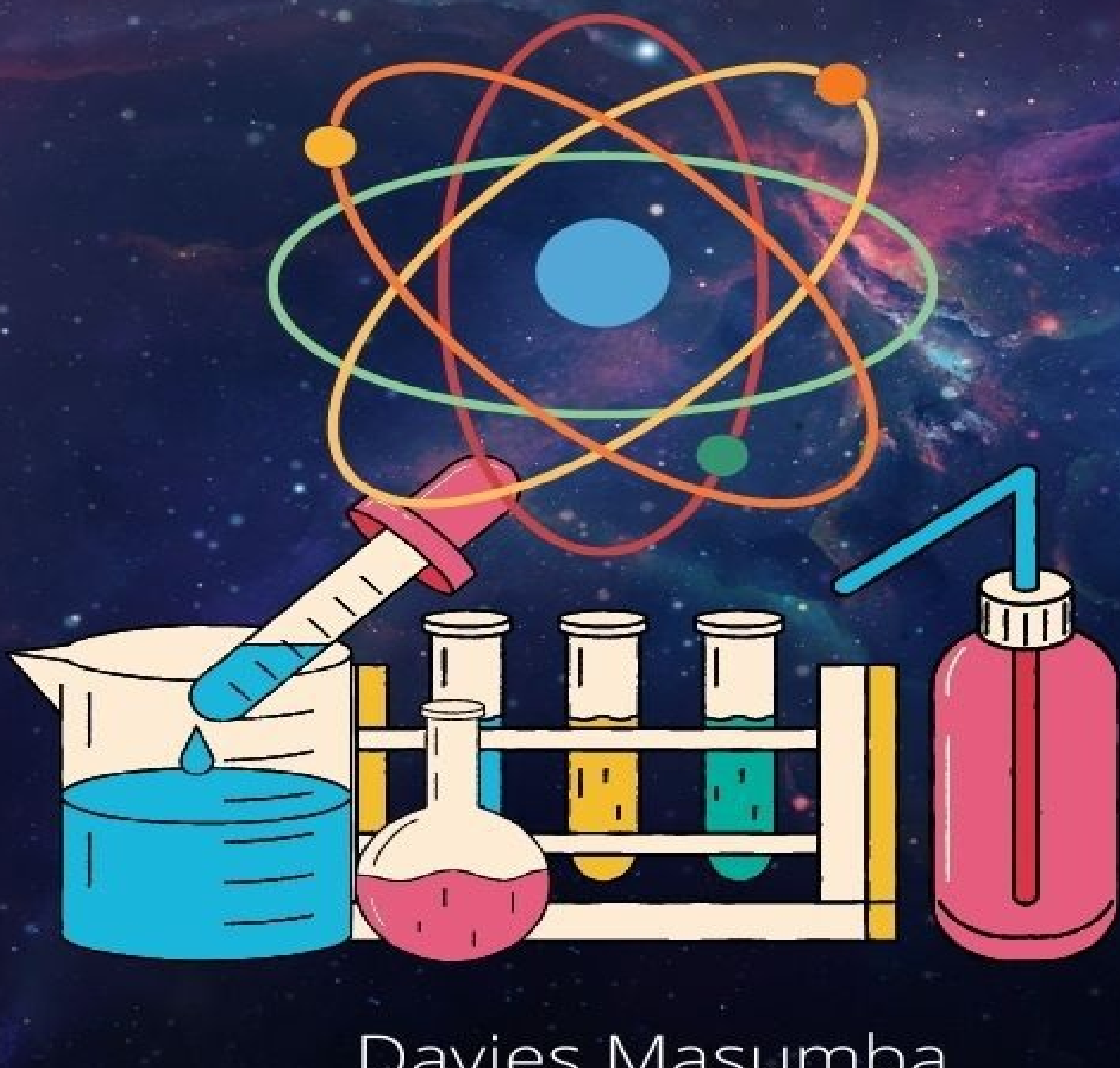
CHEMISTRY FOR SECONDARY SCHOOL VOLUME 1

INTRODUCTION TO CHEMISTRY



Davies Masumba

CHEMISTRY FOR SECONDARY SCHOOL

INTRODUCTION TO CHEMISTRY

DAVIES MASUMBA

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This book is a dedication to my mother Mary Mwanaumo for bringing up despite the hard times she was faced with bringing up children single handedred.

Contents

INTRODUCTION TO CHEMISTRY

- Chemistry is the study of the composition of matter and the chemical changes matter undergoes under different conditions of pressure and temperature.
- Chemistry is studied to know what matter is like and how it behaves, and our explanations and predictions of its behaviour.

BRANCHES OF CHEMISTRY

Chemistry is branched into several branches some of which are:

- **Analytical chemistry**: This is a branch of chemistry which deals with the study of composition of substances qualitatively and quantitatively.
- **Physical chemistry:** This is a branch of chemistry which deals with the study of physical principles that underlie structure of matter and chemical transformation.
- **Organic chemistry:** This is a branch of chemistry which deals with the study of compounds of carbon.
- **Inorganic chemistry:** This is a branch of chemistry which deals with the study of all elements except carbon but includes simple compounds of carbon such carbon dioxide.
- **Biochemistry:** This is a branch of chemistry which deals with the study of living systems, both plants and animals.

IMPORTANCE OF CHEMISTRY

Chemistry is important because it has many applications that improves life in societies. The following are some of the applications of chemistry in real life:

- **Construction industry**: Chemistry helps in the production of glass, cement, bitumen and metals used for construction of buildings, bridges and roads in the construction industry.
- **Packaging industry**: Chemistry helps in the production of plastic papers, glassware and containers used for packaging and storing things in the packaging industry.
- **Manufacturing industry**: Chemistry provides knowledge applied in manufacturing industries to produce raw materials for various

purposes.

- **Energy production industry**: Chemistry helps in the production of chemical energy in batteries and fuels such as petrol. For example, petrol is obtained from crude oil by fractional distillation.
- **Transportation industry**: Chemistry helps in the production of fuels used to power motor vehicles, trains, ships and aeroplanes for transporting people from one place to another.
- **Food industry**: Chemistry helps in the production of baking powder, yeast and knowledge used to make foods such biscuits and bread.
- **Household goods**: Chemistry helps in the production of soaps, detergent, sugar, medicine, plastics, containers, tissues, salt and many products used in households.
- **Agriculture industry**: Chemistry helps in the production of fertilizers and pesticides used in agriculture for production of food and other industrial raw materials.

- **Environmental protection**: Chemistry helps us with knowledge to better understand and care for our bodies, mind, environment and the world at large. It also enables us to better understand the benefits and hazards of this world and enables us to make intelligent decisions.

- **Career choice**: Chemistry equips us with knowledge which we need in chemistry related careers such Chemical engineering, Agricultural engineering, Biochemist, Electrical engineering, Electrical engineering, Metallurgist, Nursing, Medicine (medical doctor), Food science technology, Science teacher, Pharmacist, Chemist analyst, Research chemist, Medical laboratory technician, Environmental ecologist, Chemicals / fertilizer salesman and Veterinary.

CHALLENGES OF CHEMICAL INDUSTRIAL ACTIVITIES

Despite chemical industrial activities having many applications in real life, these activities have also negative effects such as:

- **Air pollution**: Combustion of fossil fuels in industrial machines and motor vehicles emits harmful and toxic gases into the atmosphere which causes

- **Respiratory problems**: Caused by breathing in nitrogen oxide from car exhaust and hot furnaces.
- **Eye and throat irritation**: Caused by sulphur dioxide in air from combustion of fossil fuels in power stations.
- **Acid rain**: Caused by sulphur dioxide in air from combustion of fossil fuels in power stations dissolved in rain.
- **Oxygen starvation**: Caused by breathing in Carbon monoxide from incomplete combustion of carbon compounds in car engines which in reacts with the haemoglobin in blood preventing it from carrying oxygen around the body – so one dies from oxygen starvation.

- **Water pollution**: Chemical fertilizers and nutrients run-off from farms and gardens cause the build-up of toxic algae in rivers, making uninhabitable to aquatic organisms and unpleasant to man.
- **Global warming:** This is the average rising in temperatures around the globe caused by greenhouse gases from combustion of carbon compounds. Greenhouses gases in the atmosphere accumulates to the ozone layer thereby absorbing and preventing heat from escaping into outer space. This causes the earth to warm up.
- **Climate change**: Climate change is the average change in atmospheric conditions such as temperature, rainfall, cloud cover and winds of an area over a long period of time due to global warming. Climate change leads to:

- Flooding of coastal areas due to melting of ice at arctic circles.
- Change in weather patterns such as rainfall.
- Outbreak of diseases and hunger leading to poverty.

SAFETY IN THE LABORATORY

Conducting experiment in the laboratory can be quiet dangerous and because of that, a number of rules and regulations have to be followed to make the laboratory a safe place.

1. Organized work area
2. Always have a clean workplace.
3. Remember to always wear safety googles in the lab to avoid contact of chemicals in the face.
4. Do not use a glassware with a crack.
5. Do not smell the chemical directly in the mouth of a flask. It may

irritate your nose and respiratory tract. To smell, use the waving method. Fan the air towards your nose to smell the chemical.

6. Do not taste chemicals. Most of them are poisonous and only for lab purposes.
7. Wash your hands before going out of the lab to get rid of the chemicals in your hands.
8. Clean and return the tools to where you got them. Never leave any spillage or tool on the work place.
9. Use only the tools according to their proper usage.
10. Don’t wear sandals or open shoes.
11. Long hair must be tied back.
12. Laboratory coats must be put on fastened.
13. Never eat, drink or smoke in a laboratory.
14. Never touch your face, mouth or eyes.
15. Never suck pens or chew pencils.

THE PARTICULATE NATURE OF MATTER

MATTER

Matter is defined as the material of the universe which occupies space and has mass. All the different substances around us that make up the universe are all called **matter** provided they have mass and take up space.

Matter and the Kinetic Theory

The kinetic theory of matter states that matter is made up of very tiny particles that are in continuous motion. The tiny particles which make up matter are called the basic units of matter. These are:

- Atoms,
- Molecules, and
- Ions.

The kinetic theory of matter helps to explain the following:

- Classification of matter into three states such as solid liquid and gas. According to the kinetic theory matter is classified into three states due to the way particles are arranged, motion and forces of attraction.
- Change of state of matter.

Examples of substances made of the three basic units of matter

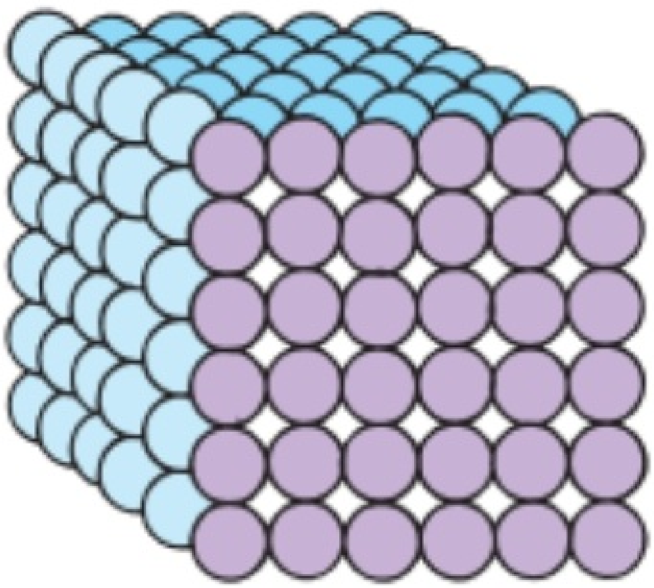
- **Atoms:** Noble gases are examples of substances made up of single atoms.
- **Ions:** Ionic compounds such sodium chloride and potassium manganate (VII) are examples of substances that consist of ions (charged atoms) joined together in large number.
- **Molecules:** Covalent compounds oxygen gas and water are some

of examples of substances that consists of two or more atoms joined together. Water, bromine, and the gases nitrogen and oxygen in air, are made up of molecules.

