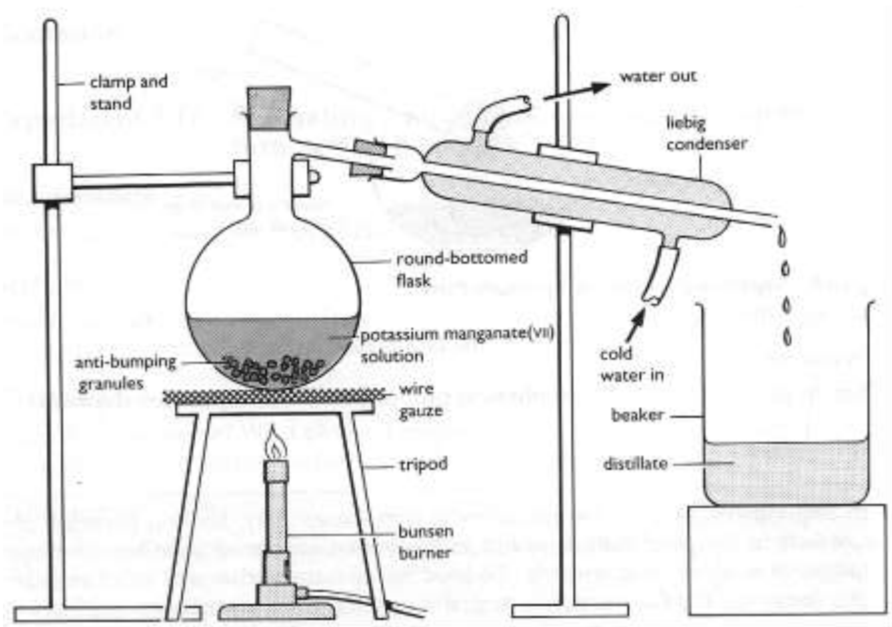
Separating Iron Fillings From Sawdust By Magnetism



The magnet attracts iron fillings only, leaving particles of sawdust in the petri dish. Iron fillings have been separated from sawdust by **magnetism**. Hence a magnet can be used to separate mixtures of substances if one component of the mixture is magnetic and the other one non-magnetic.

DISTILLATION

OBTAINING PURE WATER FROM POTASSIUM MANGANATE (VII) SOLUTION



Distillation

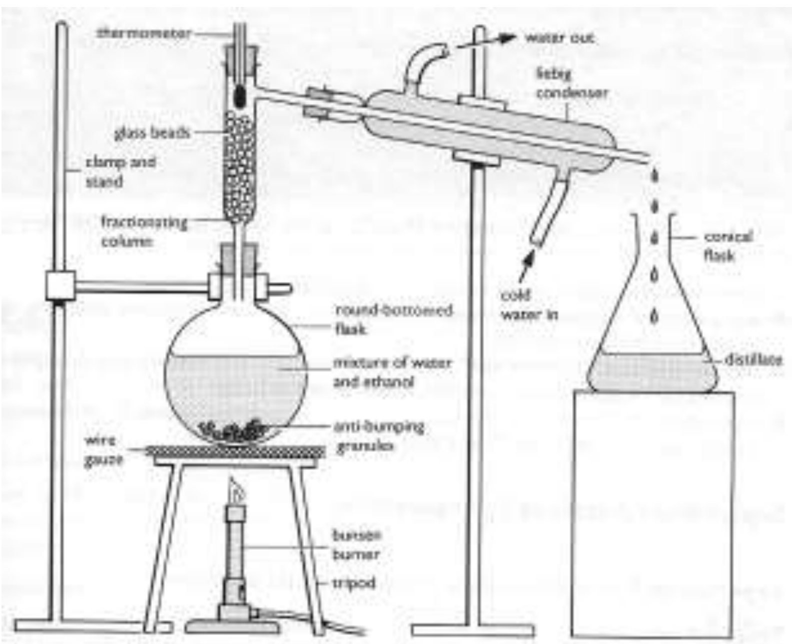
The water boils and evaporates forming vapour, leaving behind potassium manganate (VII). The water vapour passes into the condenser where it is cooled and turned back to water which drips into the collecting beaker. Water is separated from potassium manganate (VII) solution.

This process is called **distillation**. Distillation is the process of vaporising a liquid and then condensing vapour.
The colour of the solution in the distillation flask is still the same, or even darker after heating has stopped, because the potassium manganate (VII) does not distil.
The colourless liquid collected after distillation is called **distillate.**

If boiling is likely to be irregular (a phenomenon called bumping), a few granules of porcelain must be put in the flask to reduce bumping. They are known as anti-bumping granules.

SEPARATING A MIXTURE OF ETHANOL AND WATER

FRACTIONAL DISTILLATION



Fractional distillation

Ethanol has been separated from water. Ethanol and water have different boiling points. Ethanol boils at 78°C while water boils at 100°C (at standard atmospheric pressure). Therefore when the temperature reaches 78°C 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 to the flask. Only the ethanol vapour reaches the top of the fractionating column and enters the liebig condenser where it condenses. The liquid ethanol finally collects in the conical flask. Ethanol from the condenser when put in the spirit lamp and ignited, burns. The other liquid, water, which remains in the round-bottomed flask, does not burn.

This process of separating a mixture of liquids is called **fractional distillation**. This separation of a mixture of liquids depends on the difference in the boiling points. As the temperature begins to rise beyond the boiling point of ethanol we stop collecting the distillate to avoid collecting water.

It should be noted that the thermometer is placed at the top of the fractionating column so that it registers the temperature of the ethanol vapour.

Fractionating column and liebig condenser

The **fractionating column** is a long tube filled with small glass beads. The glass beads provide a large surface area for condensation to allow efficient separation of the components in their mixture.

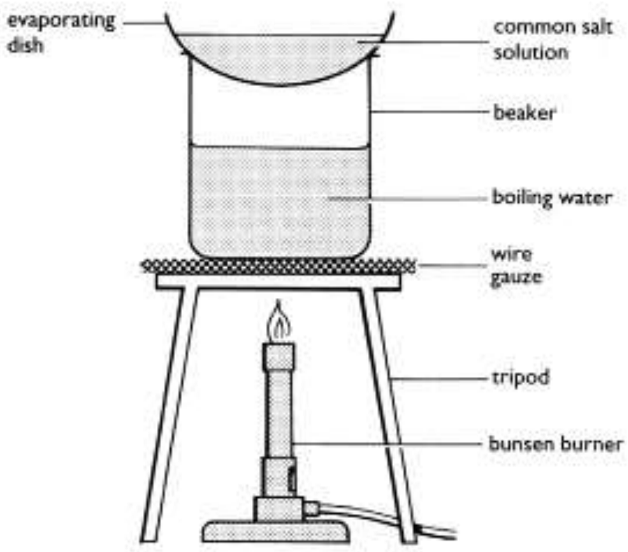
The **liebig condenser** is kept in the slanting position to avoid the distillate formed by condensation from running back into the fractionating column; and also to ensure that cold water completely surrounds the inner glass tube where the vapour passes in order to provide maximum cooling and avoid loss of vapour.

Some industrial applications of fractional distillation:

- Separation of various components of crude oil.
- The manufacture of spirits such as whisky, rum, gin and brandy.
- Separation of liquid air into nitrogen and oxygen.

SEPARATING SUBSTANCES BY EVAPORATION

OBTAINING A SOLUTE FROM ITS SOLUTION



The steam bath method of heating is preferred to avoid spitting of the salt which would occur if direct heating was used.
In order to separate crystals of common salt from solution we evaporate the solution to dryness. Solids which are soluble in water can be obtained from the resulting solutions by using this method, known as **evaporation.**

SEPARATING MIXTURES OF SUBSTANCES BY PAPER CHROMATOGRAPHY

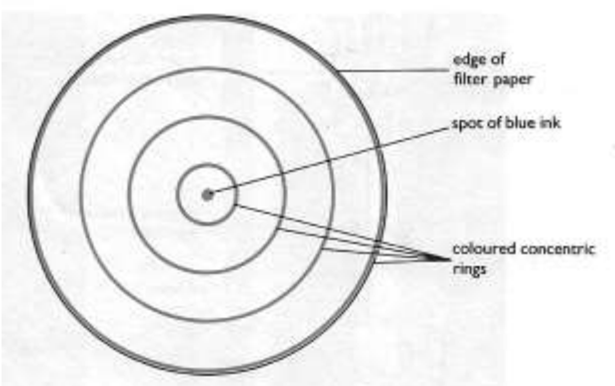
Chromatography is a technique for separating mixtures of solutes using a solvent and a separating medium. In the case of paper chromatography the separating medium is paper.

SEPARATING DYES IN INK BY RADIAL PAPER CHROMATOGRAPHY

RADIAL PAPER CHROMATOGRAPHY

The particles in ink are separated from the spot in the centre and form coloured concentric rings. This separation is called **radial** paper chromatography.

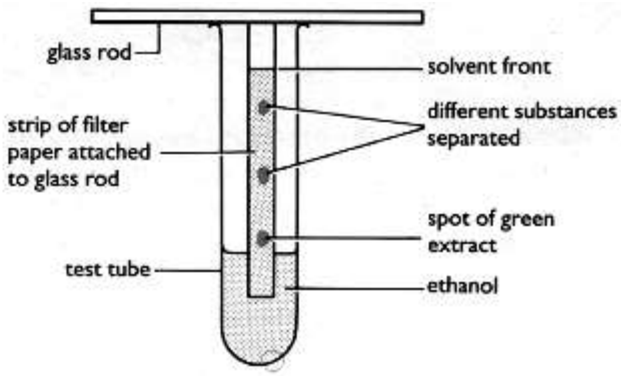
If the outside ring is yellow, it shows that the yellow substance is the most soluble in the solvent. When the ethanol drips onto the ink spot, it dissolves the yellow substance most easily and carries it the furthest way. The ethanol eventually separates all the colours, resulting in the formation of coloured concentric rings.



Result of radial paper chromatography

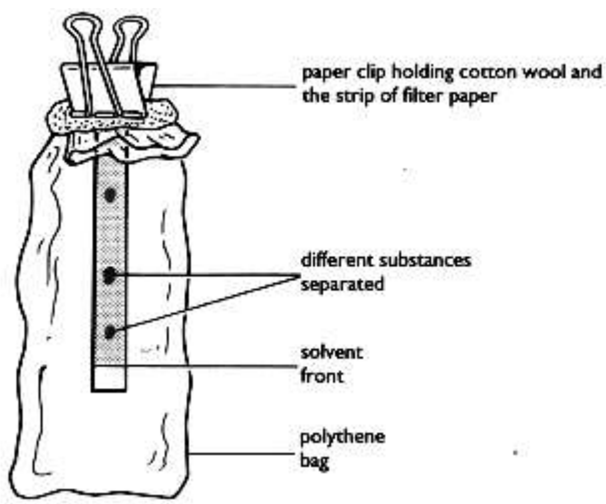
ASCENDING PAPER CHROMATOGRAPHY

The solute compounds are carried along by the solvent being continuously absorbed by the paper and redissolved. Compounds that are strongly absorbed tend to lag behind; compounds weakly absorbed are carried along more rapidly by the advancing solvent. The solute compounds eventually become separated at different levels forming coloured bands. This is called **ascending** paper chromatography.



Result of ascending paper chromatography

DESCENDING PAPER CHROMATOGRAPHY



Result of Descending Paper Chromatography

The solvent moves downwards, this is called **descending** paper chromatography. The solvent descends the strip of filter paper resulting in the separation of the components of chlorophyll. The polythene bag keeps the air saturated with solvent vapour to reduce the evaporation of the moving solvent.

Descending paper chromatography has an advantage because as the solvent flows down the paper by capillary action, it is aided by gravity. This speeds up the rate of movement of particles and the coloured bands are thus formed much more easily and faster than in ascending paper chromatography.

FACTORS AFFECTING CHROMATOGRAPHY

Speed of the moving solvent (also known as the **moving phase**) depends on;

- The separating medium (also known as the **stationary phase**),
- The surface tension of the solvent,
- The viscosity of the solvent.

The more soluble particles in a particular solvent move faster with the solvent as it soaks throughout the filter paper. Hence the particles move at different speeds and so they gradually become separate.

In general, all types of chromatography involve **two phases**, namely the stationary phase and the moving phase.

An **absorbent material** provides the stationary phase. If the stationary phase is paper, it is called paper chromatography. The absorbent material in this case is paper. The piece of paper used in chromatography and showing the results of the separation, is called a **chromatography**.

However, paper is not the only stationary phase to be used in chromatography. A glass tube or column packed with absorbent material, like alumina, is an alternative that is used in descending **column chromatography**.

Chromatography also requires the moving phase which is the solvent. Essentially the separation depends on the competition between the moving phase and the stationary phase for molecules of the mixture being tested.

R_f VALUE

R_f value is the distance travelled by a given component divided by the distance travelled by the solvent front.

WORKING OUT R_f VALUE

For example the R_f value for component A above can be shown as:

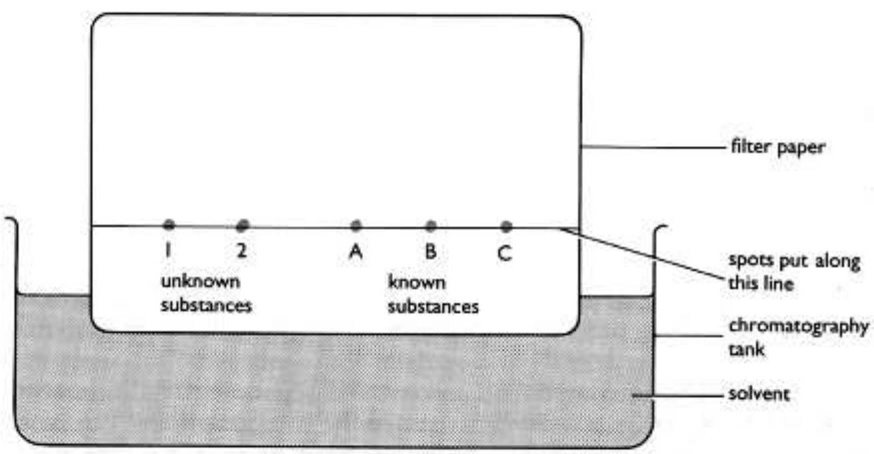
$$R_f \text{ (component A)} = \frac{X}{Y} = \frac{distance\ moved\ by\ component\ A}{distance\ moved\ by\ solvent\ front}$$

For a given system of component, solvent and separating medium, at a given temperature, the **R_f value** is a characteristic of the component and can be used to identify the component whether it is on its own or in a mixture with other substances.