Structure of states of matter

SOLIDS

- **Arrangement of particles in solids:** The particles in a solid are closely arranged (packed) in regular pattern or lattice (closely packed). This is why solids have fixed volume, fixed shape and cannot be compressed.
- **Forces of attraction between particles in solids**: Strong forces hold particles together. So the particles cannot leave their positions. The only movements they make are tiny vibrations to and fro in fixed positions. This is why solids do not flow.
- **Motion of particles in solids:** Particles in solids only vibrate to and fro in fixed position because particles holding the together are very strong.

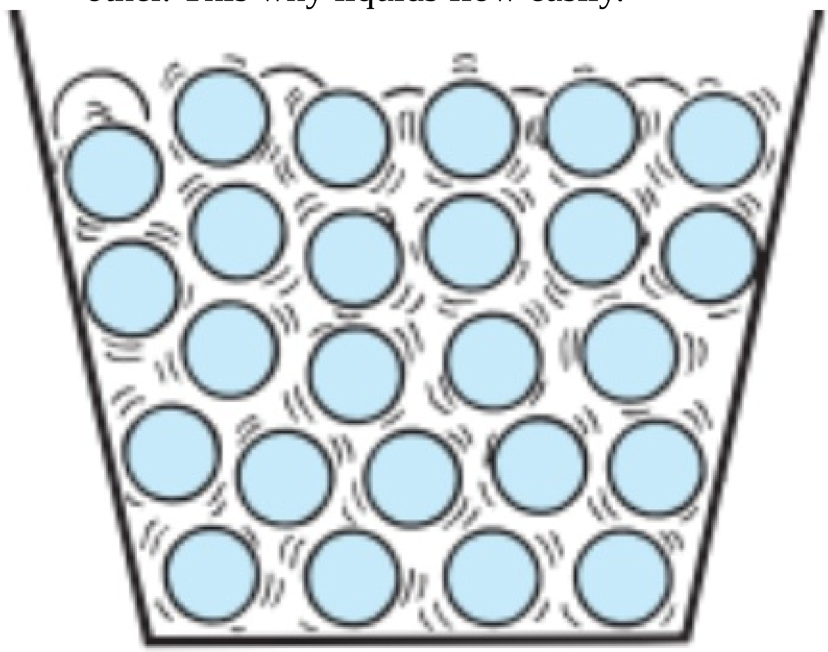


LIQUIDS

- **Arrangement of particles in liquids:** Particles in liquid are loosely packed and not regularly arranged and can move about and slide past each other. This is why liquids have fixed volume and no fixed and cannot be compressed.
- **Forces of attraction between particles in liquids:** The forces that hold

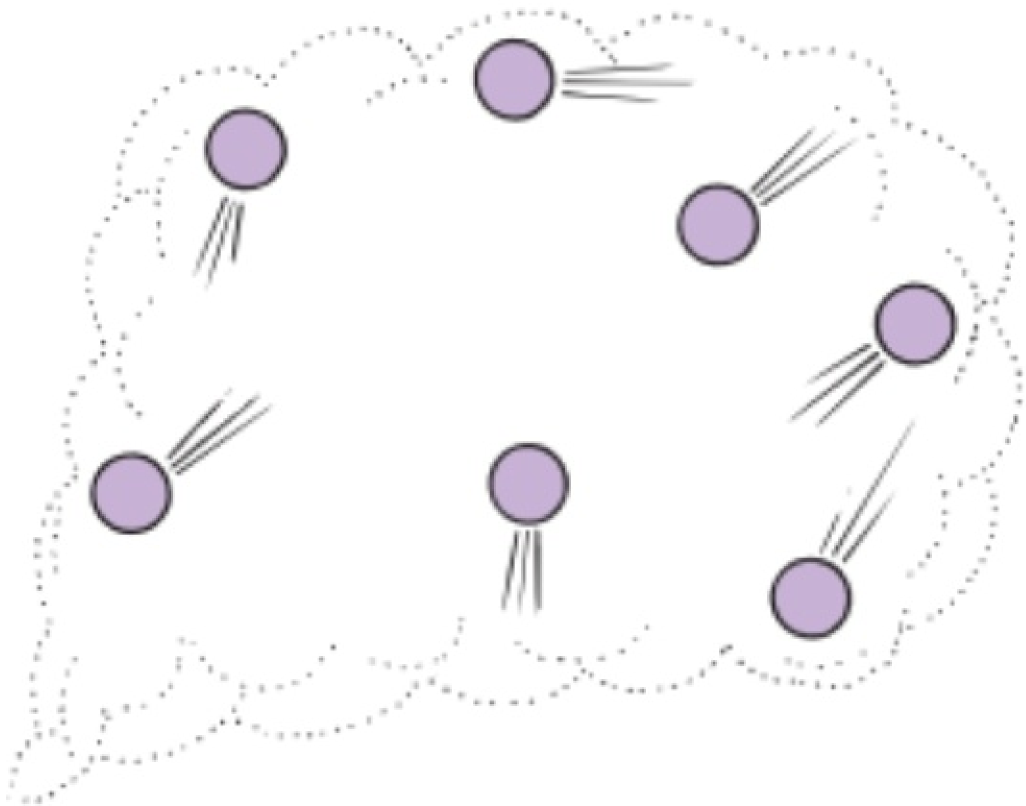
particles together in liquids are weaker than those in solids. So particles move around in a random way and often collide with one another while attracted to each other. This why liquids flow and have no fixed shape.

- **Motion of particles in liquids:** Particles in liquids move around in a random way and often collide with one another while attracted to each other. This why liquids flow easily.



GASES

- **Arrangement of particles:** In gases the particles are relatively far apart and they move about very quickly. This is why gases have no fixed volume and no fixed shape and can be compressed.



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  Forces of attraction between particles: There are almost no forces holding particles together in gases. Particles move randomly at very high speeds in the space available. This why liquids flow in all directions.
- **Motion of particles:** Particles in gases flow in all directions at very high speeds. This is because particles in gases are relatively far apart and there are almost no forces holding particles together.

Change of states of matter

- Change of state is the change of matter from one state to another. It is a physical change.
- According to the kinetic theory, matter changes state when the kinetic energy of particles is increased or decreased by heating or removing heat.
- The process which involves adding heat to a substance is called **endothermic** while that which involves removing heat is called **exothermic**.

When energy in form of heat is applied to various forms of matter, a change in state of matter occurs because the particles acquire kinetic energy and begin to move with high speed weakening the forces holding them together in the process. Likewise, when heat is removed, there is loss of kinetic energy and particles begin to slow down and get stationed in fixed positions.

The following are some of the examples of changes of states.

1) Melting,
2) Vaporization or evaporation,
3) Freezing or solidification,
4) Condensation, and
5) Sublimation.

Melting: This is the change of state of matter from solid to liquid due to heating.

- During melting the kinetic energy of particles increases as heat is increased.
- This causes particles of solid to vibrate faster at greater amplitude thereby increasing the spaces between them.
- This weakens the forces holding the particles close together.
- Hence particles start to slide past each changing positions in the process.
- When this happens, the solid has turned into a liquid.
- The temperature at which the solid changes to a liquid is called the **melting point**.

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

- During vaporisation the kinetic energy of particles increases as heat is increased.
- This causes particles in liquid to start moving faster.
- The forces of attraction between particles are greatly weakened.
- Hence particles are free to move and escape from each.
- When this happens, the liquid turns into a gas.
- The constant temperature at which a liquid turns into a gas is called **boiling point**.

Freezing: This is the change of state of matter from liquid to solid due to cooling.

- During freezing the kinetic energy of particles decreases as heat is decreased.
- This caused particles in liquids to slow down thereby decreasing the spaces between them.
- This strengthens the forces holding particles together.
- Particles begins to vibrate within fixed positions.
- When this happens, the liquid turns into a solid.
- The temperature at which the liquid turns into a solid is called **freezing point**.

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

- During condensation the kinetic energy of particles decreases as heat is decreased.
- Particles in gas slow down thereby decreasing the spaces between particles.
- This strengthens the forces of attraction between the particles.
- The weak forces are now able to hold the particles closer; the gas has turned into a liquid.
- The temperature at which a gas turns into a liquid is called **condensation point**.

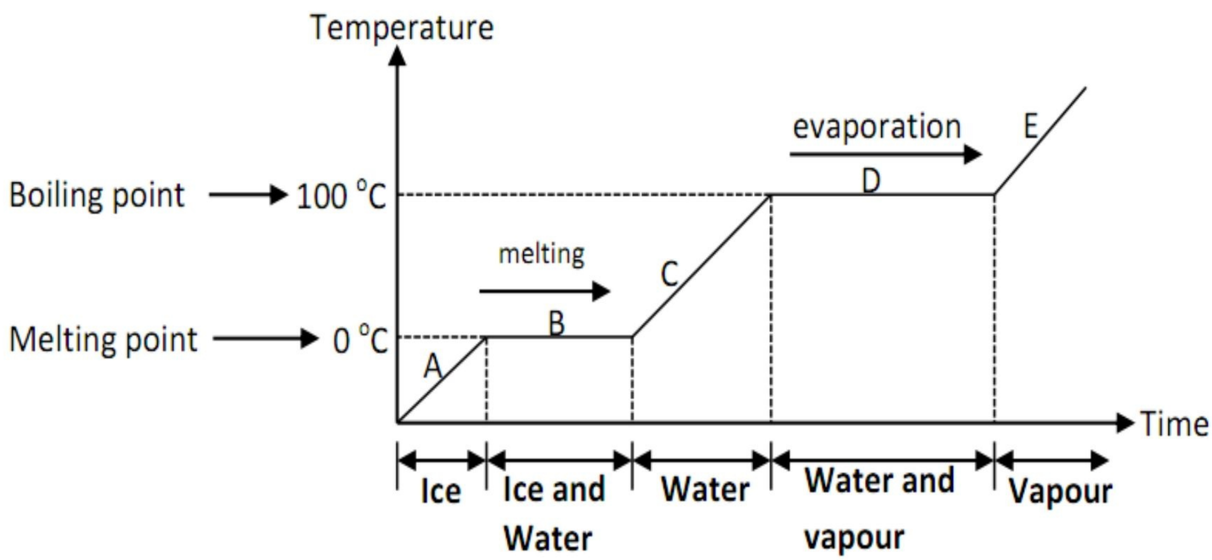
Heating and cooling curves

- These are curves which shows the variation of temperature of a substance with time during heating and cooling.
- On heating and cooling curves, the y-axis represents change in temperature while the x-axis represents time.
- The slope section represents change in temperature while the flat section represents change of state.

Heating curve for ice

- A heating curve shows the increase in temperature with time as heat is being absorbed by a substance. It is an endothermic process.

- During heat the heat content of the substance increases.



Explaining sections of the heating curve

Section A (slope):

As ice heated, it absorbs energy and its temperature rises as time elapses.

Section B (flat section):

In this section the temperature remains the same (constant) over a period of time even though ice continues to be heated. This is because the heat added is used to break bonds (forces of attraction) between particles to bring about melting of ice into liquid. There are two states in this section, solid and liquid.

Section C (slope section):

In this section all the ice is melted into water and temperature increases again with time as water continues to be heat because heat is being used to increase the temperature of particles in water.