APPLICATIONS OF CHROMATOGRAPHY

It can be used to identify substances. To do this, drops of several different solutions are usually spotted on the filter paper along a line. Some of the solutions contain known substances and others contain unknown substances. The unknown substances are the ones to be analysed.

The figure below shows a filter paper with two spots with mixtures 1 and 2 and three spots of known substances A, B and C. Substances A, B and C were chosen because the mixtures 1 and 2 are suspected to contain the substance A, B or C. After the paper has been prepared in this way, its lower end is dipped in the solvent in a chromatography tank, taking care that the spots do not come into direct contact with the solvent.



An experiment to identify two unknown substances by paper chromatography

CHROMATOGRAPHY SHOWING THE RESULTS OF THE IDENTIFICATION

The paper might look like that.
The results show that the known substances are present in the unknown substances and that the unknown substance 1 is a mixture of A and C, and the unknown substance 2 is a mixture of A and B.

Some mixtures contain colourless substances and they produce colourless chromatograms. The colourless components may be detected by viewing the chromatogram under ultraviolet radiation when they may fluoresce or by spraying the chromatogram with a substance that reacts with the components to give them colour.

Some other applications of chromatography include:

- Identifying the flavouring components in foodstuffs,
- Separating pigments from plants,
- Separating antibiotic drugs from their growing media,
- Separating amino acids from proteins.

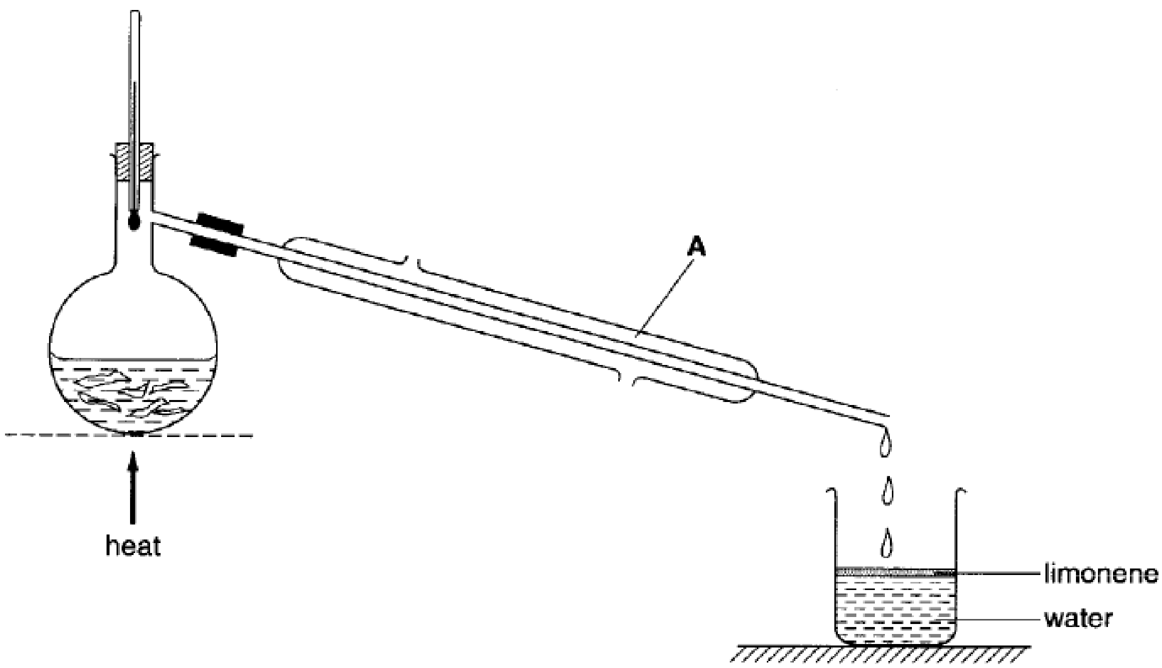
THE IMPORTANCE OF PURITY OF SUBSTANCES

Purity of substances such as food, drugs and water is of vital importance 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.

EXAM TYPE QUESTIONS WITH ANSWERS II

Attempt the following questions before looking at the answers:

1. Limonene is a liquid hydrocarbon found in orange peel. It can be extracted by boiling the orange peel with water, using the apparatus shown below. The mixture of limonene and water distils at a temperature which is 1 °C below the boiling point of water.



(a) (i) State the name of the piece of apparatus labelled **A**.

....[1]

(ii) Suggest what the reading on the thermometer will be when the limonene-water mixture is being distilled.

[1]

What information in the diagram shows this?

....[1]

2 What is always true for a pure substance?

A It always boils at 100 °C.

B It contains only one type of atom.

C It has a sharp melting point.

D It is solid at room temperature.

3 Chromatography is used to identify simple carbohydrates, such as sugars, in plant material.

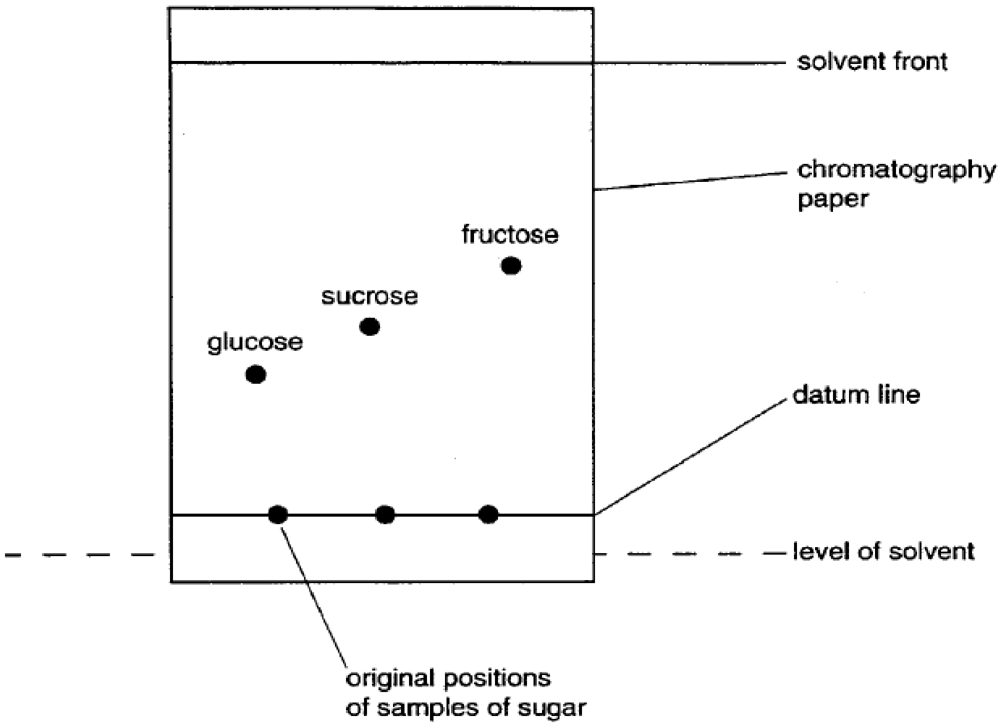


Fig. 5.2

A leaf is ground with 50% aqueous alcohol to give a colourless solution of the sugars. This solution is concentrated and a chromatogram is obtained. The paper is sprayed with resorcinol solution.

(a) **(i)** A common use of ethanol is in alcoholic drinks. In this experiment it is used as a solvent. Give one other use.

....[1]

(ii) Why is the datum line drawn in pencil?

....[1]

iii) Suggest a reason why it is necessary to spray the chromatogram with resorcinol.

....

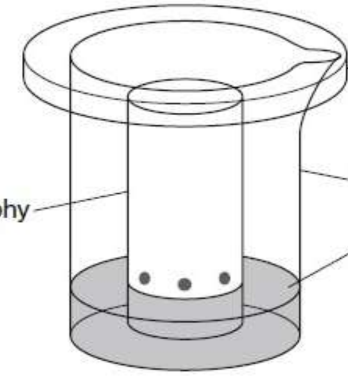
....[2]

iv) Describe how chromatography could be used to show that the hydrolysis of starch produces only one sugar, glucose.

....

....[2]

(b) Amino acids are colourless and can be separated and identified by chromatography.



What additional apparatus is required to identify the amino acids present in a mixture?

A a locating agent

B a ruler

C a ruler and a locating agent

D neither a ruler or a locating agent

ANSWERS

1. (i) **Liebig condenser.**
   (ii) **99.**
   (iii) **Limonene floats on water / on top of the water**.

2. **C**.
3. (a)
   (i) **fuel or making esters or antiseptic or ethanoic acid or vinegar or thermometers**.
   (ii) **does not dissolve or does not contain dyes**.
   (iii) two of these;
   to develop it or locating agent sample are colourless
   to make them visible
   (iv) any two of these;
   only one spot
   same position or Rf value compare with glucose
   (b) **C**.

4. THE LANGUAGE OF CHEMISTRY

4.1 PHYSICAL AND CHEMICAL CHANGES

PHYSICAL CHANGES

Freezing, evaporation, condensation and sublimation are examples of changes of state. Such changes are called **physical change.** A physical change is one which no new substance is formed. They can easily be reversed, for example water can change into ice and ice back into water.

When sulphur is heated, no new substance is formed. Sulphur melts, but the molten sulphur changes back into solid on cooling. Other examples of substances which undergo physical changes on heating are iodine, which sublimes, and candle wax which melts and then solidifies on cooling.

CHEMICAL CHANGES

When paper burns in air, it changes into an entirely different substance-ash. This is an example of a **chemical change.** A chemical change is one which a new substance is formed. In many chemical changes, energy is given out or absorbed. This is usually in the form of heat or light. Chemical changes cannot easily be reversed. It is not possible to change the ash back to paper by any known means.

Magnesium changes into a white ash when it burns. This new substance is formed as a result of magnesium chemically combining with the oxygen of the air. Since the magnesium combines with the oxygen of the air to form white ash, we can say that the composition of the new substance is different from that of magnesium. The two substances differ in their physical and chemical properties. Other examples of substances which undergo chemical changes on heating are mercuric oxide and copper (II) carbonate. Mercuric oxide forms oxygen and mercury, and copper (II) carbonate forms copper oxide and carbon dioxide.

Physical Change	Chemical Change
No new substance is formed.	A new substance is formed.
Usually the change is reversible.	Usually the change is not easily reversible.
Usually no change is given out or taken in.	Usually energy is given out or taken in.
The mass of the substance remains the same.	The mass of the new substance is different from that of the starting substance.

4.2 ELEMENTS, MIXTURES AND COMPOUNDS

ELEMENTS

An element is a pure substance which cannot be split up into two or more other simpler substances by chemical means. Magnesium and oxygen are examples of elements.

MIXTURES

A mixture is a substance which consists of two or more substances not chemically combined. It is often easy to separate a mixture into its components because each component keeps its own properties. In a mixture, the particles of each substance remain separate and the number of each can vary.

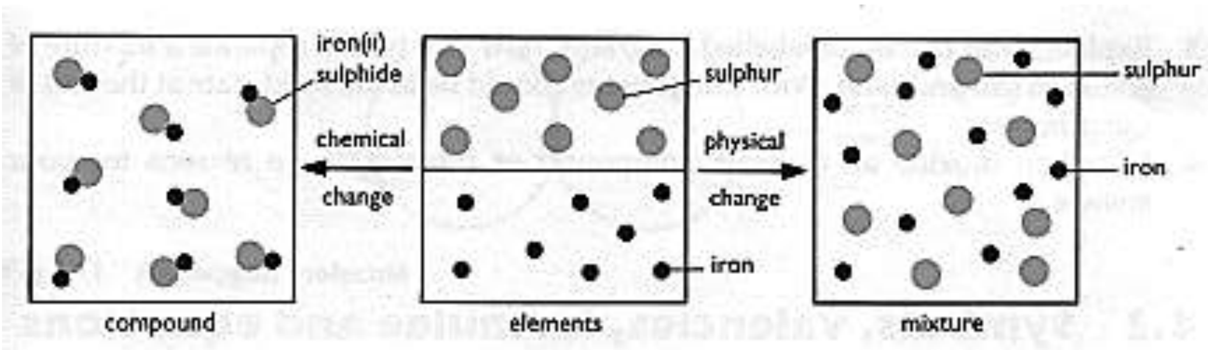
Examples of mixtures include; iron and sulphur together, air, sugar solution and brass. Air is a mixture of oxygen, nitrogen, carbon dioxide and other gases. Sugar solution is a mixture of sugar and water. Brass a mixture of zinc and copper.

COMPOUNDS

The new substance formed when iron chemically combines with sulphur is called iron (II) sulphide.

A **compound** is a substance which consists of two or more elements chemically combined. The properties of a compound are different from those of the individual elements. This is because a chemical change has taken place. In a compound, the particles of the elements are **combined** chemically in a **fixed ratio**. It is often difficult to separate a compound into its constituent elements.

Some examples of common compounds are water, common salt and sugar. Water consists of the elements hydrogen and oxygen, common salt consists of sodium and chlorine and sugar consists of carbon, hydrogen and oxygen.



Differences between mixtures and compounds

Mixture	Compound
The substances in a mixture can be separated by physical means.	The elements in a compound cannot be separated by physical means.
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 e.t.c) 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.
The composition of a mixture is variable. The substances can be present in any proportions by mass.	The composition of a compound is fixed. The elements are combined in definite proportions by mass.

4.3 SYMBOLS, VALENCIES, FORMULAE AND EQUATIONS

ATOMS

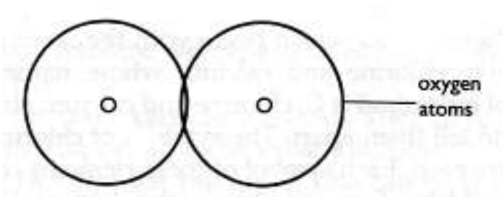
If an element is repeatedly divided, a tiny particle which is difficult to split up would eventually remain. Such a particle is called an atom. An **atom** is the smallest particle of an element that can take part in a chemical change.

MOLECULES

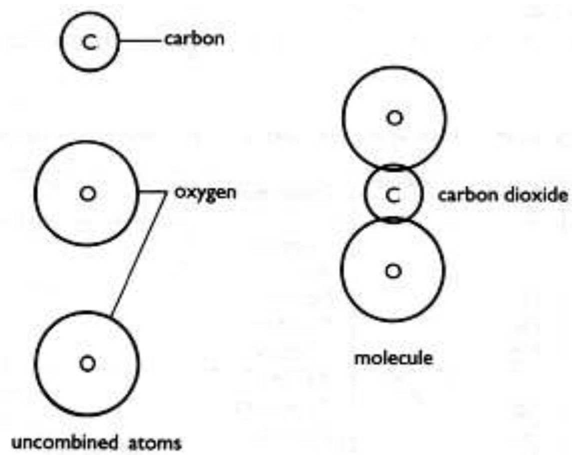
Gases and liquids normally occur as molecules. A **molecule** is the smallest particle of an element or a compound which exists independently, that is in a free state.