Section D (flat section):

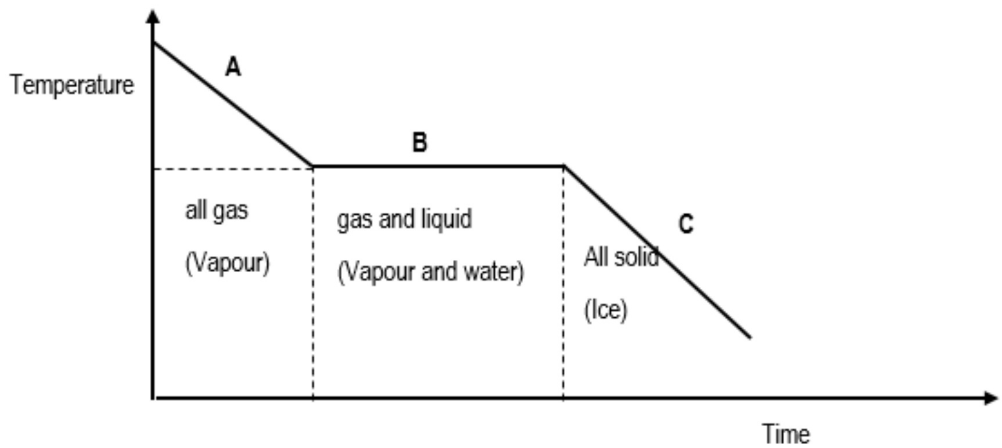
In this section temperature remains constant again over a period of time even though water continues to be heated. This is because the heat added is used to break bonds between particles in water to bring about evaporation of water into vapour. There are two states in this section, liquid and gas.

Section E (slope section):

In this section all the water has evaporated into vapour and temperature rises again with time as the vapour continues to be heated because heat is being used to increase the temperate of vapour particles.

Cooling curve of water vapour

- A cooling curve shows the decrease of temperature with time as a substance is cooled. It is an exothermic process.



- During cooling the heat content of the substance decreases.

Explaining sections of the heating curve

Section A (slope section)

In this section the temperature decreases with time as the vapour is cooled.

Section B (flat section)

In this section the temperature remains constant over a period of time even though the vapour continues to be cooled. This is because particles of vapour continue to release heat as they collide with each as they condense from vapour to liquid.

Section C (slope section)

In this section all the vapour has condensed to water and the temperature of particles decrease with time as particles start to lose while cooling continues.

The table below shows some boiling and melting points of some common substances.

Substance	Boiling point ^oC	Melting point ^oC

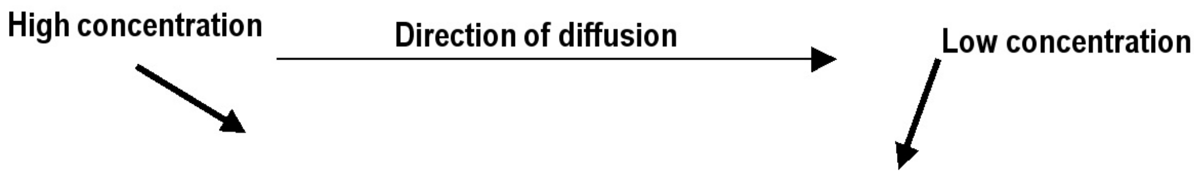
Sulphur	445	119
Water	100	0
Ethanol	78	-15
Sodium	890	98
Oxygen	-183	-219
Iron	2900	1540
Diamond	4832	3550

Evidence in support of the kinetic theory of matter

There is evidence all around us that proves that matter is made up of tiny of particles that move around in liquids and gases but vibrate about fixed positions in solids. One such great evidence for the particulate nature of matter is **diffusion**.

Definition of diffusion

Diffusion is the movement of particles from a region of higher concentration to a region of lower concentration.



Demonstrating diffusion in fluids

Experiment to demonstrate diffusion in liquids

Requirements:

- Beaker, water, potassium manganate (VII) crystals or ammonium dichromate (VI) crystals or copper (II) sulphate crystals.

Procedure:

- Water is poured into a beaker to half-full level.
- Small quantities of potassium crystal are inserted at the bottom of the beaker of water and left to stand for some time (10 minutes).

Observations:

- 

  As the crystal dissolves the colour slowly spreads through the liquid, first covering the bottom. Eventually the colour distributes itself evenly throughout the liquid.

Explanation:

- The self-uniform spreading of potassium manganate (VII) in water shows that potassium manganate dissolved into small particles which moved through the water due to collision with moving water molecules.

Experiment to demonstrate diffusion in gases
Requirements:

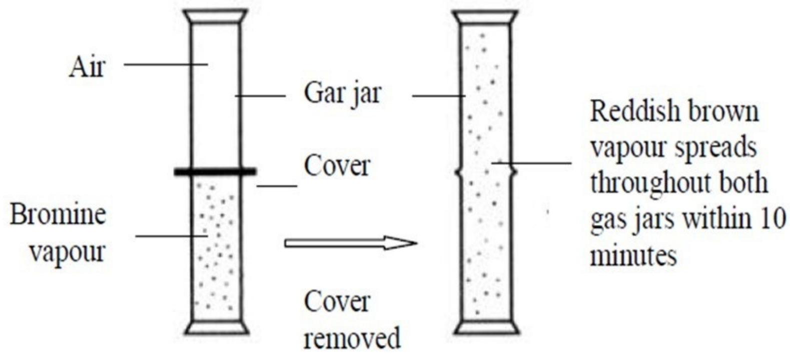
- Gas jars (2), gas jar covers, dropping pipette, white full scarp paper and liquid bromine.

Procedure:

- Using a dropping pipette/teat pipette, 2-3 drops of liquid bromine in transferred into an open gas jar and covered.
- An open gas jar of air is placed upside down on an open gas jar containing a few drops of red-brown bromine as shown below.
- The cover is then removed the jars left to stand.

Observations:

- After some good time, the brown–red fumes of gaseous bromine spread evenly throughout both gas jars from the liquid present in the lower gas jar.



Explanations:

- Bromine particles move from lower jar into spaces between air particles in upper jar.

Conclusions:

- Diffusion proceeds from a region of higher concentration to region of lower concentration to occupy the total available space.
- It is a random motion (called Brownian movement) and a long-time process since molecules experience numerous collisions while moving from one end to another.
- Diffusion is faster in gases since particles are farther apart from each other and have high kinetic energy levels.

- Diffusion in liquids is lower since particles are close together, collide more frequently and have lower kinetic energy.
- Diffusion does not occur in solids because particles do not move instead they only vibrate in fixed positions.

Evidence for random motion of particles

Brownian motion discovered by Robert brown is the continuous random motion of particles. It is evidence for the random motion of particles in gases and liquids.

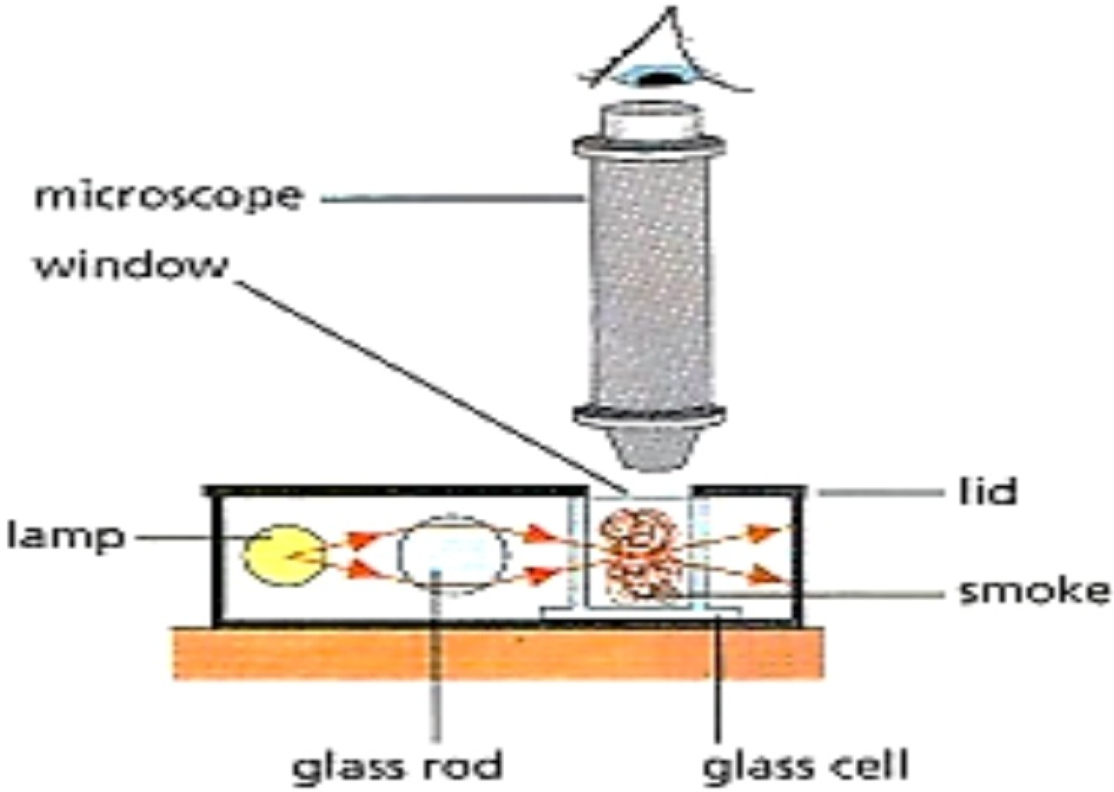
An experiment to demonstrate Brownian motion in gases

Requirements:

- A straw, smoke cell, a source of fire, a microscope and lamp.

Procedure:

- An experiment is set up as shown below, and a microscope is used to look at the smoke cell which contains smoke particles as well as air molecules (not seen by the naked eye).



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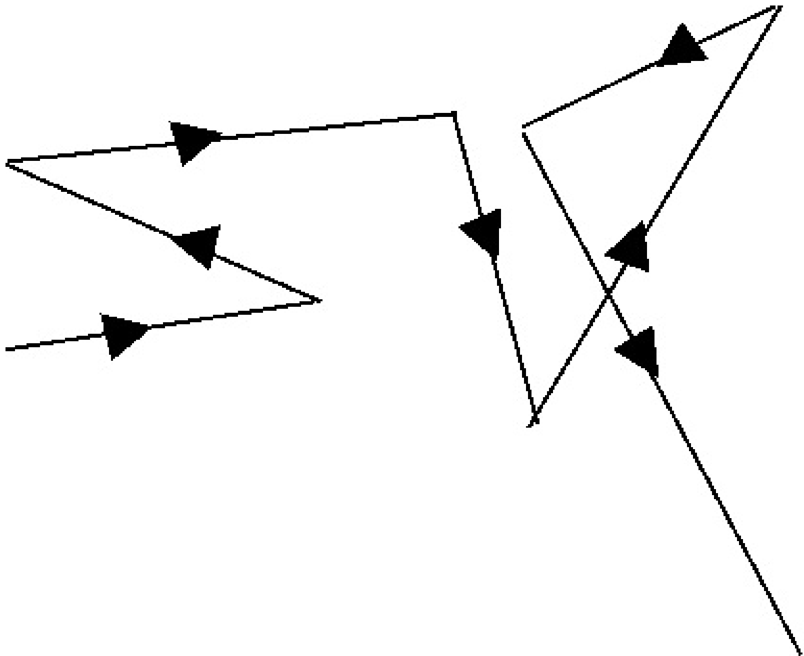
Observations:

- In bright light, smoke particles show up as bright spots which move in continuous and jerky random motion.