Molecules can be thought of to be of three types. The first type consists of those elements in which a single atom forms the molecule. These molecules are called **monoatomic**. Examples of elements which form molecules from single atoms are gases such as helium, neon and argon which are known as the noble gases.

The second type of molecule consists of atoms of the same element combined together. These molecules are called **diatomic.** Single atoms of these elements do not occur in a free state. Examples of elements, each of which has two of its atoms combined to form a molecule, include oxygen, hydrogen, nitrogen and chlorine.



The third type of molecule consists of atoms of different elements combined together. Here the atoms form molecules of compounds. Examples of such molecules include carbon dioxide, water and sugar.



SYMBOLS

A chemical symbol of an element is a letter or letters derived from the name of the element. It represents one atom of the element.

A symbol is, in some cases, the first letter of the English name of the element written in capital form. Examples of such are:

Boron	B	Hydrogen	H	Oxygen	O
Carbon	C	Iodine	I	Phosphorus	P
Fluorine	F	Nitrogen	N	Sulphur	S

Others consist of two letters. Each symbol of these consists of a capital of the first letter and a small letter from the name of the element. Examples are:

Aluminium	Al	Calcium	Ca	Neon	Ne
Argon	Ar	Chlorine	Cl	Silicon	Si
Barium	Ba	Magnesium	Mg	Zinc	Zn
Bromine	Br	Manganese	Mn	Zirconium	Zr

VALENCY (COMBINING POWER)

The names of compounds can only be written in short hand if the combining power of the elements is known. The combining power is known as the valency.

The **valency** of an element may be defined as the number of hydrogen atoms which will combine with or displace one atom of the element. It is usually a small whole number and indicates an element's power of combining with other elements to form compounds. Hydrogen is assigned a standard valency of one, so it then becomes easy to work out the valencies of the other elements.

For instance, since one atom of chlorine will combine with one atom of hydrogen to form the compound hydrogen chloride, the valency of chlorine is one. Similarly, two atoms of hydrogen will combine with one atom will combine with one atom of oxygen to form water and so the valency of oxygen is two.

Some elements, however, have variable valencies. They have different valencies under different conditions. Copper and iron exhibit variable valencies. The valency of copper can be one or two and the valency of iron can be two or three. For this reason, when copper and iron form compounds, the valency of the copper or iron is indicated in brackets in the name of the compound. For instance, copper combines with oxygen to form copper (I) oxide and copper (II) oxide. Iron combines with chlorine to form iron (II) chloride and iron (III) chloride.

Valencies of common 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(II) 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	P

RADICALS

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

All sulphates contain the sulphate radical, SO_4. The sulphate radical has a valency of two. Some compounds containing the sulphate radical are:

Sodium sulphate	Na_2SO_4
Iron (II) sulphate	$FeSO_4$
Zinc sulphate	$ZnSO_4$
Copper (II) sulphate	$CuSO_4$

Radicals and their valencies

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

The ammonium radical, NH_4 forms compounds similar to the compound of potassium and sodium.

CHEMICAL FORMULAE

A chemical formula consists of a symbol or symbols showing the number of atoms in one molecule of an element or a compound. There are basically two types of formula. One represents one molecule of an element and the other represents one molecule of a compound.

Most of the elements that are gases consist of molecules each containing two atoms. The molecules of hydrogen, oxygen, nitrogen and chlorine are all of this type. One molecule of each of these elements is represented by the symbol of the element with, at the bottom right-hand side of the symbol, a small subscript digit showing the number of atoms in the molecule.

HYDROGEN

- H represents one atom of hydrogen.
- H_2 represents one molecule of hydrogen with two atoms combined.
- 2H represents two separate atoms of hydrogen.
- $2H_2$ represents two molecules of hydrogen each containing two atoms.

NITROGEN

- N represents one atom of nitrogen.
- N_2 represents one molecule of nitrogen with two atoms combined.
- 3N represents three atoms of nitrogen.
- $3N_2$ represents three molecules of nitrogen each containing two atoms.

The same system is used to represent a molecule of a **compound.**

MOLECULES OF COMPOUNDS

- H_2O represents one molecule of water in which two atoms of hydrogen are combined with one atom of oxygen.
- $2CO_2$ represents two molecules of carbon dioxide each having one atom of carbon combined with two atoms of oxygen.

- $3CaCO_3$ represents three molecules of calcium carbonate each having one atom of calcium combined with one atom of carbon and three atoms of oxygen.
- $4Na_2SO_4$ represents four molecules of sodium sulphate each having two atoms of sodium combined with one atom of sulphur and four atoms of oxygen,

The large digit in front of the formula is used to show the number of molecules of a compound or element involved in a reaction. This digit is a multiplier for the whole formula and not just for the first element in the formula, for example $4Na_2SO_4$ contains:

Na	4 x 2= 8 atoms
S	4 x 2= 4 atoms
O	4 x 4= 16 atoms

WRITING FORMULAE

Writing correct formulae for substances is simple when the valencies of the elements and radicals are known. Any formula can be written by following **three simple steps.** For example the formulae for calcium chloride and sodium carbonate may be written by following these steps.

1. Write the symbols for the combining elements and radicals.

   Ca Cl Na CO_3

2. Write the valency of each element or radical at its top right-hand side.

   Ca^2 Cl^1 Na^1 CO_3^2

3. Rewrite the symbols and formulae and 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.

Ca^2 Cl^1 becomes $CaCl_2$
Na^1 CO_3^2 becomes Na_2CO_3

Sometimes formulae can be written in the simplest way e.g. in the case of calcium and oxygen which have a valency of two. It would be written as Ca_2O_2 but it has a common factor of two so the formula will be CaO, which is used for calcium oxide.

In some formulae, radicals can be written in brackets followed by a small subscript digit. This is how the formula for calcium hydroxide is written in three simple steps:

1. Ca – OH
2. Ca^2 – OH^1
3. $Ca(OH)_2$

Two hydroxide (–OH) groups are present in calcium hydroxide. To make this clear the –OH group is placed in brackets followed by a small subscript '2' at the bottom right-hand side of the brackets. $Ca(OH)_2$ is the same as CaO_2H_2 as the subscript '2' multiplies everything inside the brackets. $Ca(OH)_2$ shows more clearly than CaO_2H_2 that two hydroxide groups are present in the formula which is why we write it this way.

CHEMICAL EQUATIONS

Chemical changes, which are usually called **chemical reactions**, can be represented by **equations**. An equation is a chemical sentence which describes what is going on in a chemical reaction. The substances that take part in a reaction are called **reactants** and are written on the left-hand side of the equation. The new substances formed by the reaction are called **products** and are written on the right-hand side of the equation.

reactants	**form**	**products**
A + B	→	C + D

The plus sign on the left-hand side of the equation means 'reacts with' and on the right-hand side it means 'and'. The reactants and products are separated by an arrow → which means 'to form'. Of course a reaction takes place only if the conditions, for example temperature, are suitable.

STATE SYMBOLS

For completeness there is a need to show the physical states of the substances involved in a chemical reaction. This is done by placing a small letter or letters in brackets after the name of each substance in the equation.

Solids have (s) **written after them. If a substance is a liquid,** (l) **is put after it, and if it is a gas,** (g) **is used. If a substance is dissolved in water** (aq) **is the symbol used. It stands for 'in aqueous solution'. These letters, used to denote the physical states of substances are known as** state symbols.

WORD EQUATIONS

Equations may be written in words. Such equations are called **word equations**. For instance, the word equation, including state symbols, for a reaction may be written as: 'iron(s) + sulphur(s) → iron(II) sulphide(s)'.

This equation is a short way to describe the reaction 'solid iron reacts with solid sulphur to form solid iron(II) sulphide'. All chemical reactions can be represented by word equations. Word equations can be written for all the reactions referred to so far in this unit, as follows:

- magnesium(s) + oxygen(g) → magnesium oxide(s)
- mercuric oxide(s) → mercury(l) + oxygen(g)
- dilute hydrochloric acid(aq) + iron(s) → iron(III) chloride(aq) + hydrogen(g)
- hydrogen(g) + oxygen(g) → water(g)
- dilute hydrochloric acid(g) + iron(II) sulphide(s) → iron(II) chloride(aq) + hydrogen sulphide(g) + sulphur(s) + water(g)
- hydrogen sulphide(g) + oxygen(g) → sulphur(s) + water(g)

EQUATIONS USING SYMBOLS

Word equations such as these are not complete because they do not show the actual numbers of atoms involved in a chemical reaction. If we were to write the

numbers of atoms in words the word equations would become rather long so instead of using words, chemical symbols and formulae may be used. This is a more convenient and shorter way to describe a chemical reaction and it is also more complete as it shows the actual numbers of atoms involved in the chemical reaction.
The reaction ‘hydrogen burns in oxygen to produce water vapour’ may be written using symbols and formulae:

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

This equation may be illustrated in a diagram to show clearly the numbers of hydrogen and oxygen atoms on both sides of the equation:

$$H{-}H + O{-}O \rightarrow H{-}O{-}H$$

Both the equation and the diagram may **appear** to be a complete way of describing the reaction, but this is not so. Let us look at the numbers of the hydrogen and oxygen atoms on the left-hand and right-hand sides of the equation:

	left	**right**
number of hydrogen atoms	2	2
number of oxygen atoms	2	1

Atoms cannot be created or destroyed in a chemical reaction so all the atoms present at the beginning of a reaction must still be there at the end. No new atoms can appear and none can disappear.

This means that the number of each kind of atom in a chemical reaction must be the same on the left and on the right-hand sides of the equation. So our equation is unbalanced and we must think again.

BALANCING EQUATIONS

The number of hydrogen atoms should equal on both sides of the equation. Similarly, the number of oxygen atoms should be equal on both sides of the equation. This is achieved by writing **whole** number multiples in front of the symbols and formulae as fractions of atoms and molecules do not exist. The complete equation for the reaction between oxygen and hydrogen becomes:

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

The equation has **four** hydrogen and **two** oxygen atoms on the left-hand side and on the right-hand side.
This process of making the number of each type of atom equal on both sides of the equation is called **balancing** the equation.
The balanced equation may also be illustrated as shown below:

H–H H–O–H
+ O–O →
H–H H–O–H

The numbers of hydrogen and oxygen atoms on the left-hand and right-hand sides of the equation can again be compared with ease, as follows:

	left	**right**
number of hydrogen atoms	4	4
number of oxygen atoms	2	2

The following steps are used to write chemical equations for example the reaction between magnesium and oxygen to produce magnesium oxide.

1.Write the correct symbols and formulae for the reactions and products.

$$Mg + O_2 \rightarrow MgO$$

2. Count the numbers of each type of atom on both sides of the equation and then write the **smallest** whole number multiple in front of the symbols and formulae in order to balance the equation. Since we have two oxygen atoms on the left-hand side and need two on the right-hand side, we must have two magnesium atoms and two molecules of magnesium oxide.

$$2Mg + O_2 \rightarrow 2MgO$$

3. Check that the number of each type of atom is equal on the left and right-hand sides of the equation.

	left	**right**
Mg	2	2
O	2	2

4. To complete the picture, indicate the state of each substance in the equation.

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

ATOMIC STRUCTURE

5.1 THE COMPOSITION AND CHARACTERISTICS OF ATOMS

PARTICLES IN AN ATOM

Atoms are made up of three fundamental particles. These are the **proton,** the **neutron** and the **electron**.

An atom has a central body, the **nucleus** which is made up of protons and neutrons.

Electrons revolve around the nucleus like planets around the sun. The paths taken by electrons are referred to as **shells**.

Each of the three fundamental particles of an atom differs from the others by possessing certain characteristics in terms of **mass** and **electrical charge.**

The **proton** has a unit positive electrical charge and a mass of 1. The **neutron** has no charge and its mass is almost equal to that of a proton such that it is also taken as being 1. The **electron** has a unit negative charge and has approximately 1840 electrons have the same mass as one proton, electrons are said to have negligible mass.

Characteristics of atomic particles

Particle	Mass (a.m.u.)	Charge
Proton	1	+1
Neutron	1	Neutral
Electron	$\frac{1}{1840}$	-1

The value of 1 for the mass of a proton or neutron is an arbitrary unit known as the **atomic mass unit** (a.m.u.). It represents the **relative** masses of the proton, neutron and electron and therefore is a ratio. It has no unit.

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

ATOMIC NUMBER AND MASS NUMBER

The atoms of each element have a characteristic number of protons in their nucleus. The number of protons in an atom is called the **atomic number (Z)** of the element.

Since atoms are electrically neutral, the atomic number is also equal to the number of electrons around the nucleus of an atom:

atomic number= number of protons= number of electrons

It is also very useful to remember that protons and neutrons have almost equal masses. Protons and neutrons make up most of the mass of an atom, which means the greater part of the mass of an atom is concentrated in the nucleus.

The greater the number of protons and neutrons in an atom, the greater the mass of the atom. The total number of protons and neutrons in an atom is called the **mass number (A)** of the atom.

If the number of neutrons is represented as N, the mass number of an atom can be calculated using the formula:

$$A = Z + N$$

or:

mass number = number of protons + number of neutrons

EXAMPLE

Sodium has atomic number 11. If sodium has 12 neutrons, calculate the mass number of sodium.

$$A = Z + N$$
$$= 11 + 12$$
$$= 23$$

The mass number of sodium is 23.
The atomic numbers and mass numbers of the first 20 elements are shown below.

Atomic and mass numbers

Element	Atomic number (Z)	Mass number (A)
Hydrogen	1	1
Helium	2	4
Lithium	3	7
Beryllium	4	9
Boron	5	11
Carbon	6	12
Nitrogen	7	14
Oxygen	8	16
Fluorine	9	19
Neon	10	20
Sodium	11	23
Magnesium	12	24
Aluminium	13	27
Silicon	14	28
Phosphorus	15	31
Sulphur	16	32
Chlorine	17	35
Argon	18	40
Potassium	19	39
Calcium	20	40

Nuclides

The atomic number and mass number of an atom are sometimes written together with the symbol of the element. This represents a nuclide.

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. Thus a nuclide is represented as ${}^{A}_{Z}X$ where X is the symbol of the element.

For instance, the sodium and magnesium nuclides may be written as:

$${}^{23}_{11}Na \quad \text{and} \quad {}^{24}_{12}Mg$$

respectively.