Explanation:

- Smoke particles being very light but large collide continuously with small air molecules that are a bit massive than them. This continuous collision of smoke

particles with air molecules causes them to change direction of motion continuously and hence seen to move in a random zigzag path.



Random paths of
a smoke particle

Conclusion:

- The movements of smoke particles are due to their random collision with the air molecules.
- The result of Brownian motion depicts the kinetic theory explanation which states that matter is made up of particles which in continuous **random** motion in liquids and gases.

FACTORS AFFECTING DIFFUSION

We have learned that diffusion is fastest in gases than liquids and that it does not occur in solids. Thus diffusion is directly proportional to the following:

1. **Distance between particles:**

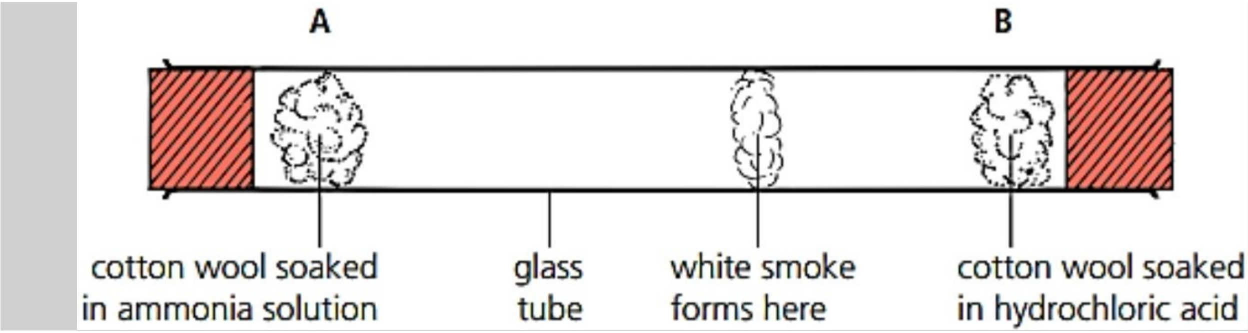
- Diffusion is the movement of particles from high concentration to lower concentration as a result of collision to fill the total available space, the more spaced particles are, the more they diffuse into the available spaces. This why diffusion is faster in gases than liquids.

2. **The kinetic energy or temperature of the particles:**

- Kinetic energy of particles depends on temperature. The higher the temperature, the faster particles move and therefore the faster they diffuse. Particles in gases have more kinetic energy than particles in liquids, this is another reason why diffusion is faster in gases.

3. **The size of the particles:**

- The lighter particles are, the faster they move. Therefore, light particles diffuses faster than heavy particles because they move fast. This is proven through an experiment set up as shown below:



- A cotton soaked in aqueous ammonia (which gives off ammonia gas) and another soaked in hydrochloric acid (which gives off hydrogen chloride gas) are placed on opposite sides of the tube.
- It is observed that, a white smoke of ammonium chloride forms closer to **B** than **A.** This is so because ammonia gas is lighter than HCl gas and diffuses faster than hydrogen chloride gas. The molecular mass of ammonia gas is less than that of hydrogen chloride.

4. **Concentration of particles:**

- The higher the concentration, the faster the diffusion. This is because an increase in concentration increases the pressure of the molecules to diffuse into the area of low concentration, (more particles trying to get to the same area at the same time).

CLASSIFICATION OF MATTER

The various kinds materials which make up matter are classified into the following classes:

1. PURE SUBSTANCES
2. MIXTURES

PURE SUBSTANCES

- **A pure substance is a substance which consists of the same substances or other substances chemically combined in fixed proportions**.
- Components of a pure substance cannot be separated into simpler substances by physical methods but can only be changed in identity and properties by chemical methods.
- Examples of pure substances includes **compounds** and **elements**.

COMPOUNDS

- **A compound is a pure substance made up of two or more elements chemically combined together in definite ratio**.
- Examples of compounds includes <u>**water**</u>, <u>**copper (II) sulphate**</u> and <u>**sodium hydroxide**</u>.
- Compounds can be decomposed into simpler substances called **elements** by chemical changes. For example, <u>**water**</u> can be decomposed into **hydrogen** and **oxygen** by chemical change called **electrolysis**, <u>**copper (II) sulphate**</u> can be decomposed into **copper**, **sulphur** and **oxygen** by a chemical process called thermal **decomposition**. Hydrogen, oxygen, copper, and sulphur are examples of elements.

- The process by which compounds are formed is called a **chemical change** or **chemical reaction**.
- **A chemical reaction is a change which results in the formation of new substances.**

Examples of chemical reactions includes:

1. **Synthesis reactions:** They involve combination of two or more substances to form one compound. Examples of synthesis reactions includes

(i) Combustion

the burning of metals in air to new compounds called metal oxide

Eg

Magnesium metal burns in oxygen to produce magnesium oxide. This reaction can be written in a simple as

Magnesium + oxygen => magnesium oxide

2. **Decomposition reaction:** The chemical reaction which involves breaking down one single compound into two or more elements and compounds.

Examples includes

(i) Thermal decomposition of calcium carbonate into calcium oxide and carbon dioxide through heating.

Calcium carbonate => calcium oxide + carbon dioxide.

(ii) Electrolysis of acidified water by electrolysis

Water is split into hydrogen and oxygen gas by passage of electric current.

Water => hydrogen gas + oxygen gas

3. **Single replacement:** A chemical reaction in which one substance replaces another substance to form a new substance. Examples of single replacement reactions are

(i) acid-metal reactions. When metals are reacted with acids, a compound called salt is formed alongside hydrogen gas.

Eg

Magnesium reacts with hydrochloric acid to form magnesium chloride and hydrogen gas as shown below:

Magnesium + hydrochloric acid => magnesium chloride + hydrogen gas.

ELEMENTS

- **An element is a pure substance made up of same types atoms and cannot be decomposed into simpler substances by chemical changes**.
- Examples of elements includes **coppe**r **atom**, **sulphur atom**, **oxygen atom**, **hydrogen atom**, **iron atom**, **nitrogen atom** etc.

MIXTURES

- **A mixture is a physical combination of two or more pure substances in varied proportions.**
- In any mixture, the composition can be varied and each component of the mixture retains its own properties and may be separated into pure substances by physical methods.
- Examples of mixtures includes

(i) **Crude oil**: mixture of oil.
(ii) **Soil**: soil is a mixture of different types of substances.
(iii) **Air**: air is a mixture of gases such as oxygen, nitrogen, carbon dioxide, water vapour etc.
(iv) **Brass**: Brass is a solid mixture of copper and zinc
(v) **Sea water**: sea water is a mixture of water, salt and solid particles such as clay.
(vi) **Blood**: blood is a mixture of water. blood cells and other components.

- Mixtures are classified into **homogenous** and **heterogeneous** mixtures.
- **HOMOGENOUS MIXTURE**: A mixture in which components mix uniformly and are indistinguishable from each other. Examples of homogeneous mixtures includes air, brass, salt solution etc.
- **HETEROGENEOUS MIXTURE**: A mixture in which components do not mix uniformly and components are distinguishable.

Examples of heterogeneous mixtures includes cooking oil and water; water, stones and sand. The question is how can you tell whether a substance is pure or a mixture? **Criteria of purity is the answer**.

CRITERIA OF PURITY

- Criteria of purity is a method used to determine whether a substance is pure or a mixture.
- In the laboratory chemists determine the purity of a substance using the following methods:

1. Boiling and melting point of the substance.
2. Density of the substance.
3. Refractive index of the substance.
4. Chromatography.

- Every substance has a different pair of melting and boiling points, density, and refractive index.
- A pure substance has a definite and sharp melting and boiling point, fixed density, fixed refractive index and produces only one well-defined spot on a chromatogram.

	Boiling point	**Melting point**	**chromatogram**
Pure substance	Sharp	Sharp	Produces only one spot on a chromatogram.
Mixture	Boils over a temperature range and at a higher temperature than the pure liquid	Melts over a temperature range and at a lower temperature than the pure solid	Produces more than one spot.

IMPORTANCE OF PURITY OF SUBSTANCES

- To ensure that manufactured drinks, foodstuffs, medicine, fuels and many other products are pure so that they do not harm us or other things related to us.

- To identify counterfeit substances.

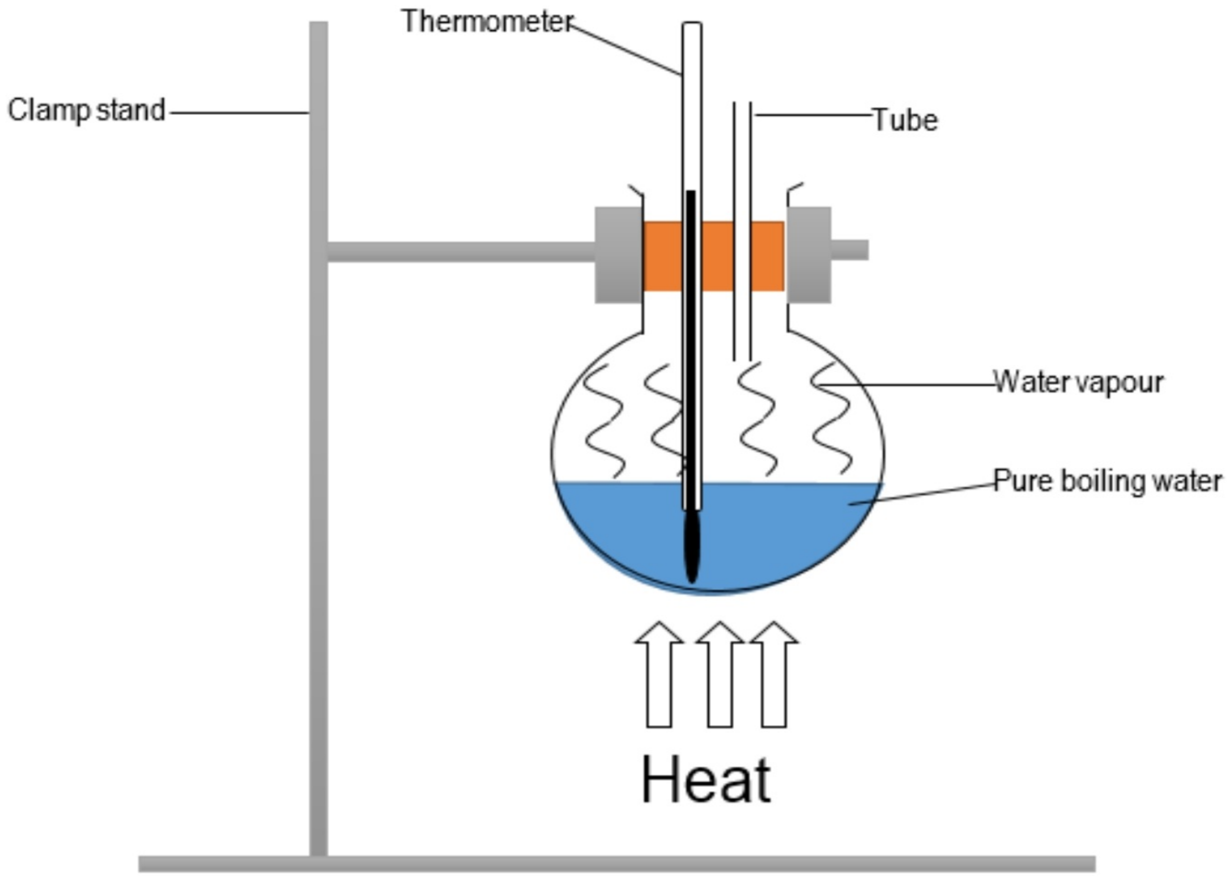
Experiment to determine the boiling point of pure water

Requirements:

Safety tube, round bottomed flask, laboratory thermometer, Bunsen burner, clamp stand and a rubber stopper distilled water, matches and graph paper.