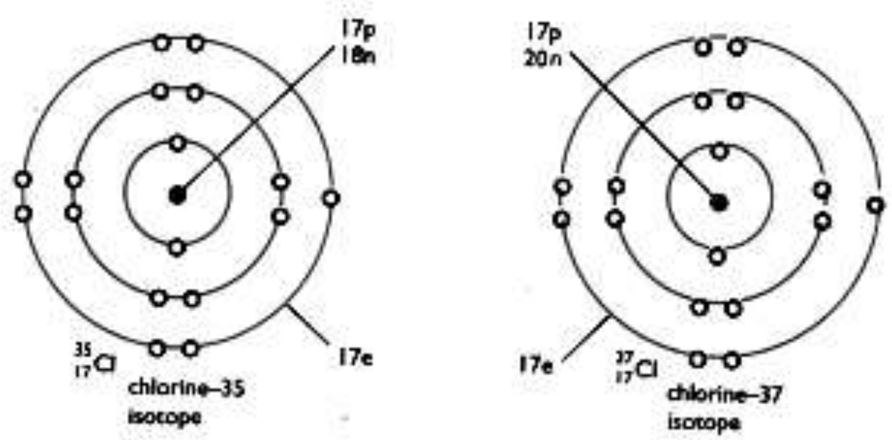
ISOTOPES

These are atoms of the same element having the same number of protons but different numbers of neutrons. Hence isotopes have the same atomic number but different mass numbers.

The mass number of an isotope is usually quoted to distinguish it from other isotopes. The two isotopes of chlorine-35, which has 18 neutrons and chlorine-37, which has 20 neutrons.

Many elements occur with several isotopes. Carbon has three isotopes namely, carbon-12, carbon-13 and carbon-14. The carbon-12 isotope has six protons and six neutrons in its nucleus, the carbon-13 has six protons and seven neutrons in its nucleus and the carbon-14 isotope has six protons and eight neutrons in its nucleus.

Three isotopes of hydrogen are known to exist. These are hydrogen-1, hydrogen-2 and hydrogen-3. The hydrogen-1 isotope, also called **ordinary hydrogen**, has one proton and no neutrons. The hydrogen-2 isotope, also called **deuterium**, has one proton and one neutron in its nucleus. Deuterium combines with oxygen to form a compound called heavy water. The third isotope of hydrogen, hydrogen-3, which is also called **tritium**, has one proton and three neutrons in its nucleus.



USES OF ISOTOPES

Isotopes have a wide range of uses. Hydrogen isotopes are important in the study of nuclear energy. Some isotopes are artificially made by firing neutrons or protons at elements.

These usually emit radioactive rays and are called **radioisotopes.** They are used as tracers to investigate such things as the flow of liquids in chemical plants and for the study of wear of machines.

THE ARRANGEMENT OF ELECTRONS IN AN ATOM

Electrons are arranged in a definite pattern in an atom. They lie along rings or shells around the nucleus. The shells are assigned the letters K, L, M, and so on.

Forces of attraction exist between the electrically charged protons and electrons. These are called **electrostatic** forces and they prevent the electrons from drifting away from the rest of the atom.

The electrostatic forces of attraction in an atom increase as the distance between the nucleus and electrons decreases. Those nearest to the nucleus are held strongly and are very stable so more energy is required to pull these electrons away from the nucleus.

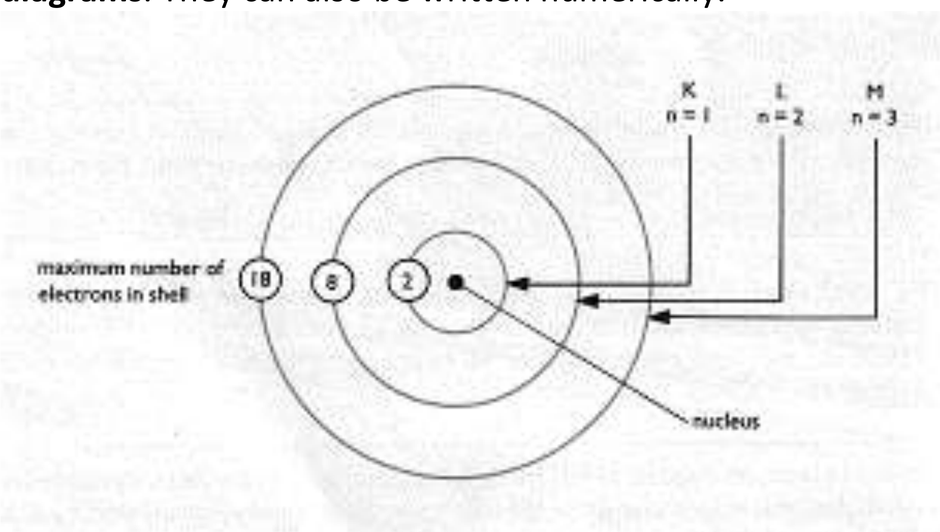
Shells of electrons have different levels of energy. So these electron shells are called **energy levels.**

Generally, the maximum number of electrons that can fill a shell is given by the formula $\mathbf{2n^2}$ where n is the number of the shell.

Numbers of electrons in shells K, L and M

Letter of shell	K	L	M
Number of shell	1	2	3
$\mathbf{2n^2}$	2×1^2	2×2^2	2×3^2
Maximum number of electrons	2	8	18

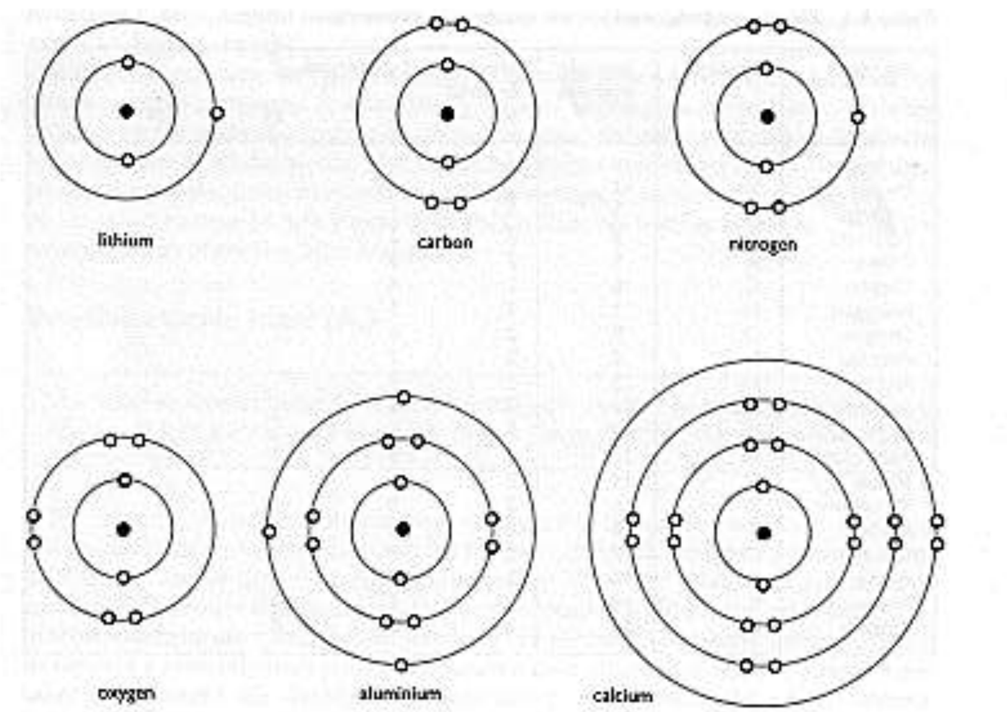
An easier way to draw electron shells is by drawing **electron configuration diagrams**. They can also be written numerically.



STRUCTURES OF NOBLE GASES

The elements helium, neon and argon are called **noble gases**. They have one similarity which is that each atom of these elements has a completely filled outermost shell.

Noble gases are highly **unreactive** as a result of this stability.



Electronic configuration for the first 20 elements

Element	Symbol	Atomic number	E.C
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
PHOSPHORUS	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

5.2 RELATIVE ATOMIC AND MOLECULAR MASSES

RELATIVE ATOMIC MASS (A_R)

The **relative atomic mass** A_r of an element is the number of times that the mass of one atom of the element is greater than $\frac{1}{12}$ the mass of an atom of carbon-12.

For instance, the relative atomic mass of oxygen is 16 and this means that oxygen has a mass 16 times heavier than $\frac{1}{12}$ of the mass of carbon. Notice that as the relative atomic mass is a **ratio** it has no units. Most elements are composed of two or more naturally-occurring isotopes. Some isotopes of a given element are more abundant than others.

For example, chlorine contains approximately 75 per cent of an isotope with mass number 35 and approximately 25 per cent of an isotope with mass number 37. Consequently the relative atomic mass of a sample of chlorine is:

$$\frac{75}{100} \text{ x } 35 + \frac{25}{100} \text{ x } 37 = 26.25 + 9.25 = 35.5$$

So the relative atomic mass of chlorine given in tables is 35.5. Clearly, this value is closer to 35, the mass number ofchlorine-35, which is the more abundant isotope.

Similarly, the relative atomic mass of naturally-occurring carbon which is composed mainly (98.89 per cent) of carbon-12 and 1.11 per cent of carbon-13 is:

$$\frac{98.89}{100} \text{ x } 12 + \frac{1.11}{100} \text{ x } 13 = 12.01$$

RELATIVE MOLECULAR MASS (M_R)

The relative molecular mass (M_r) of an element or compound is the mass of one molecule of that element or compound compared with $\frac{1}{12}$ of the mass of an atom of carbon-12.

To find the relative molecular mass, the relative atomic masses of all the atoms in the formula are added up.

Calculating relative molecular mass

Name of compound	Formula	Relative molecular mass (M_r)
Hydrogen chloride	HCl	1 x H = 1x 1 = 1 1 x Cl = 1 x 35.5 = 35.5 **36.5**
Water (hydrogen oxide)	H_2O	2 x H = 2 x 1 = 2 1 x O = 1 x 16 = 16 **18**
Carbon dioxide	CO_2	1x C = 1 x 12 = 12 2 x O = 2 x 16 = 32 **44**
Sodium carbonate	Na_2CO_3	2 x Na = 2 x 23 = 46 1 x C = 1 x 12 = 12 3 x O = 3 x 16 = 48 **106**
Aluminium sulphate	$Al_2(SO_4)_3$	2 x Al = 2 x 27 = 54 3 x S = 3 x 32 = 96 12 x O = 12 x 16 = 192 **342**

6. BONDING AND STRUCTURE OF MATTER

6.1 BONDING OF ATOMS

Atoms combine with each other by forming bonds and this combination is known as **bonding.**

WHY DO ATOMS FORM BONDS?

Atoms of some elements have incompletely filled outermost shells, and as a result, these atoms are **unstable** and therefore, **reactive**. They react with one another in order to acquire full outermost shells like those of the noble gases.

Atoms of elements can acquire a full outermost shell by losing or gaining electrons and also by sharing electrons when they react with each other. When an atom loses or gains electrons it becomes an **ion.**

ELECTRONS AND VALENCY OF ELEMENTS

The number of **electrons** in the outermost shell of an atom is closely related to the **valency** of an element.

For example, the electronic configuration of the element sodium (Na) is 2,8,1. The valency of this element is 1, because it has only one electron in the outermost shell. Magnesium (Mg) has the electronic configuration 2,8,2 and its valency is 2.

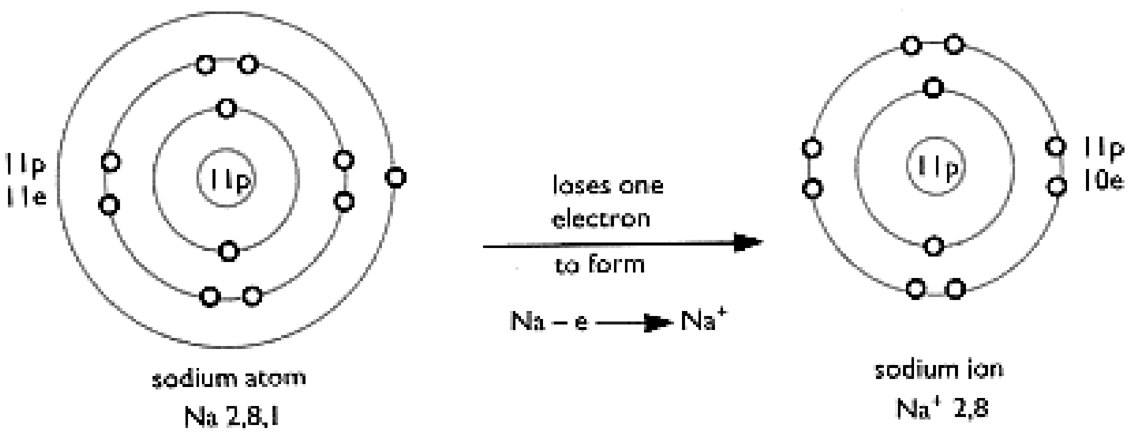
Sometimes the valency of an element may be equal to the maximum number in a shell minus the number of the outermost electrons. For example, Oxygen has six electrons in its outermost shell (2,6). It has a valency of (8-6) = 2.

The **valency** of an element determines the number of bonds that an element can form.

6.2 IONIC BONDING

STRUCTURE OF SODIUM CHLORIDE

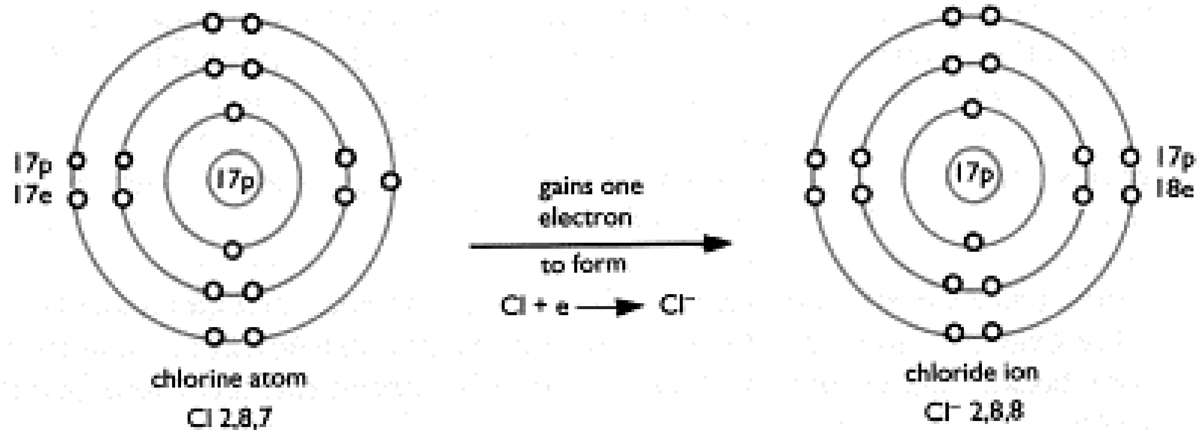
Sodium has the electronic configuration 2, 8, 1. It has one electron in its outermost shell. It can attain a full outermost shell by losing this single electron to another atom. When this happens, a sodium **ion** is formed.