Procedure:

- Distilled water is added in the round bottomed and clamped. The experiment is set up as shown in the diagram below.



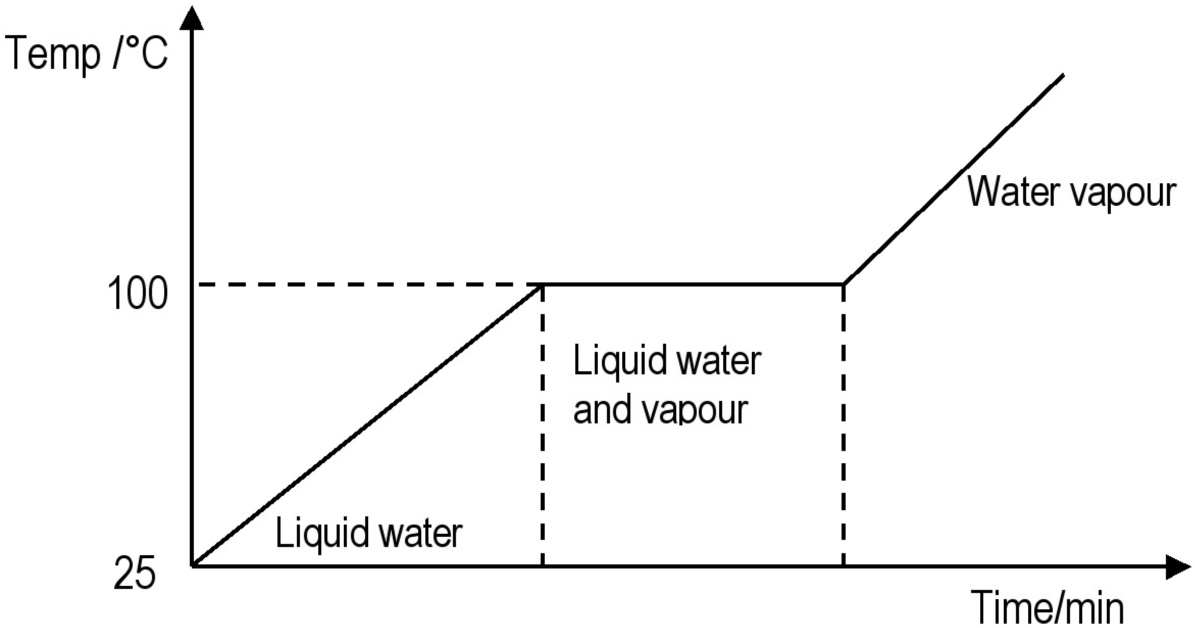
- Take the reading on the thermometer every after a minute and record your data in a table like the one shown below.

Temp °C	
Time /min	

- Plot the data in the table on the graph with temperature on the y-axis and time on the x-axis.

Observations:

- As water continues to be heated, the temperature continues to rise steadily from 25°C until water starts to boil at 100 °C. The temperature stays at 100 °C as water boils for several minutes.
- When data was plotted on the graph, it came out as shown below.



Water boiling point

Explanations:

- The temperature remained at 100°C as pure water boiled because pure water has a sharp and clear boiling point.

Conclusion:

- Pure substances have sharp and clear boiling points.

An experiment to investigate the effect of impurities on the boiling point of a pure water

Requirements:

Safety tube, round bottomed flask, laboratory thermometer, Bunsen burner, clamp stand, a rubber stopper, distilled water, table salt and matches.

Procedure:

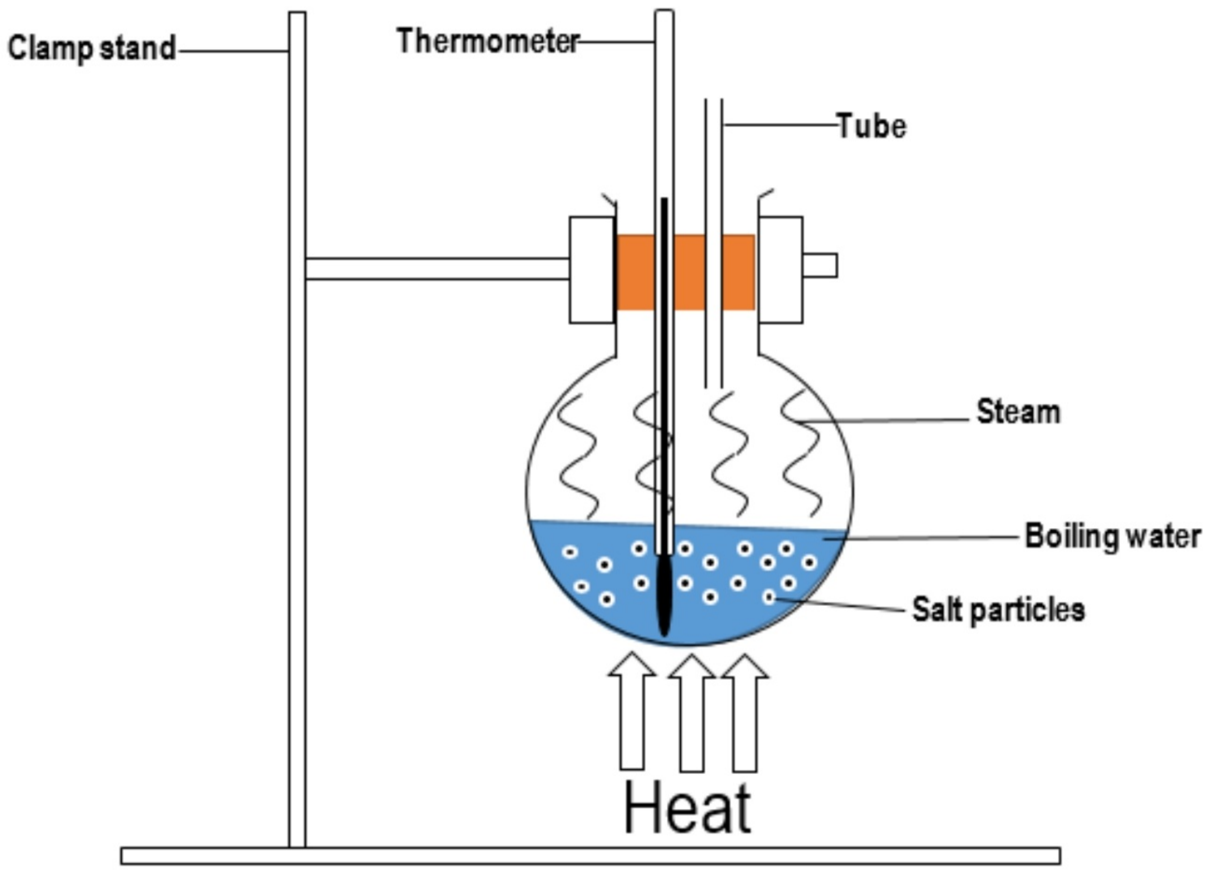
- Salt was added to the distilled water and stirred. The solution was put in the round bottomed and clamped. The whole experiment is set up as shown in the diagram below.
- Take the reading on the thermometer every after a minute and record you data in a table like the one shown below.

Temp °C									
Time /min	1	2	3	4	5	6	7	8	9

- Plot the data in the table on the graph with temperature on the y-axis and time on the x-axis.

Observation:

- As water continues to be heated, the temperature continues to rise steadily from 25°C until water starts to boil at 105 °C. The temperature stays at 105 °C as the water boils for several minutes.



- When temperature and time are plotted on the graph, the boiling point occurred over a temperature range and at a higher temperature than the pure liquid.

Explanation

- The increase in boiling point is caused by the addition of salt particles to distilled water because the impurities absorb some of the heat that is being supplied to the system.

Conclusion

- The boiling point of a mixture is higher than the boiling point of a pure substance.

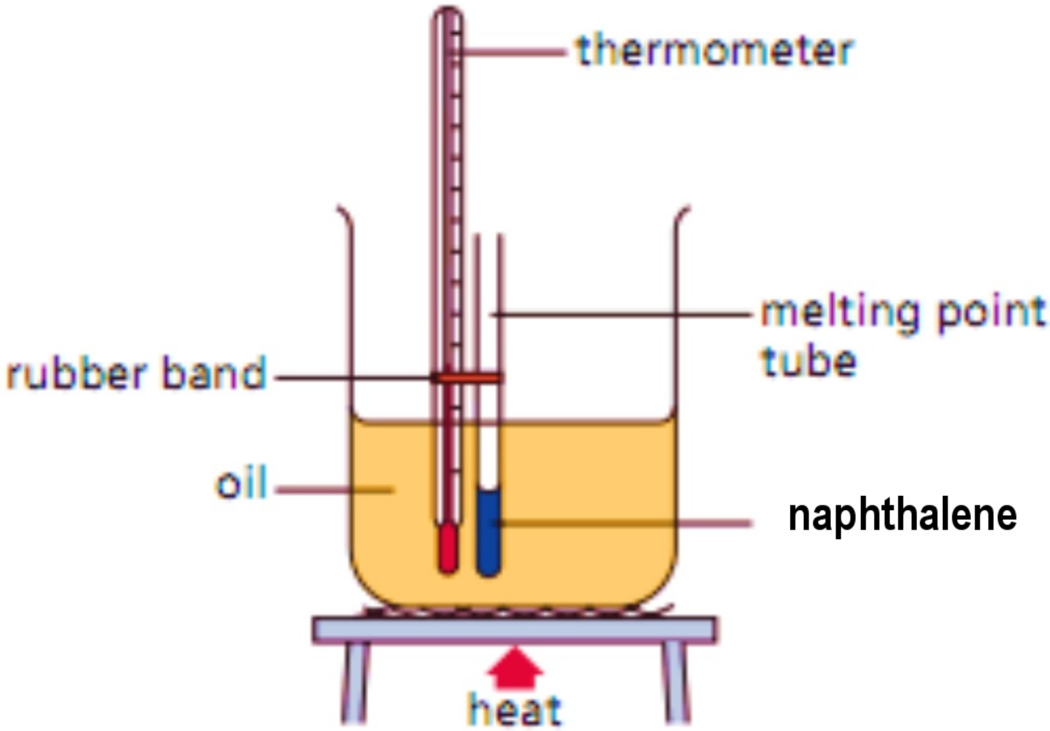
Experiment to determine the melting point of naphthalene

Requirements:

Naphthalene, boiling tube, thermometer, rubber band stopper, gas tube, beaker, water, Bunsen burner, and graph paper.

Procedure:

Solid or powdered naphthalene is put into a boiling tube the boiling tube and then attached to the bulb of the thermometer by a rubber band. The whole experiment is set up as shown below.



- Take the reading on the thermometer every after a minute and record your data in a table like the one shown below.

Temp °C									
Time /min	1	2	3	4	5	6	7	8	9

Observations

- As water continues to be heated, the temperature on the thermometer rises from 25 ^{o}C to 46 ^{o}C, to 62 ^{o}C to 71 ^{o}C and comes to remain steady at 80 ^{o}C. At 80 ^{o}C naphthalene starts to melt. The temperature remains constant at 80 ^{o}C until all the naphthalene has melted. When all the naphthalene has melted the temperature starts to rise again.

Explanation

- The constant temperature at which naphthalene starts to melt is the melting point of naphthalene because melting point of a pure substance is sharp and steady.

Conclusion

- In this experiment, the melting point of naphthalene is found to be 80 ^{o}C.

Application in life:
We put salt in beans after it has cooked and we put salt in meat just before starting to cook. Explain why?

SEPARATING MIXTURES

- In chemistry most substances are required and used in pure form. However, most substances exist as mixtures.
- We can easily separate pure substances from mixtures using physical methods because each component of the mixture retains its own physical properties such as;

1. Different solubility
2. Different melting and boiling points.
3. Different densities

- The physical methods used to separate pure substances depends on these physical properties of pure substances.
- Physical methods are examples of **physical changes**, this means that they do not result into formation of new substances and they are reversible. Examples of physical changes are:

1. **Evaporation:** which happens when liquids are heated,
2. **Melting:** which happens when solids such as ice are heated,
3. **Filtration:** which involves separating insoluble solids from a solution using a filter paper,
4. **Sublimatio**n: which involves cooling or heating certain substance,
5. **Solidification:** which happens when liquids are cooled.
6. **Condensation:** which happens when gases are cooled.

Differences between physical and chemical changes

Physical change	Chemical change
No new chemical substance is formed	A new substance is formed
The change is reversible	The change is not easily reversible
The mass of the original substance does not change	The mass of the original substance changes
No energy is involved	Energy is involved

METHODS OF SEPARATING MIXTURES

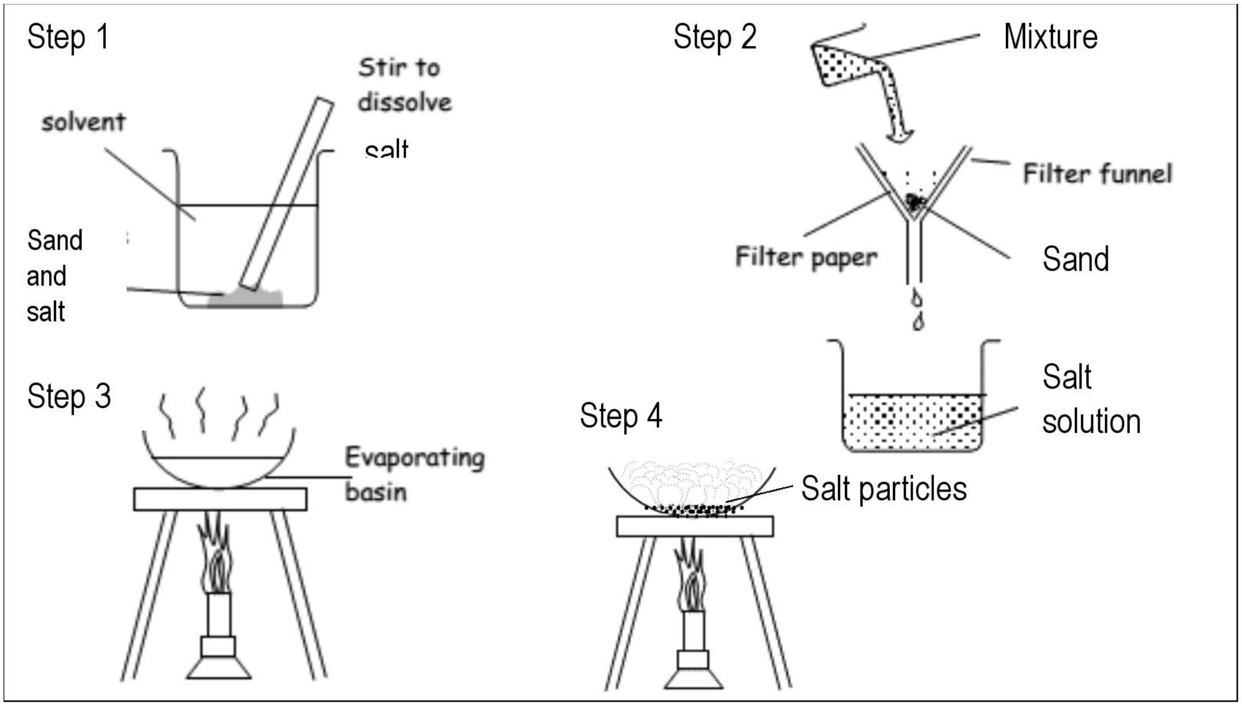
The following are the physical methods for separating mixtures.

1. Filtration, evaporation and crystallization.
2. Simple distillation
3. Fractional distillation
4. Chromatography

Filtration, evaporation and crystallization

This method is used to separate two solids from a solution in which one is soluble.

- This method can be used to separate a mixture of common salt and sand.
- The mixture of salt and sand is first dissolved in water, then filtered to obtain sand, then evaporated and finally crystallized to collect salt as shown below;

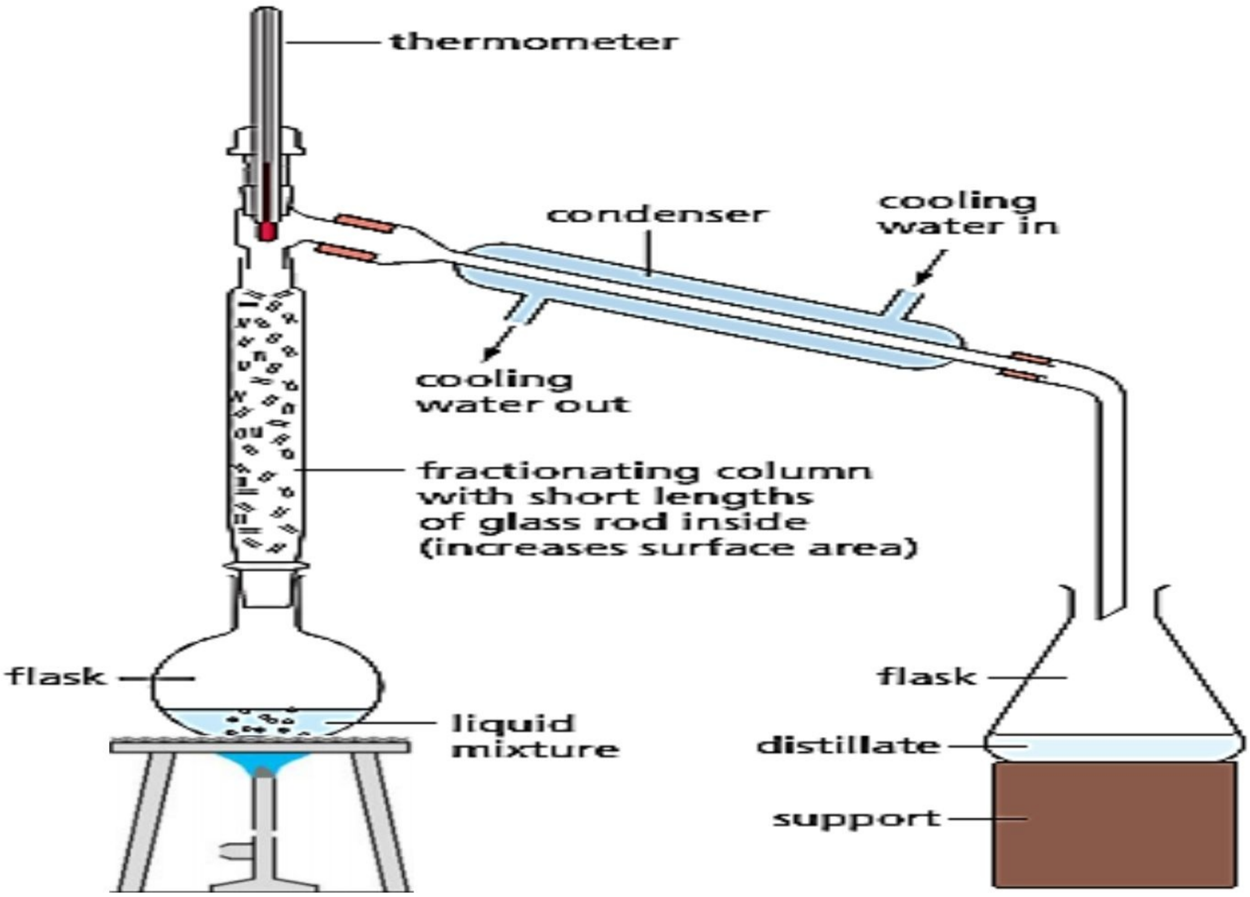


- This method can also be used to separate a mixture of copper (II) sulphate and sand.

FRACTIONAL DISTILLATION

- Fractional distillation is a method used to separate a mixture of two or more **miscible** liquids which boils at different temperatures.
- Components in the mixture can be separated by fractional distillation because they are miscible and have different boiling points.
- This method can be used to separate a mixture of water which boils

at **100 ^oC** and ethanol which boils at **78 ^oC** which mixes uniformly. The apparatus used is set up as shown below;



- The mixture of ethanol and water is heated so that the vapours of ethanol and water boil off at different temperatures and be condensed and collected separately at different temperatures with ethanol boiling out first at 78 °C whereas water boils at 100 °C.
- The portions of liquids collected differently are called **fractions**, this is why the method is called **fractional distillation**.

Functions of the parts used in fractional distillation

- The thermometer is used to check the temperature of the fraction being distilled
- The glass beads condense the fractions whose boiling point has not yet reached so that they drop back into the distillation flask.
- The fractionating column offers a large surface area for condensing less volatile vapours.

Application of fractional distillation

Fractional distillation is used;

1. To separate pure oxygen and pure nitrogen from liquefied air,
2. To separate substances in petroleum (crude oil) into fractions (petrol, paraffin etc.), and
3. In the production of ethanol.

CHROMATOGRAPHY

Chromatography is a separation technique used to separate complex mixtures such as dyes in ink and chlorophyll, amino acids in proteins, monosaccharide in starch, poisons in blood etc.

We are able to separate components in a mixture by chromatography because components have different solubility in a given solvent.

How chromatography is carried

A mixture whose components is to be separated is put on a material called **stationary phase** which can be a filter paper. A stationary phase with a mixture on it is dipped into a solvent called **mobile phase** which can be ethanol etc. If the stationary phase is a paper, then the chromatography is described as **paper chromatography** and if the separation takes place on a solid it is described as **solid chromatography.**

The components of the mixture dissolve in the solvent and moves and separate in different components on a stationary phase. The mixture separates in components because each component of the mixture sticks to the stationary phase at different intervals due to their different **solubility** in the solvent and **densities**. The **less soluble solute sticks closer to the starting point** while the **more soluble ones stick near the end of the stationary phase**. The components move in the stationary phase by **capillary property**. The separated components of the mixture as they appear on a stationary phase are called **chromatogram**.