The ion is positively charged and has the symbol Na^+. The formation of the sodium ion can be expressed as the **ionic half reaction:**

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

An example showing a gain of electrons is chlorine. This element has the electronic configuration 2, 8, 7.It can attain a full outermost shell by gaining one electron from another atom. When this happen it becomes a chloride ion. It is negatively charged because it has one more electron than proton. Its symbol is Cl^-



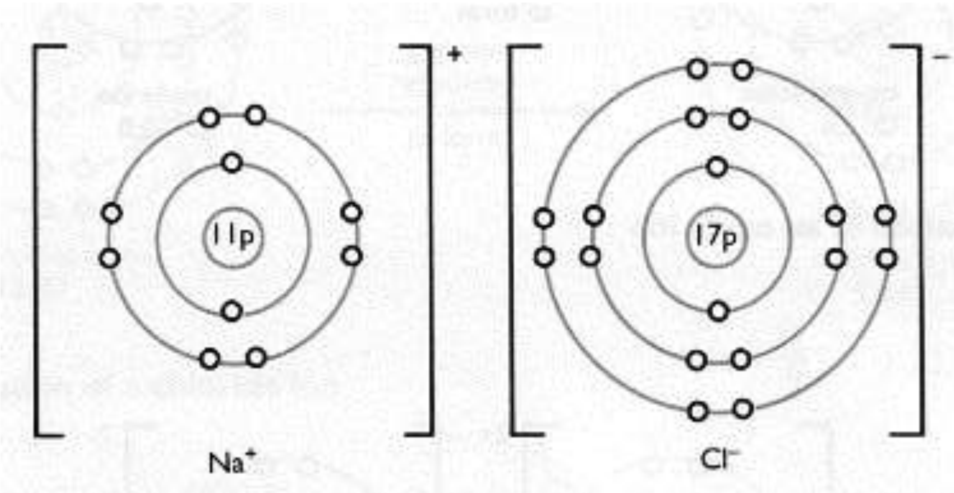
Na^+ and Cl^- are charged particles because they have unequal numbers of protons and electrons. They are now ions and are formed by lose or gain of electrons.

The sodium ion (Na^+) and the chloride ion (Cl^-) are unreactive because they have full outermost shells. But because they are charged, an electrostatic force of attraction is created.

This results in an **electrovalent** or **ionic bond** between the two oppositely charged ions. An electrovalent bond, is therefore, formed by the transfer of electrons. It takes place between **metals and non-metals**. The metal loses electrons while the non-metal gains electrons.

When sodium and chlorine combine, they form NaCl. There is one sodium ion for each chloride ion hence the compound is neutral. This reaction can be expressed by the **ionic equation**:

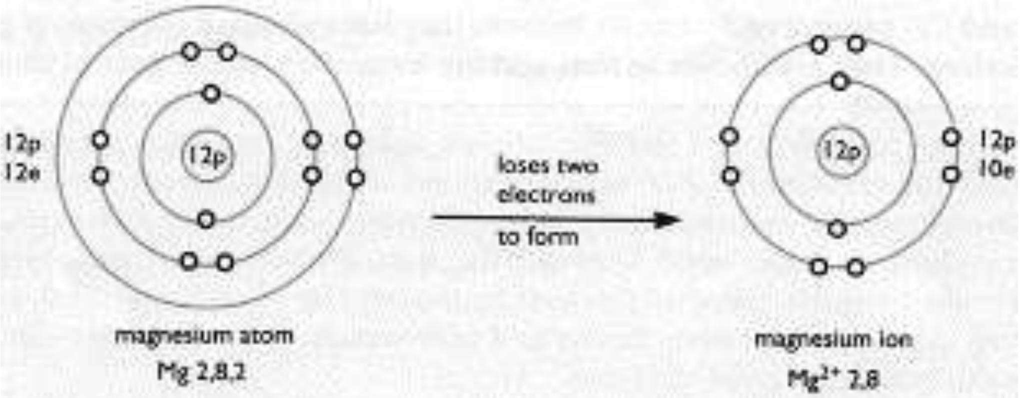
$$Na^+ + Cl^- \rightarrow NaCl$$



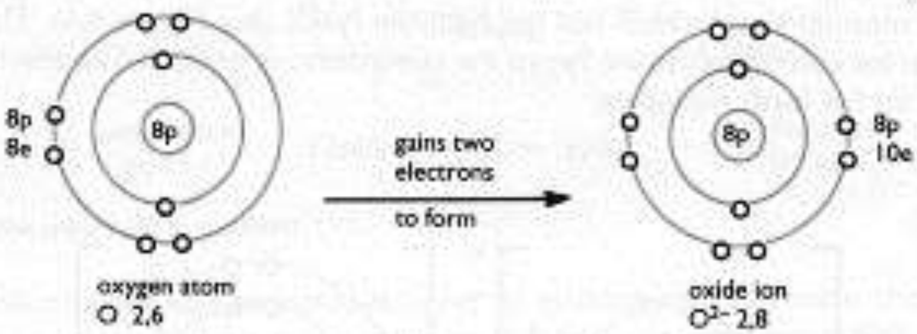
STRUCTURE OF MAGNESIUM OXIDE

During reaction, magnesium loses its two outermost electrons to form a magnesium ion with a charge of +2. Oxygen gains the two electrons lost by magnesium to fill its outermost shell. An oxide ion with a charge of -2 is formed.

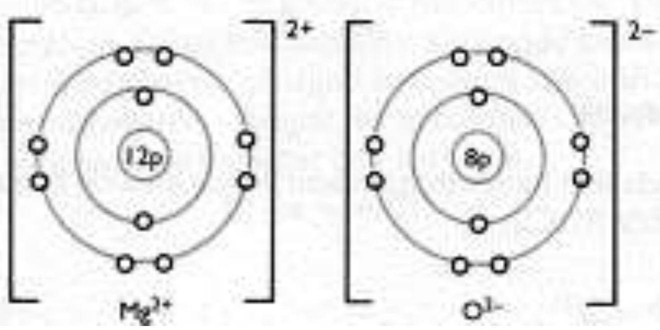
The oppositely charged ions attract each other and form a neutral compound, magnesium oxide. This compound has the formula MgO.



Formation of a magnesium ion



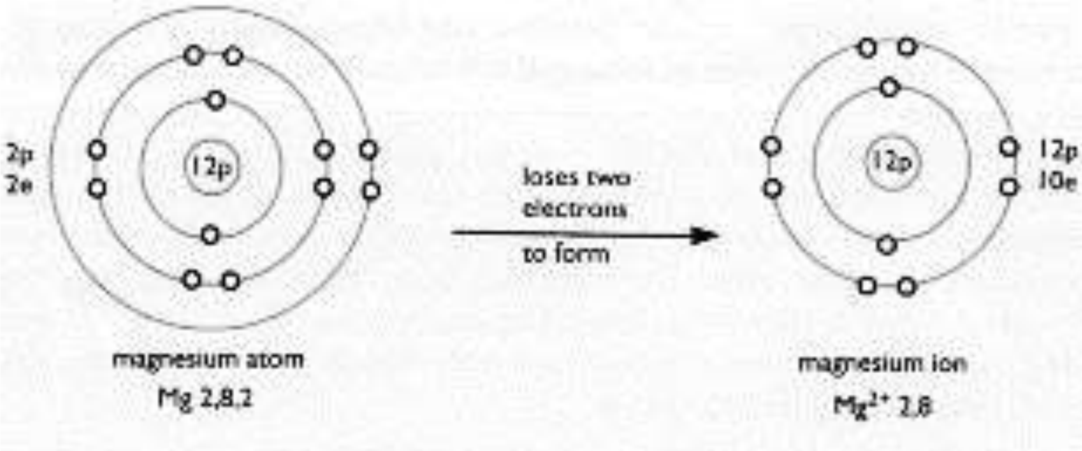
Formation of an oxide ion



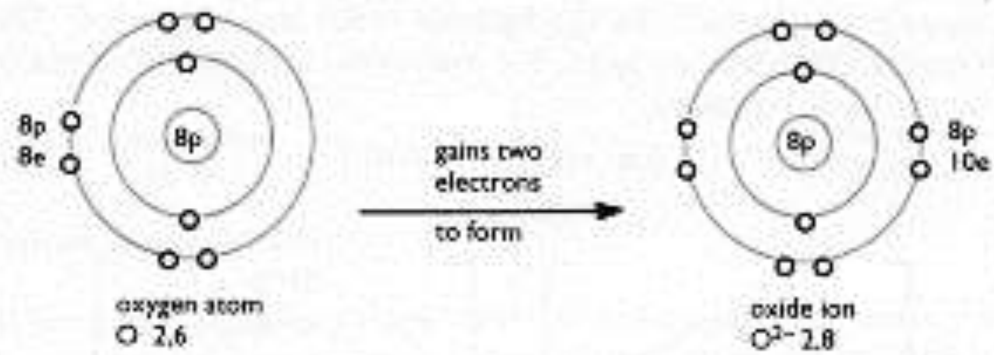
Magnesium oxide.

STRUCTURE OF CALCIUM CHLORIDE

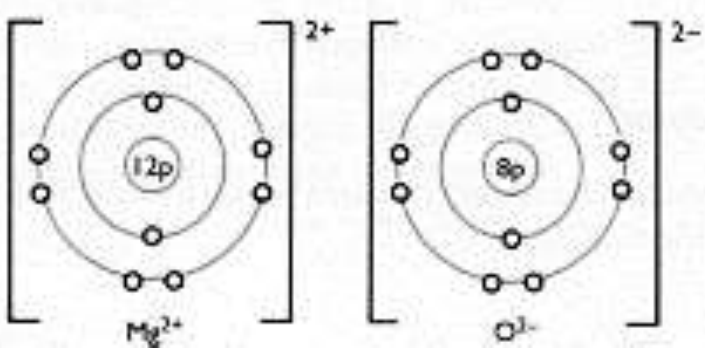
Calcium loses two electrons in the outermost shell. These two electrons are gained by two chlorine atoms because each chlorine atom can gain only one electron to form a chloride ion. Each calcium atom reacts with two chlorine atoms to form calcium chloride, $CaCl_2$.



Formation of a magnesium ion



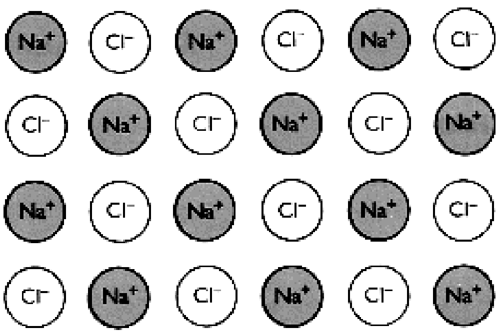
Formation of an oxide ion

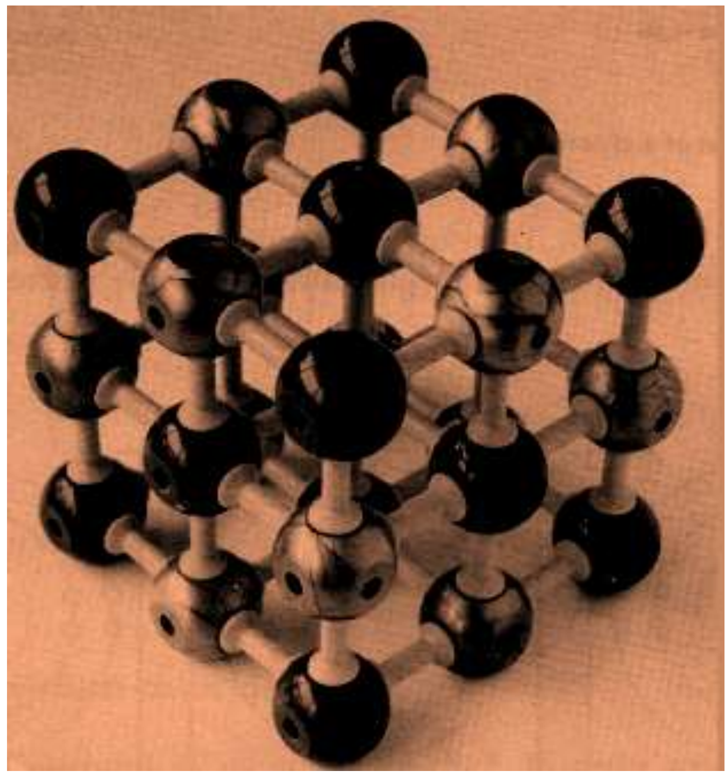


Calcium chloride

IONIC LATTICES

Sodium chloride, magnesium oxide and calcium chloride are examples of ionic compounds. These have giant ionic structures or lattices. In the sodium chloride lattice, each sodium is surrounded by six chloride ions while each chloride ion is also surrounded by six sodium ions.





CHARACTERISTICS OF IONIC/ELECTROVALENT COMPOUNDS

- They are made up of positively charged and negatively charged ions.
- In aqueous solution and in molten state, they conduct electricity.
- They are soluble in water but are insoluble in organic solvents such as ethanol.
- They are non-volatile and are generally solids at room temperature.
- They have high melting and boiling points.
- The high melting and boiling points are due to the strong electrostatic forces of attraction between the ions. Breaking such strong bonds requires great energy. Because of these high boiling and melting points some of these

compounds are used in industry to line furnaces, for example calcium oxide, which has a melting point of 2614°C.

Melting points and boiling points of some ionic compounds

Compound	Melting point °C	Boiling point °C
Sodium iodide	647	1227
Magnesium chloride	714	1407
Sodium chloride	801	1467
Lithium fluoride	867	1667

6.3 COVALENT BONDING

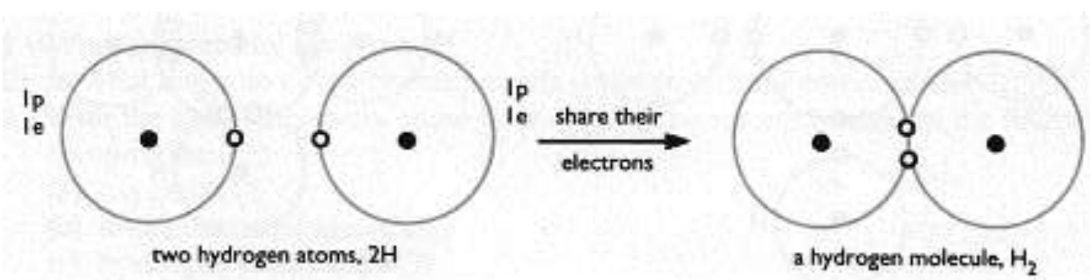
This is the type of bond formed by the sharing of electrons. It is also known as a **molecular bond**. It consists of a shared pair of electrons and is formed between **non-metallic** elements which share one or more pairs of electrons.