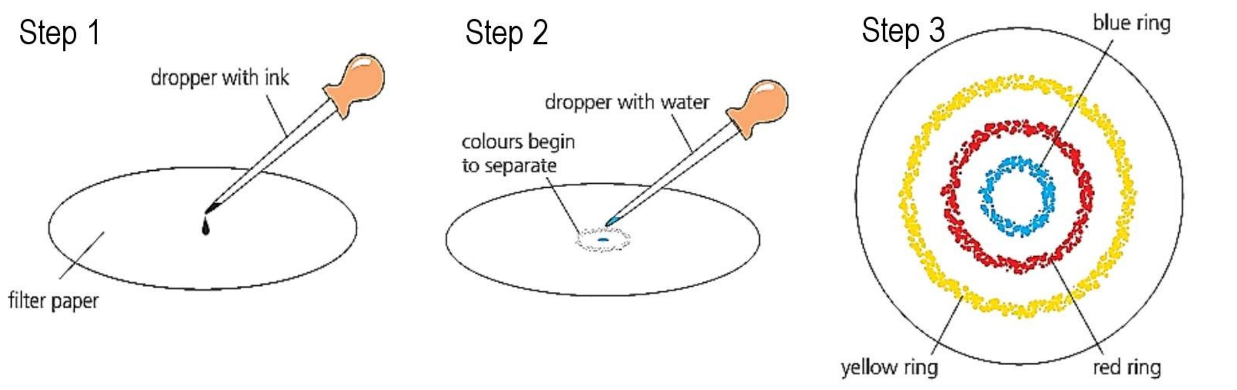
Types of paper chromatography

There are three types of chromatography namely

1. **Radial paper chromatography**
2. **Ascending paper chromatography**
3. **Descending paper chromatography**

Radial paper chromatography

The type of paper chromatography in which the stationary phase is circular. In this type of paper chromatography, the stationary phase is placed horizontally and a spot of mixture to be separated is place at the centre of stationary phase. Then drops of mobile phase are added as shown below:



1. Place a drop of black ink in the centre of some filter paper. Let it dry. Then add three or four more drops on the same spot, in the same way.

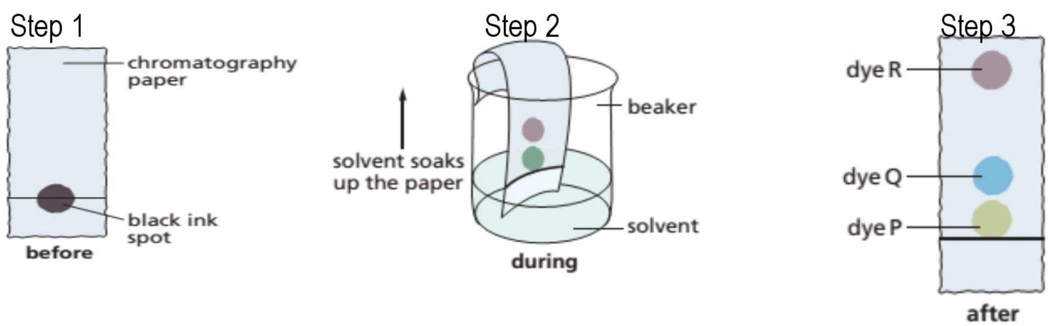
2. Now drip water onto the ink spot, one drop at a time. The ink slowly spreads out and separates into rings of different colours.

3. Suppose there are three rings: yellow, red and blue. This shows that the ink contains three dyes, coloured yellow, red and blue.

- The advantage of using this method is that gravity has little effect on the movement of components.

Ascending paper chromatography

- The type of paper chromatography where the stationary phase is placed vertically. In this type of paper chromatography, the mobile phase moves up the stationary phase by **capillary action.**
- To separate the different coloured dyes in a sample of black ink using ascending paper chromatography, a spot of the ink is put onto a pencil line drawn on the filter paper, then the paper is set in a suitable solvent as shown in the figure below;



- The dye will dissolve in solvent and travel up the paper at different speed and be separated. The chromatogram in step 3 shows how the ink contains three dyes, P, Q and R.

Advantage;

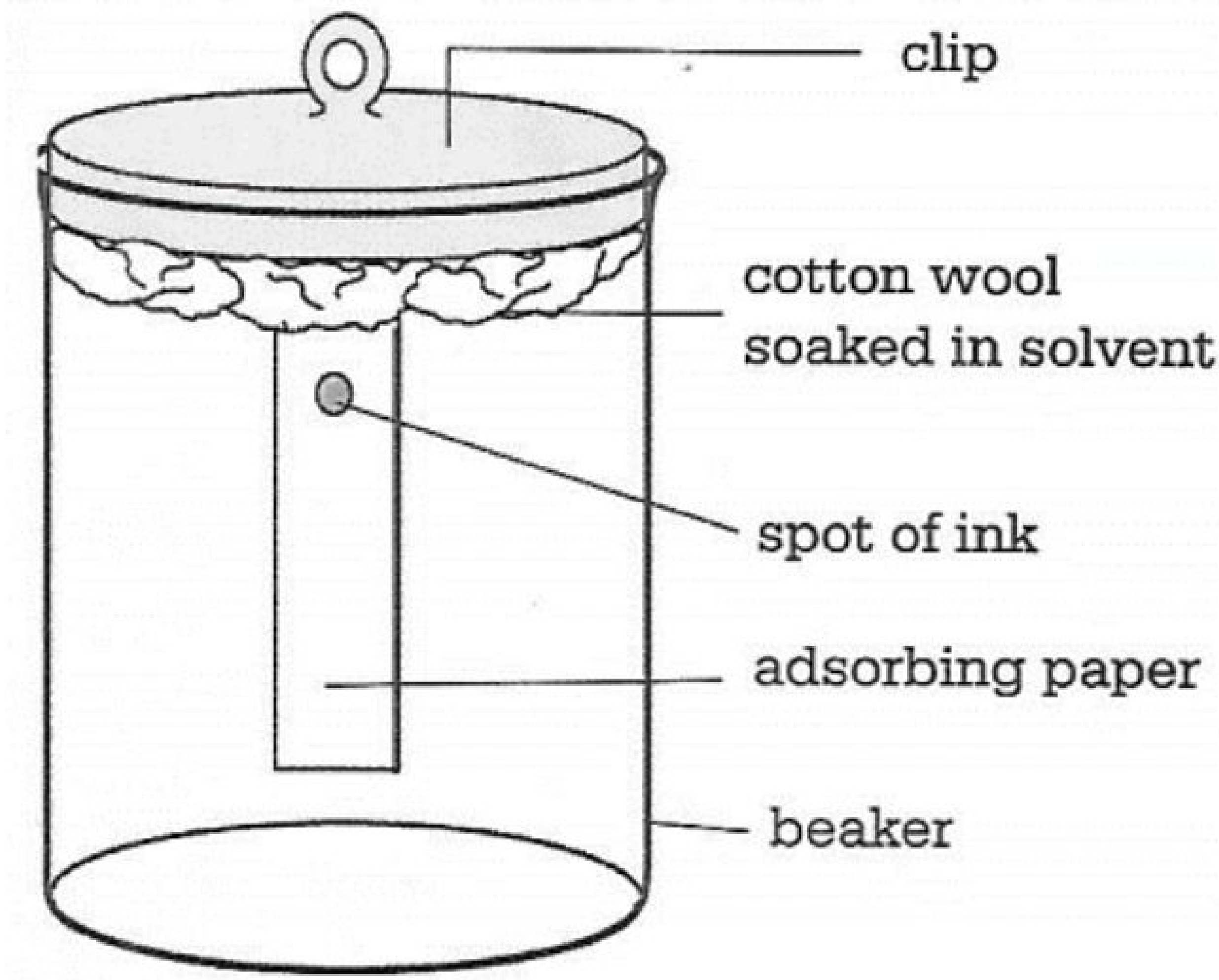
- Colours are clearly seen and distinct

Disadvantage;

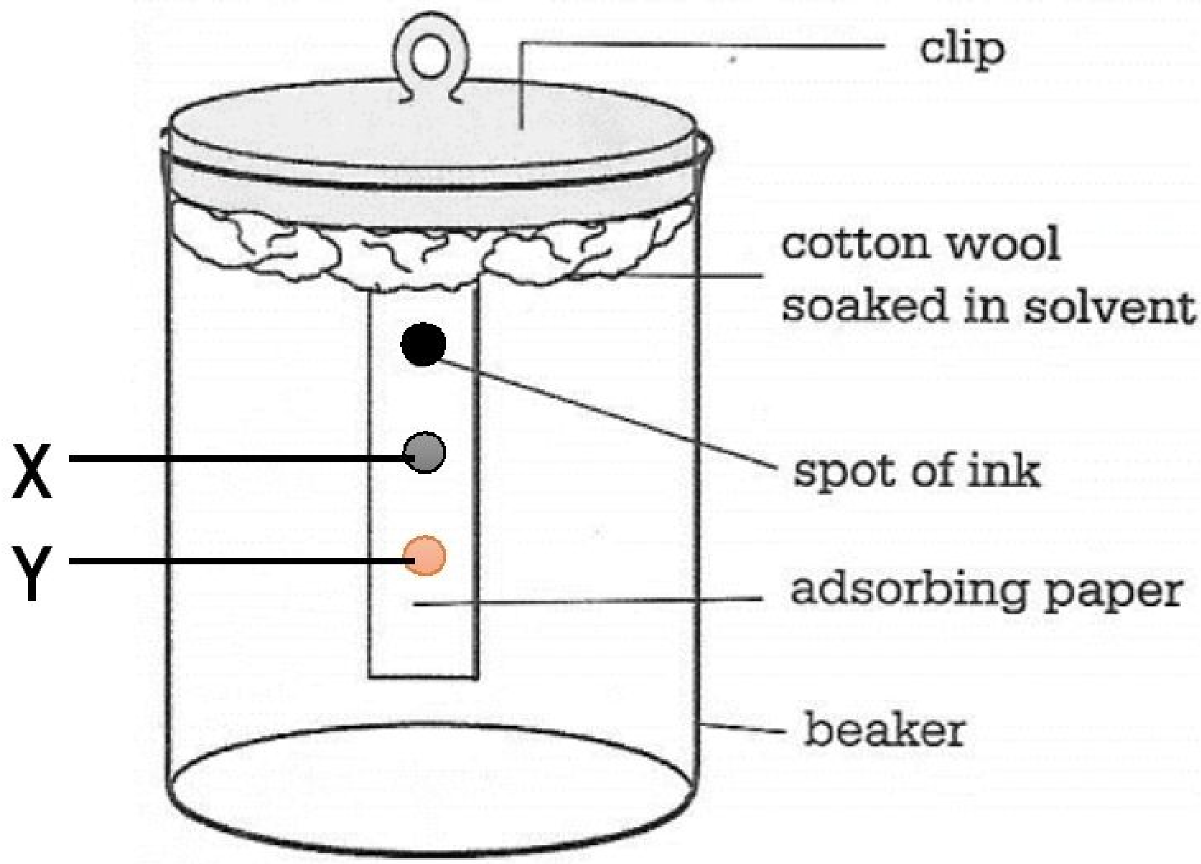
- It takes time for colours to be seen as solvents move against gravity.

Descending paper chromatography

- The type of paper chromatography where the stationary phase is placed vertically. In this type of paper chromatography, the mobile phase moves downwards the stationary phase by **capillary action and pull of gravity.**
- To separate the different coloured dyes in a sample of black ink using descending paper chromatography, a spot of the ink is put onto a pencil line drawn on the filter paper, then the paper is fixed to cotton wool soaked in a solvent as shown in the figure below;



- The solvent will travel downwards dissolving the black ink and separating the components of the ink in the process. The chromatogram in the figure below shows how the ink contains two dyes, X and Y.