The **covalency** of an atom is the number of electrons it shares with other atoms.

SOME COVALENT MOLECULES AND HOW THEY ARE FORMED

HYDROGEN

A hydrogen atom has only one electron. When two hydrogen atoms combine, they share electrons. It has the formula H_2.

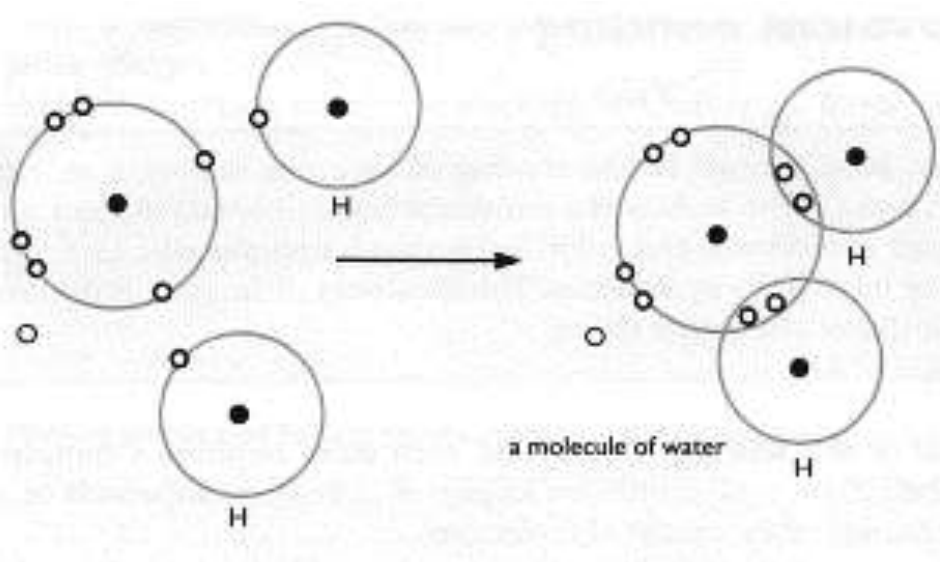


The shared pair of electrons in a hydrogen molecule can also be shown as:

H:H or **H–H**

WATER

Water has the formula H_2O. In each molecule, an oxygen atom shares electrons with two hydrogen atoms so that all atoms obtain full outermost shells.

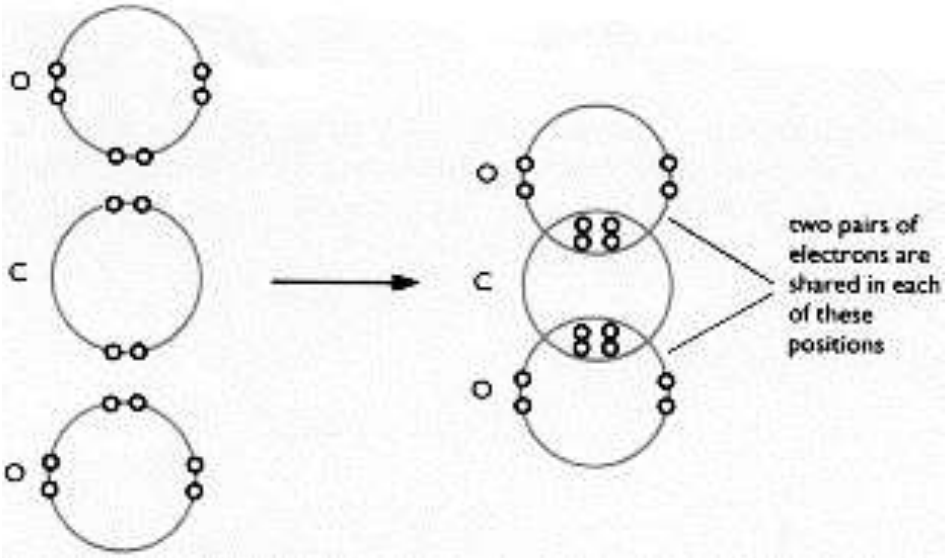


A molecule of water can also be shown as:

H : O : H or **H–O–H**

CARBON DIOXIDE

Carbon atom acquires a full outermost shell by sharing electrons with two oxygen atoms. Since each covalent bond consists of two pairs of electrons, a **double** covalent bond is formed.



A molecule of carbon dioxide can also be shown as:

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

Two pairs of electrons are shared between a carbon atom and each oxygen atom.

CHARACTERISTICS OF COVALENT COMPOUNDS

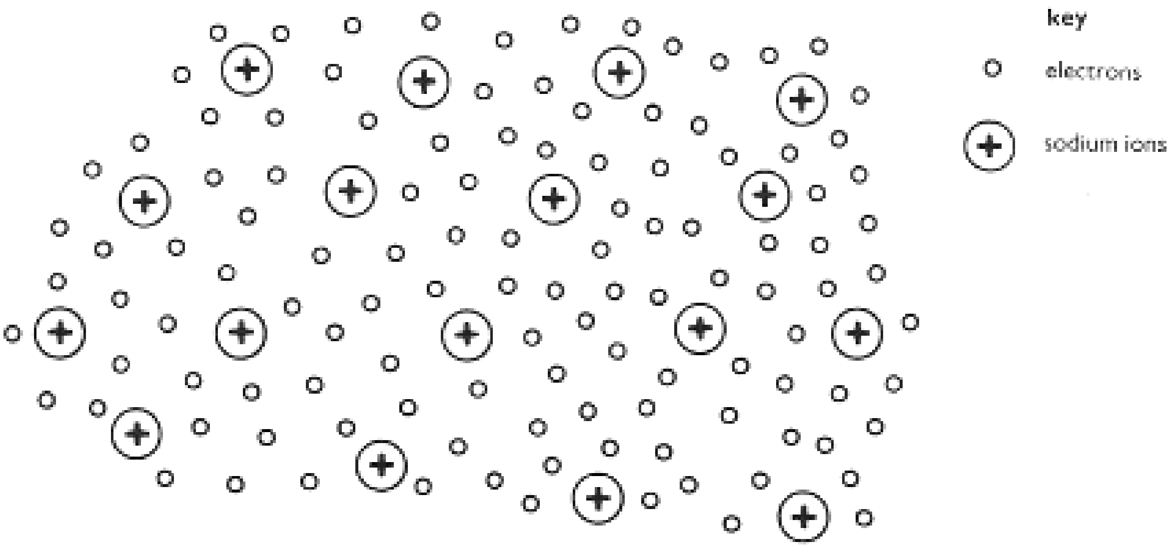
- They are made up of molecules.
- They have low melting and boiling points.
- The low melting and boiling points are due to the fact that the forces holding the molecules together are weak. Hence not much energy is required to separate them.
- They are usually insoluble in water but are soluble in organic solvents such as petrol and ethanol.
- They do not conduct electricity in solid or molten state.
- They are volatile.

Melting and boiling points of some covalent compounds

Compound	Melting point °C	Boiling point °C
Oxygen	-218	-183
Chlorine	-101	-35
Water	0	100
Naphthalene	80	218

6.4 METALLIC BONDING

Atoms of metals acquire a full outermost shell by losing electrons. Sodium atoms, for example, lose a single electron from the outermost shell. When a large number of sodium atoms lose these electrons, the result is many free electrons.



The sodium ions which are positively charged, repel each other but they are held together by the negatively charged free electrons between them. The attractive forces between the electrons and the sodium ions are greater than the repulsive forces between the sodium ions.

The electrons are free to move anywhere in the metallic lattice. They are said to be **delocalised**. The attraction between the electrons and the positively charged ions is known as **metallic bonding**.

Note: The presence of free electrons makes metals good conductors of electricity.

EXAM TYPE QUESTIONS WITH ANSWERS III

Attempt the following questions before looking at the answers.

1. (a) Magnesium is an element. It has an atomic number of 12 and the main isotope of magnesium has a mass number of 24.

(i) Define the term "atomic number"

____ [1]

(ii) What are isotopes?

____[2]

(iii) How many neutrons are present in the nucleus of an atom of magnesium?

____[1]

(b) Graphite, like calcium oxide, has a high melting point yet its structure is very different from calcium oxide. Describe the bonding in graphite.

____ [3]

(c) (i) Barium is a typical Group II element. Complete the table below to show the formulae of the compounds named. [3]

Name Of Compound	**Formula of Compound**
Barium Oxide	
Barium Chloride	
Barium Sulphate	

(ii) Using your knowledge of the Periodic Table, predict whether barium will react more vigorously or less vigorously than calcium. Explain your answer.

____ [2]

(iii) Write a balanced, symbol equation to show how barium reacts with cold water.

____ [2]

d) Carbon and silicon are elements in group IV of the Periodic Table. They have similar properties and form compounds with similar formulae.

Methane and silane are compounds of the group IV elements with hydrogen.

(i) Suggest a formula for silane.

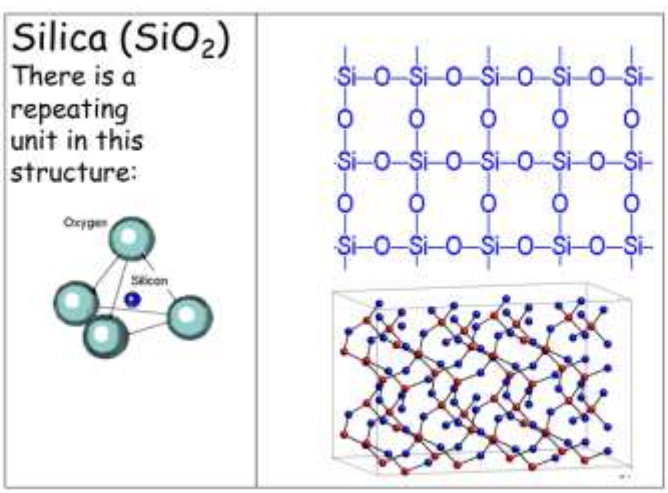
____ [1]

(ii) Draw a diagram to show how atoms of carbon and hydrogen bond to form a molecule of methane. Only outer shell electrons need to be shown.[4]

(iii) Describe the structure of quartz.

____ [2]

(iv) Describe fully why quartz has a high melting point.[3]



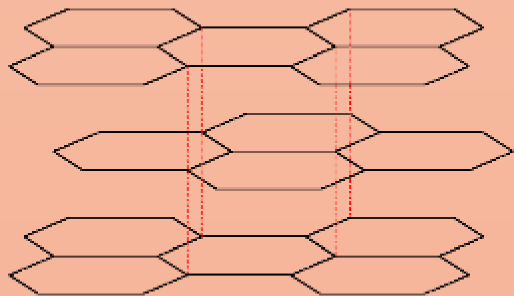
ANSWERS

1.(a)(i)Number of protons (in the nucleus)

(ii) Atoms of the same element/same atomic number of protons [1] with different mass numbers/different numbers of neutrons [1]

(iii) 12

(b) Layered structure [1] each carbon covalently bonded [1] to three [1] other carbon atoms weak forces between the layers [1] Max [3] / [4].



For a diagram only the marks can be awarded for two or more layers [1] two or more hexagons [1] labelled weak bonds drawn between the layers

(c)(i)

Name Of Compound	Formula of Compound
Barium Oxide	BaO
Barium Chloride	$BaCl_2$
Barium Sulphate	$BaSO_4$

(ii) Barium will react more vigorously [1] (than calcium). Barium is below Ca in Group II/the reactivity increases as the group is descended [1]

(iii) $Ba + 2H_2O \rightarrow Ba(OH)_2 + H_2$ [2]

(d)(i) SiH_4

(ii) 1 C and 4 H atoms [1]

4 electrons in outer shell of C atom [1]

1 electron in outer shell of each H atom [1]

2 electrons in each bond [1]

(iii) Each silicon atom bonded to 4 oxygen atoms [1] each oxygen atom bonded to 2 silicon atoms [1] (Max [2])

(iv) Large amount of energy [1] needed to break [1] strong bonds[1]

7. ACIDS, BASES AND ALKALIS

7.1 ACIDS

An **acid** is a substance which, when dissolved in water, produces hydrogen ions (H^+) as the only positively charged ions. It is the hydrogen ion that gives an acid its properties. They have a sour taste.

For example, when hydrogen chloride is dissolved in water, it produces hydrogen ions (positively charged) and chloride ions (negatively charged). This can be shown by:

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

The aqueous solution of hydrogen chloride in this equation, is called hydrochloric acid. It takes the form of ions. The process in which ions are formed is referred to as **ionisation**.

Common acids

Name of acid	Formula
Nitric acid	HNO_3
Sulphuric acid	H_2SO_4
Phosphoric (V) acid	H_3PO_4
Ethanoic acid	CH_3COOH
Carbonic acid	H_2CO_3

Each acid in the table ionises in aqueous solution producing hydrogen ions and negatively charged ions

THE EFFECT OF DILUTE ACIDS ON INDICATORS

Blue litmus paper and methyl orange, turn red in dilute acids, bromothymol blue turns yellow whereas phenolphthalein becomes colourless. Substances such as these, which show particular colour change in an acidic or alkaline medium are called **indicators.**

The indicators show characteristic colour changes, in solutions of different acids. The colours described are the characteristic colours of the indicators in **dilute** solutions of acids. Blue litmus paper turns red in ethanoic acid.

Acids found in some common substances

Name of substance	Acid contained
Orange juice	Citric acid
Lemon juice	Citric acid
Lime juice	Citric acid
Grape juice	Tartaric acid
Sour milk	Lactic acid
Rancid butter	Butanoic acid
Vinegar	Ethanoic acid
Olive oil	Oleic acid

STRONG AND WEAK ACIDS

A **strong** acid is one which ionises completely in aqueous solution. In this case ionisation is complete:

$$HCl(aq) \rightarrow H^{+}(aq) + Cl^{-}(aq)$$

A **weak** acid, on the other hand is one which is partially ionised in aqueous solution. In such a case the ionisation is reversible and this is indicated by the sign $\rightleftharpoons$

The sign in the equation below means an aqueous solution of ethanoic acid consists of molecules and ions at the same time. That is ethanoic acid is a weak acid.

$$CH_3COOH(aq) \rightleftharpoons H^{+}(aq) + CH_2COO^{-}(aq)$$

Other examples of weak acids are carbonic acid (H_2CO_3), phosphoric (v) acid (H_3PO_4) and organic acids in general.

BASICITY OF AN ACID

The **basicity** of an acid is the number of hydrogen ions that can be formed from one molecule of that acid. The basicity of sulphuric acid is two since two hydrogen ions are formed when one molecule of sulphuric acid ionises in solution.

Basicities of common acids

Acid	Formula	Basicity
Hydrochloric acid	HCl	1
Nitric acid	HNO_3	1
Ethanoic acid	CH_3COOH	1
Sulphuric acid	H_2SO_4 HNO_3	2
Carbonic acid	H_2CO_3	2
Phosphoric (V) acid	H_3PO_4	3

TYPES OF ACIDS

There are two classes. These are mineral acids and organic acids.
Mineral acids are prepared from minerals obtained from the earth. Some examples include hydrochloric, sulphuric and nitric acid.

Organic acids are prepared from organic sources such as plants. Vinegar contains ethanoic acid and this gives vinegar its characteristic smell and taste. It is used as an antiseptic and for flavouring and preserving foods.

DETERMINE THE EFFECT OF DILUTE ACIDS ON METALS

When zinc or magnesium metal was added to the acids, there was a vigorous reaction and a gas was given off. On testing with a burning splint, a ‘pop’ sound was produced indicating that the gas given off was hydrogen.

When copper metal was added to the acids, however, there was no reaction and no gas was given off. As a result no ‘pop’ sound was produced when a burning splint was brought near the mouth of the test tube.

From these results, it can be seen that some but not all metals displace hydrogen from dilute acids, for example, copper metal does not. It depends on the reactivity of the metals.

The reactive metals for example sodium, magnesium, aluminium and zinc displace hydrogen from acids. The less reactive metals such as copper, mercury and silver do not displace hydrogen from acids.

The reaction which takes place between zinc metal and dilute hydrochloric acid can be represented by the following equation:

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

This reaction can be rewritten, as follows, to show the ions present in solution and how they interact:

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

If we leave out ions that appear on both sides of the equation, we remain with the following equation:

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

This is called the **ionic** equation of the reaction. An ionic equation of a reaction is very useful because it shows only the particles taking part in the reaction. Those ions which do not take part in the reaction (the Cl^- ions in this case) are called **spectator ions.**

The reaction between magnesium metal and dilute hydrochloric acid can be written as follows:

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

The ionic equation for this reaction is:

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

EFFECTS OF ACIDS ON CARBONATES AND HYDROGEN CARBONATES

The addition of dilute hydrochloric acid to calcium carbonate or sodium hydrogen carbonate causes **effervescence** to take place. This is the name of the effect when a gas is produced very fast with the formation of many bubbles. The gas produced is carbon dioxide which extinguishes a lighted splint and turns lime water milky.

The reactions taking place can be represented by the following equations:

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

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

Ionically these can be written as:

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

$$NaHCO_3(s) + H^+(aq) \rightarrow Na^+(aq) + CO_2(g) + H_2O(l)$$

Any carbonate or hydrogen carbonate will give a similar reaction with a dilute acid, which is they will produce carbon dioxide. This reaction is used as a test for a carbonate or a hydrogen carbonate. It is also a common method for preparing carbon dioxide in the laboratory.

7.2 BASES AND ALKALIS

Sodium hydroxide is used in industry for making soap and plastics. It is also used to remove fat and oil from ovens in our homes. Calcium hydroxide can be used to purify sugar in industry.

A **base** is a compound which consists of oxide (O^{2-}) or hydroxide (OH^-) ions. A base reacts with an acid to form a compound, called a salt, and water only. Bases are oxides and hydroxides of metals. An **alkali** is a compound which, when dissolved in water, forms hydroxide ions (OH^-) as the only negatively charged ions.

CLASSIFYING SUBSTANCES AS BASES AND ALKALIS

The elements display different properties on being held in a bunsen flame. Magnesium and calcium burn easily in air to form their respective oxides.

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

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

Zinc, iron, lead and copper do not burn. (However, it must be noted that they do form oxides but by other methods.)

The oxides of magnesium and calcium dissolve slightly in water to form their respective hydroxides.

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

These solutions feel soapy and turn red litmus blue. The hydroxides of zinc, iron, lead and copper also exist, but are insoluble in water.
All metallic oxides and hydroxides are bases because of the ions they contain. However, some bases are soluble in water whilst others are insoluble. **Soluble** bases are called **alkalis.**
The common alkalis and their ionisation in aqueous solution are:

Potassium hydroxide $KOH(aq) \rightarrow K^+(aq) + OH^-(aq)$
Sodium hydroxide $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$
Calcium hydroxide $Ca(OH)_2(aq) \rightarrow Ca^{2+}(aq) + OH^-(aq)$
Ammonia solution $NH_4OH(aq) \rightarrow NH_4^+(aq) + OH^-(aq)$

Sodium and potassium hydroxide are **corrosive** substances.

Relative solubilities of common bases

Base	Formula	Relative solubility
Potassium hydroxide	KOH	Soluble
Sodium hydroxide	NaOH	Soluble
Ammonia solution	NH_4OH	Soluble
Calcium oxide	CaO	Slightly soluble

Calcium hydroxide	$Ca(OH)_2$	Slightly soluble
Magnesium oxide	MgO	Slightly soluble
Copper (II) oxide	CuO	Insoluble
Iron (II) hydroxide	$Fe(OH)_2$	Insoluble
Zinc hydroxide	$Zn(OH)_2$	Insoluble

STRONG AND WEAK ALKALIS

A **strong** alkali is one which ionises completely in aqueous solution, for example:

$$KOH(aq) \rightarrow K^{+}(aq) + OH^{-}(aq)$$

$$NaOH(aq) \rightarrow Na^{+}(aq) + OH^{-}(aq)$$

$$Ca(OH)_2(aq) \rightarrow Ca^{+}(aq) + OH^{-}(aq)$$

A **weak** alkali is only partially ionised in aqueous solution, for example:

$$NH_4OH(aq) \rightleftharpoons NH_4^{+}(aq) + OH^{-}(aq)$$

The reversible symbol is used in place of the arrow to indicate partial ionisation.

THE EFFECT ON ALKALIS ON INDICATORS

Each one of the indicators shows a characteristic colour in a solution of a particular alkali. Red litmus paper turns blue in sodium hydroxide, calcium hydroxide and ammonia solution.

Methyl orange turns yellow whereas phenolphthalein turns pink in the above solutions. Litmus paper would turn blue in potassium hydroxide solution.

Action of acids and alkalis on indicators

Indicator	Colour in:	
	Acid	**Alkali**
Litmus	Red	Blue
Methyl orange	Red	Yellow
Phenolphthalein	Colourless	Pink or red

Bromothymol blue	Yellow	Blue

THE EFFECT OF ALKALIS ON ACIDS

As the sodium hydroxide is slowly added to the dilute hydrochloric acid, the colour of the litmus paper at a particular time turns from red to pale blue. This is because as the sodium hydroxide is added to the acid, the acid becomes weaker until there is no acid left.

This process is called neutralisation.

Neutralisation is the formation of molecules of water from oxide (O^-) or hydroxide (OH^-) ions form a base and hydrogen ions (H^+) from an acid.

$$2H^+ (aq) + O^{2-} (aq) \rightarrow H_2O(l)$$

$$H^+ (aq) + OH^- (aq) \rightarrow H_2O(l)$$

Complete neutralisation of the acid by the alkali occurs when the litmus paper changes colour from red to pale blue.

A salt is also formed at the same time as the water. On evaporation, the resulting solution gives a white crystalline solid which is sodium chloride, the salt.

A **salt** is a compound formed when the hydrogen ions in an acid are replaced by metal or ammonium ions.

THE EFFECT OF ALKALIS ON SOLUTIONS OF SALTS

When sodium hydroxide solution is added to a solution of copper (II) sulphate, an insoluble compound, copper (II) hydroxide is formed as a **blue precipitate:**

$$CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$$

Ionically:

$$Cu^{2+} (aq) + 2OH^- (aq) \rightarrow Cu(OH)_2(s)$$

Similarly, sodium hydroxide solution produces a white precipitate of insoluble zinc hydroxide when it is added to zinc chloride solution:

$$ZnCl_2(aq) + 2NaOH(aq) \rightarrow Zn(OH)_2(s) + 2NaCl(aq)$$

Ionically:

$$Zn^{2+}(aq) \rightarrow Zn(OH)_2(s)$$

There is no precipitate formed when sodium hydroxide reacts with magnesium sulphate.

To sum up, alkalis **precipitate** insoluble hydroxides from solutions of some salts.

THE EFFECT OF ALKALIS ON AMMONIUM SALTS

When sodium hydroxide solution is added to ammonium chloride solution and the mixture warmed, a pungent smelling gas is evolved which turns moist red litmus paper blue. The gas is alkaline and is ammonia.

$$NH_4Cl(aq) + NaOH(aq) \rightarrow NaCl(aq) + NH_3(g) + H_2O(l)$$

Ionically:

$$NH_4^+(aq) + OH^-(aq) \rightarrow NH_3(g) + H_2O(l)$$

Alkalis react with ammonium salts to produce **ammonia gas.**

USES OF SOME ALKALINE SUBSTANCES

Ash and lime are often used to treat acidic soil in agriculture. Most crop plants do not grow well in soils that contain too much acid. To neutralise the acid, lime and sometimes ash, is added to the soil. Apart from neutralising the acid in the soil, these substances also act as fertilisers.

Sodium hydroxide (caustic soda) is used in reaction with fatty acids, in the chemical industry, to make soap.

Sodium carbonate (washing soda) is used to soften hard water and remove grease. It is also used in the manufacture of glass and in soap making.

7.3 PH SCALE AND CLASSIFICATION OF OXIDES

pH SCALE

The degree of acidity or alkalinity of a substance can be measured by the pH scale using universal indicator.

Universal indicator is a mixture of dyes. The indicator is available in either solution or paper form. It undergoes various colour changes depending on the **pH**, which is the degree of acidity or alkalinity of a solution.

Colours of universal indicator according to the pH value

pH value	Colour of universal indicator
0-2	Red
3	Pink
4	Brown
5	Yellow
6-8	Green
9-10	Blue
11-12	Indigo
13-14	Violet

The pH scale consists of numbers ranging from 0 to 14. These pH numbers are used to compare the strengths of acids and alkalis.

When the solution is neutral, the pH is equal to 7. Pure water for instance has a pH equal to 7. The smaller the pH value, the stronger the acid. The larger the pH value, the stronger the alkali. Alkalis have a pH value greater than 7. The diagram below shows the pH scale.

THE pH SCALE

THE IMPORTANCE OF pH VALUE IN AGRICULTURE

Agriculturalists and good farmers are interested in obtaining high yields of crop plants in the fields. They know which plants grow well in which type of soil. Some plants grow well in acid soils while others grow in alkaline soils.

Taking a pH measurement of a soil will show the acidity or alkalinity of the soil. If, for example, a soil is discovered to be too acidic for the good growth of crops, some lime is added to the soil to neutralise the acid. Conversely, if the soil is found to be too alkaline for the good growth of plants, then a fertiliser such as ammonium sulphate may be applied to improve the condition of the soil. Such fertilisers improve not only the nature of the soil but also the fertility of the soil.

CLASSIFICATION OF OXIDES

Acidic oxides are oxides of non-metals which dissolve in water to form acids. They are formed when non-metals burn in oxygen. For example:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

These oxides form acidic solutions with water. For example:

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

$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$$

$H_2SO_3(aq)$ is called **sulphurous acid.**

Basic oxides are oxides of metals. They are called bases. Most basic oxides are insoluble in water. Those which are soluble form alkalis. For example:

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

Amphoteric oxides are oxides which are both acidic and basic. They neutralise both acids and alkalis. Examples include the oxides aluminium, zinc, tin and lead. Zinc oxide, for example, neutralises sulphuric acid to form zinc sulphate:

$$ZnO(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2O(l)$$

Zinc oxide also reacts with sodium hydroxide to form a complex salt, sodium zincate.

$$ZnO(s) + 2NaOH(aq) \rightarrow Na_2ZnO_2(aq) + H_2O(l)$$

Neutral oxides are oxides which have neither acidic nor basic properties. These oxides are usually the lower oxides of non-metals, for example carbon monoxide (CO), hydrogen oxide (water, H_2O) and dinitrogen oxide (N_2O).

EXAM TYPE QUESTIONS WITH ANSWERS IV

1 Zirconium (Zr) is a metal in Period 5. Its main oxidation state is +4.

(a) The following are all zirconium atoms: $^{90}_{40}Zr$, $^{91}_{40}Zr$ and $^{92}_{40}Zr$.

In terms of numbers of electrons, neutrons and protons, how are these three atoms the same and how are they different?

They are the same because ....

....

They are different because ....

.... [3]

(b) Containers for fuel rods in nuclear reactors are made of zirconium.
Nuclear reactors are used to produce energy and to make radioactive isotopes.

(i) Which isotope of a different element is used as a fuel in nuclear reactors?

.... [1]

(ii) State one medical and one industrial use of radioactive isotopes.

....

.... [2]

(iii) Above 900 °C, zirconium reacts with water to form zirconium(IV) oxide, ZrO_2, and hydrogen. Write an equation for this reaction.

.... [2]

(iv) In a nuclear accident, water may come in contact with very hot zirconium.
Explain why the presence of hydrogen inside the reactor greatly increases the danger of the accident.

.... [1]

(c) It is possible to determine whether zirconium(IV) oxide is acidic, neutral, basic or amphoteric using an acid and an alkali. Complete the table of possible results. If the oxide is predicted to react write 'R', if it is predicted not to react write 'NR'.

if the oxide is	predicted result with hydrochloric acid	predicted result with aqueous sodium hydroxide
acidic		
neutral		
basic		
amphoteric		

[4]

ANSWERS

1 **(a** same number of protons [1]
same number of electrons [1]
different number of neutrons [1]

(b) ^{235}U / ^{239}Pu [1]
NOTE: need symbol or name and nucleon number

(ii) treating cancer / chemotherapy / radiographs / tracer studies / x–ray (scans) / sterilise surgical instruments / diagnose or treat thyroid disorders / radiotherapy [1]

paper thickness / steel thickness / radiographs / welds / tracing / fill levels in packages / food irradiation / smoke detectors [1]
ACCEPT: any other uses

(iii) $Zr + 2H_2O \rightarrow ZrO_2 + 2H_2$ [2]
not balanced = (1) only

(iv) hydrogen explodes / fire (risk) [1]

(c)

if the oxide is	predicted result with hydrochloric acid	predicted result with aqueous aqueous sodium hydroxide
acidic		R
neutral		NR
basic		NR
amphoteric		R

(1) per l [4]

[Total: 13]

2 Sulfuric acid is a strong acid. Hexanesulfonic acid is also a strong acid. It has similar properties to sulfuric acid.

(a) Sulfonic acids are made from alkanes and oleum, $H_2S_2O_7$.

$$C_6H_{14} + H_2S_2O_7 \rightarrow C_6H_{13}SO_3H + H_2SO_4$$

(i) Describe how oleum is made from sulfur by the Contact process. Give equations and reaction conditions.

....

....

....

....

....

....

.... [6]

(ii) How is concentrated sulfuric acid made from oleum?

.... [1]

(b) The formula of the hexanesulfonate ion is $C_6H_{13}SO_3^-$.

The formula of the barium ion is Ba^{2+}. What is the formula of barium hexanesulfonate?

.... [1]

(c) Complete the following equations.

(i) magnesium + hexanesulfonic acid → .... + ... [1]

....

(ii) calcium oxide + hexanesulfonic acid → .... + ... [1]

....

(iii) .....$C_6H_{13}SO_3H$ + Na_2CO_3 → .... + ... + ... [2]

(d) (Sulfuric acid is a strong acid.
You are given aqueous sulfuric acid, concentration 0.1 mol/dm^3, and aqueous hexanesulfonic acid, concentration 0.2 mol/dm^3. Describe how you could show that hexanesulfonic acid is also a strong acid.

....

.... [2]

(ii) Deduce why, for a fair comparison, the two acid solutions must have different concentrations.

....

.... [1]

(iii) Explain the terms *strong acid* and *weak acid*.

....

....

.... [2]

[Total: 17]

ANSWERS

2	**(a**	**(i)**	$S + O_2 \rightarrow SO_2$ **or** sulfur burnt / roasted / heated in air to form sulfur dioxide	[1]
			$2SO_2 + O_2 \rightleftharpoons 2SO_3$ unbalanced = (1) only	[2]
			(catalyst) vanadium(V) oxide / vanadium pentoxide (temperature) 440 to 460 °C (dissolve) sulfur trioxide in sulfuric acid (to form oleum) ignore comments about pressure	[1] [1 [1]
		(ii)	add oleum to water	[1]
	(b)		$Ba(C_6H_{13}SO_3)_2$ / $(C_6H_{13}SO_3)_2Ba$	[
	(c)		→ magnesium hexanesulfonate + hydrogen	[1]
		(ii)	→ calcium hexanesulfonate + water	[1]
		(iii)	$2C_6H_{13}SO_3H + Na_2CO_3 \rightarrow 2C_6H_{13}SO_3Na + CO_2 + H_2O$	
			$C_6H_{13}SO_3Na$ = (1) remaining species correct and equation balanced = (1)	[1] [1]

(d) measure pH / add universal indicator [1]
both acids have a low value / pH 0–2 / same colour / red [1]
or
measure rate with named reactive metal, Mg, Zn (1)
both fast reactions (1)
or
measure rate using piece of insoluble carbonate, $CaCO_3$ (1)
both fast reactions (1)
NOTE: must be insoluble for first mark
or
measure electrical conductivity (1)
both good conductors (1)

(ii) to have same concentration of H^+ / one acid is H_2SO_4, the other is $C_6H_{13}SO_3H$ / sulfuric acid is dibasic, hexanesulfonic is monobasic [1]

(iii) a strong acid is completely ionised, [1]
a weak acid is partially ionised [1]

[Total: 17]

8. INTRODUCTION TO ORGANIC CHEMISTRY

8.1 ORGANIC AND INORGANIC COMPOUNDS

Carbon has a large variety of compounds. Most of such compounds produce water and carbon dioxide when heated with copper (II) oxide. These carbon-containing compounds are called **organic compounds.**

When naphthalene is heated with copper (II) oxide, water and carbon dioxide are produced. The presence of water is tested using cobalt (II) chloride paper. Water changes the colour of cobalt (II) chloride paper from blue to pink. The presence of carbon dioxide is tested using lime water. When carbon dioxide is bubbled through lime water a cloudy white precipitate is formed.

Many organic compounds burn forming carbon dioxide and water, for example:

$CH_4(g)$	+	$2O_2(g)$	→	$CO_2(g)$	+	$2H_2O(g)$
Methane (organic compound)		Oxygen		Carbon Dioxide		Water

Most carbonates and oxides of carbon do not form water and carbon dioxide when heated with copper(II) oxide. This indicates that they are not organic compounds. Some carbonates undergo thermal decomposition releasing carbon dioxide.

SOLUBILITY OF ORGANIC AND INORGANIC COMPOUNDS

Sodium chloride is soluble in water but insoluble in acetone and ethanol. Naphthalene is insoluble in water but soluble in acetone and ethanol.

Acetone and ethanol are organic liquids but water is inorganic. Sodium chloride is an inorganic substance whereas naphthalene is organic.

Generally inorganic compounds dissolve readily in water whereas organic compounds dissolve in organic liquids such as ethanol, acetone, carbon disulphide and tetrachloromethane.

DO ORGANIC AND INORGANIC COMPOUNDS CONDUCT ELECTRICITY

Generally solutions of organic compounds do not conduct electricity whereas inorganic ones do, whether in molten or in solution form.

Some characteristics of organic and inorganic compounds

Organic Compounds	Inorganic Compounds
Insoluble in water but soluble in organic solvents	Soluble in water but insoluble in organic solvents
Do not conduct electricity	Conduct electricity
Have low melting points	Have high melting points
Volatile	Non-volatile
Flammable	Not easily flammable

8.2 HYDROCARBONS

Hydrocarbons are organic compounds whose molecules consist of the elements carbon and hydrogen only. There are two types of hydrocarbons, namely, **saturated** and **unsaturated** hydrocarbons.

Saturated hydrocarbons have molecules consisting of **single covalent bonds** only. An example of a group of saturated hydrocarbons is the **alkanes**. Unsaturated hydrocarbons have multiple bonds and an example of this group is the **alkenes.**

ALKANES

Alkanes are saturated hydrocarbons. They possess the general molecular formula C_nH_{2n+2}, where n is a whole number.

The simplest alkane is methane consisting of only one carbon atom bonded to four hydrogen atoms by single covalent bonds.

```
    H
    |
H– C–H
    |
    H
```

This representation of methane is its **structural formula**. Its **molecular formula** is **CH_4.** The names and formulae of some alkanes are listed below.

The first five alkanes

Alkane	Structural formula	Molecular formula
Methane	**H | H– C–H | H**	**CH_4**
Ethane	**H H | | H– C– C–H | | H H**	**C_2H_6**
Propane	**H H H | | | H– C– C– C–H | | | H H H**	**C_3H_8**
Butane	**H H H H | | | | H– C– C– C– C–H | | | | H H H H**	**C_4H_{10}**
Pentane	**H H H H H | | | | | H– C– C– C– C– C–H | | | | | H H H H H**	**C_5H_{12}**

Note that the ending for the names of all of the alkanes is '-ane'.

There is another way to write molecular formulae. For example the molecular formula of ethane may be written as $CH_3{\cdot}CH_3$, of propane as $CH_3{\cdot}CH_2{\cdot}CH_3$, of butane as $CH_3{\cdot}CH_2{\cdot}CH_2{\cdot}CH_3$ and of pentane as $CH_3{\cdot}CH_2{\cdot}CH_2{\cdot}CH_2{\cdot}CH_3$.

PHYSICAL PROPERTIES OF ALKANES

Alkanes of up to four carbon atoms in each molecule are **gases** at room temperature. Alkanes with five to 17 carbon atoms are **liquids** at room temperature. Alkanes with 18 or more carbon atoms in each molecule are **solids** at room temperature.

Generally alkanes have low melting and boiling points. Alkanes are insoluble in water, but they dissolve in organic solvents.

REACTIONS OF ALKANES

COMBUSTION REACTION

Alkanes burn in air forming carbon dioxide and water.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

When the air supply is limited, incomplete combustion occurs, producing carbon monoxide.

$$2CH_4(g) + 3O_2(g) \rightarrow 2CO(g) + 4H_2O(g)$$

SUBSTITUTION REACTIONS

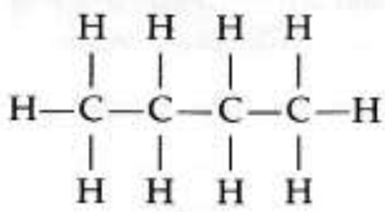
Alkanes react with members of the halogen group, for example chlorine, bromine and iodine. Atoms of these elements take the place of the hydrogen atom or atoms in the alkane. Hydrogen chloride gas is produced.

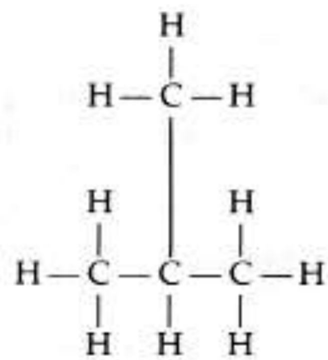
$$CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g)$$

ISOMERISM

Some organic compounds have the same molecular formula but different structural formula.

For example the formula C_4H_{10}.





The structure on the left is butane. The structure on the right is a different substance called methylpropane with a branched carbon chain.

Despite having the same molecular formula they have different structures and because of this they have different characteristics, or properties.

Characteristics of butane and methylpropane

Characteristic	Butane	Methylpropane
Melting point (°C)	-138	-160
Boiling point (°C)	-0.5	-11.7

Such compounds which have the same molecular formula but different structural formula are called **isomers**, and the phenomenon is called **isomerism.**

All alkanes with four or more carbon atom exhibit isomerism. The greater the number of carbon atoms the more the number of isomers.

HOMOLOGOUS SERIES

This is a series where each member differs from the next by one $\cdot CH_2$ unit and conforms to general molecular formula.

The following is a summary of the characteristics of a homologous series.

- All members can be represented by a general molecular formula.
- Each member in the series differs from the next by one $\cdot CH_2$ unit.
- The members in the series show similar chemical properties.
- The physical properties of the members show a gradual change along the series.

- The members in the series have similar methods of preparation.

Some elements and some of their properties

Element		Atomic number	Approximate relative atomic mass	Melting point (°C)	Boiling point (°C)
Name	**Symbol**				
Hydrogen	H	1	1	-259	-253
Helium	He	2	4	-272	-269
Lithium	Li	3	7	181	1342
Beryllium	Be	4	9	1278	2970
Boron	B	5	11	2300	2550
Carbon	C	6	12	3652	4827
Nitrogen	N	7	14	-210	-196
Oxygen	O	8	16	-218	-183
Fluorine	F	9	19	-220	-188
Neon	Ne	10	20	-248	-246
Sodium	Na	11	23	98	883
Magnesium	Mg	12	24	649	1107
Aluminium	Al	13	27	660	2467
Silicon	Si	14	28	1410	2355
Phosphorus	P	15	31	44	280
Sulphur	S	16	32	119	445
Chlorine	Cl	17	35.5	-101	-35
Argon	Ar	18	40	-189	-186
Potassium	K	19	39	63	760
Calcium	Ca	20	40	839	1484

Atomic Number — 6
Symbol — C
Name — Carbon
Average Atomic Mass — 12.011

metals (pink) / nonmetals (blue) / metaloids (green)

1 H Hydrogen 1.008																		2 He Helium 4.003
3 Li Lithium 6.94	4 Be Beryllium 9.012												5 B Boron 10.81	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998	10 Ne Neon 20.180
11 Na Sodium 22.990	12 Mg Magnesium 24.305												13 Al Aluminum 26.982	14 Si Silicon 28.085	15 P Phosphorus 30.974	16 S Sulfur 32.06	17 Cl Chlorine 35.45	18 Ar Argon 39.948
19 K Potassium 39.098	20 Ca Calcium 40.078		21 Sc Scandium 44.956	22 Ti Titanium 47.867	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.938	26 Fe Iron 55.845	27 Co Cobalt 58.933	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.723	32 Ge Germanium 72.630	33 As Arsenic 74.922	34 Se Selenium 78.97	35 Br Bromine 79.904	36 Kr Krypton 83.798
37 Rb Rubidium 85.468	38 Sr Strontium 87.62		39 Y Yttrium 88.906	40 Zr Zirconium 91.224	41 Nb Niobium 92.906	42 Mo Molybdenum 95.95	43 Tc Technetium [97]	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.906	46 Pd Palladium 106.42	47 Ag Silver 107.868	48 Cd Cadmium 112.414	49 In Indium 114.818	50 Sn Tin 118.710	51 Sb Antimony 121.760	53 Te Tellurium 127.60	53 I Iodine 126.904	54 Xe Xenon 131.293
55 Cs Cesium 132.905	56 Ba Barium 137.327	* 57 - 70	71 Lu Lutetium 174.967	72 Hf Hafnium 178.49	73 Ta Tantalum 180.948	74 W Tungsten 183.84	75 Re Rhenium 186.207	76 Os Osmium 190.23	78 Ir Iridium 192.217	79 Pt Platinum 195.084	80 Au Gold 196.997	81 Hg Mercury 200.592	81 Tl Thallium 204.38	82 Pb Lead 207.2	83 Bi Bismuth 208.980	84 Po Polonium [209]	85 At Astatine [210]	86 Rn Radon [222]
87 Fr Francium [223]	88 Ra Radium [226]	** 89 - 102	103 Lr Lawrencium [262]	104 Rf Rutherfordium [267]	105 Db Dubnium [270]	106 Sg Seaborgium [269]	107 Bh Bohrium [270]	108 Hs Hassium [270]	109 Mt Meitnerium [278]	110 Ds Darmstadtium [281]	111 Rg Roentgenium [281]	112 Cn Copernicium [285]	113 Nh Nihonium [286]	114 Fl Flerovium [289]	115 Mc Moscovium [289]	116 Lv Livermorium [293]	117 Ts Tennessine [293]	118 Og Oganesson [294]

*Lanthanide series	57 La Lanthanum 138.905	58 Ce Cerium 140.116	59 Pr Praseodymium 140.908	60 Nd Neodymium 144.242	61 Pm Promethium [145]	62 Sm Samarium 150.36	63 Eu Europium 151.964	64 Gd Gadolinium 157.25	65 Tb Terbium 158.925	66 Dy Dysprosium 162.500	67 Ho Holmium 164.930	68 Er Erbium 167.259	69 Tm Thulium 168.934	70 Yb Ytterbium 173.045
**Actinide series	89 Ac Actinium [227]	90 Th Thorium 232.038	91 Pa Protactinium 231.036	92 U Uranium 238.029	93 Np Neptunium [237]	94 Pu Plutonium [244]	95 Am Americium [243]	96 Cm Curium [247]	97 Bk Berkelium [247]	98 Cf Californium [251]	99 Es Einsteinium [252]	100 Fm Fermium [257]	101 Md Mendelevium [258]	102 No Nobelium [259]

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