Advantage of descending paper chromatography;

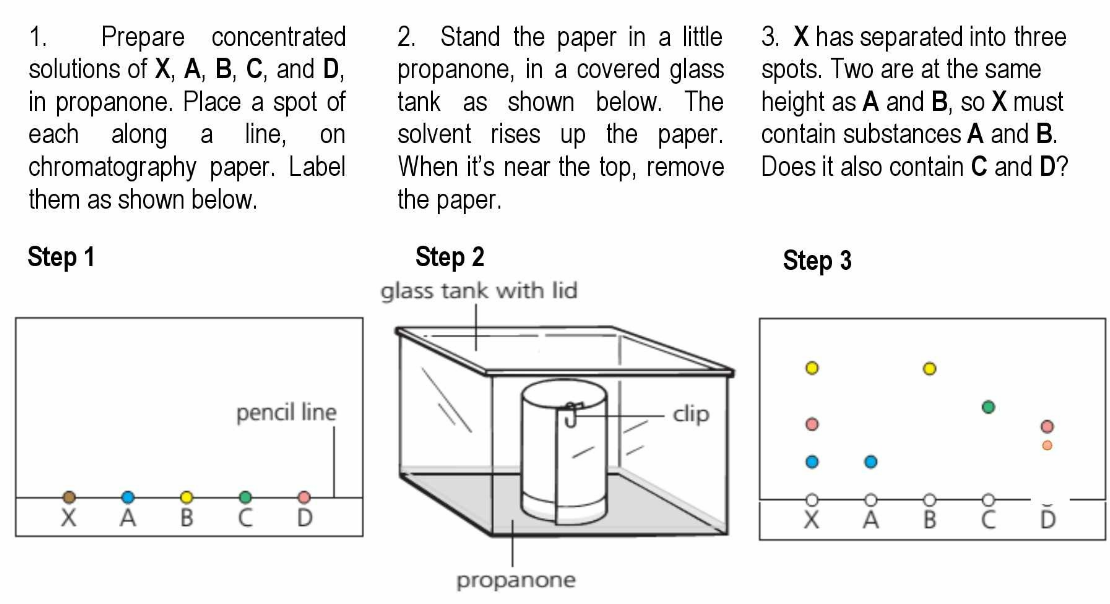
- Colours form fast due to the pull of gravity.

Disadvantage of descending paper chromatography;

- Component colours may overlap making it difficult to distinguish them.

INTERPRETATION OF RESULTS ON

CHROMATOGRAPHY



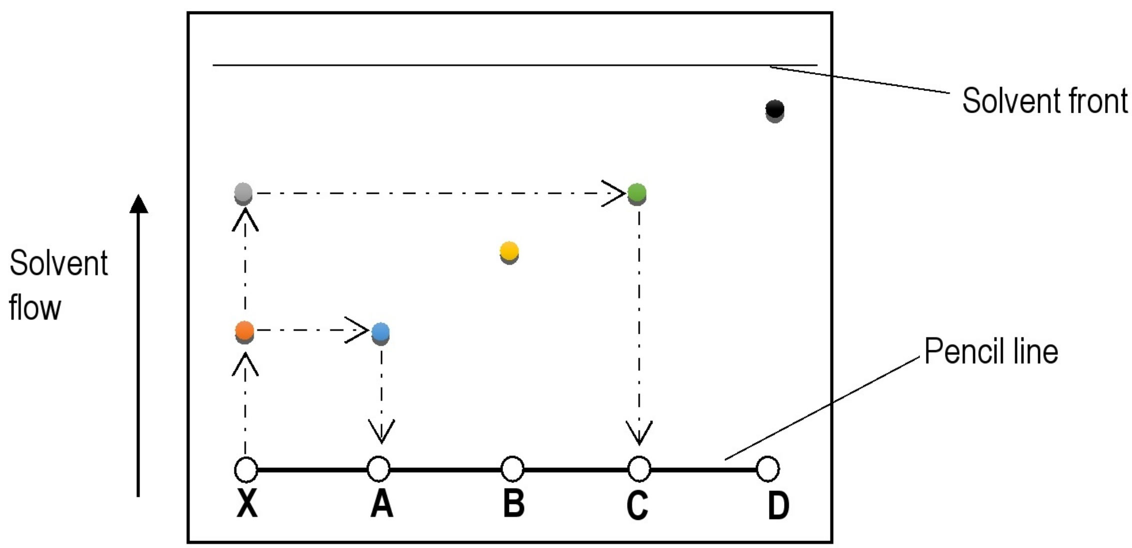
The following experiment was carried to identify mixtures of coloured substances in dyes X. For example, mixture **X** is thought to contain substances **A**, **B**, **C**, and **D**, which are all soluble in propanone. You could check the mixture like this:

The following steps are followed to interpret the results

(i) From unknown dye (i.e. **X in the figure below)**; follow direction of solvent flow until you find a spot.

(ii) Move across the chromatogram until you find a corresponding spot.

(iii) Move in the other direction to identify the known dye as shown below.

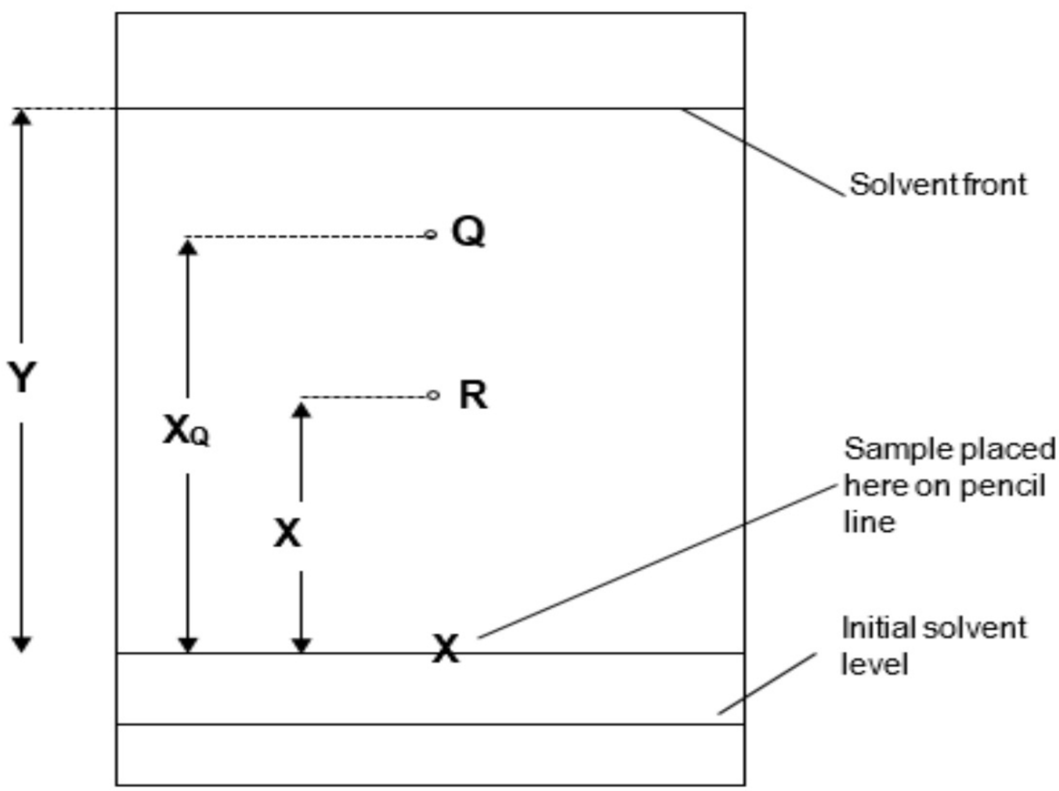


Unknown Dye X = Mixture of Dyes A + C

- The dyes at the same distance from pencil line with the unknown are of the same composition (or are part of the unknown). Hence dye X contains mixture of dyes A and C.

IDENTIFYING THE COMPONENTS OF A MIXTURE USING RETARDATION FACTOR (Rf)

- $$R_f = \frac{\textit{Distance travelled by solute } (X)}{\textit{Distance travelled by solvent } (Y)}$$



R_f of solute Q is:

$$R_f = \frac{\text{distance travelled by Q}}{\text{distance travelled by solvent}}$$

$$R_f = \frac{\boldsymbol{X_Q}}{\boldsymbol{Y}}$$

R_f of solute R is:

$$\mathbf{R_f} = \frac{\text{distance travelled by R}}{\text{distance travelled by solvent}}$$

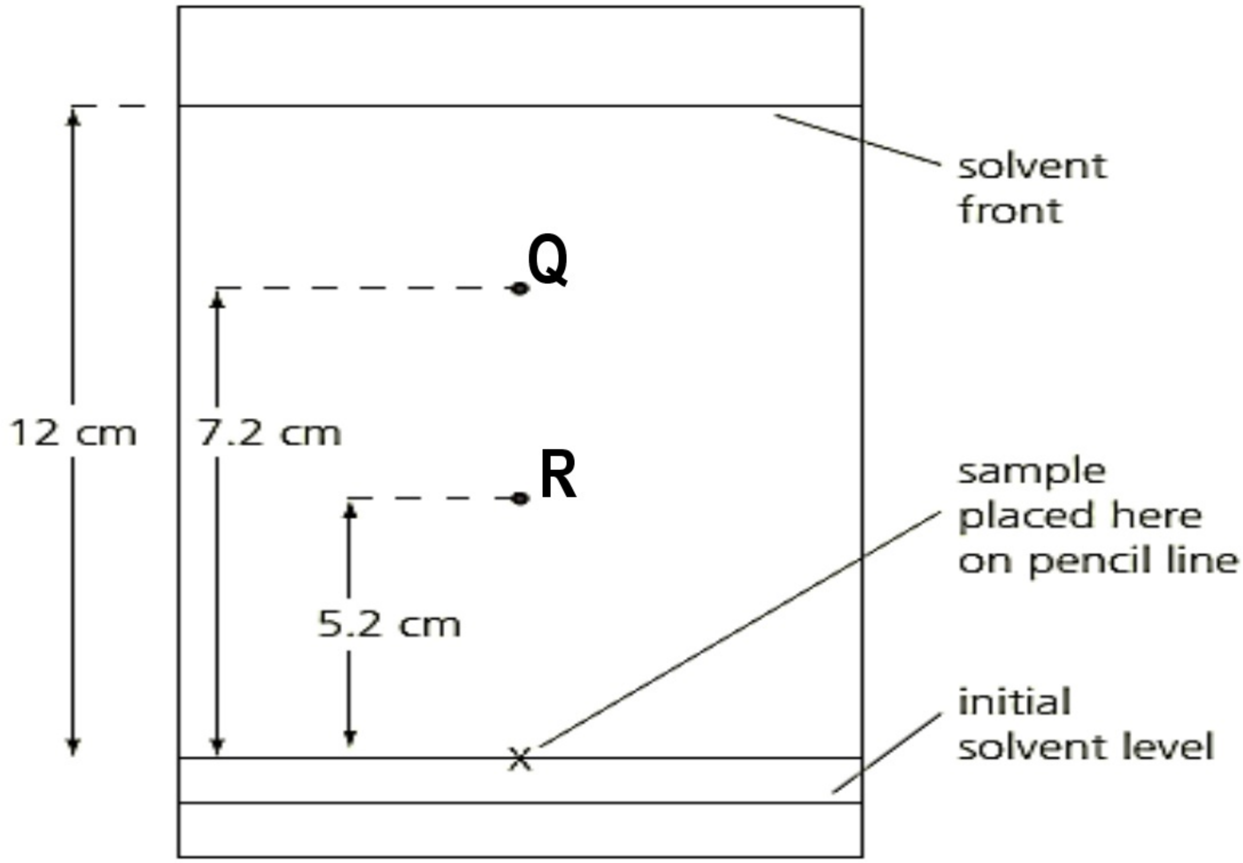
$$\mathbf{R_f} = \frac{\boldsymbol{X_R}}{\boldsymbol{Y}}$$

The table below gives R_f values for amino acid when water, butanol, and ethanoic acid are used as solvents:

Amino acid	R_f values
cysteine	0.08
lysine	0.14
glycine	0.26
serine	0.27
alanine	0.38
proline	0.43
valine	0.60
leucine	0.73

AN EXAMPLE ON HOW TO USE THE RF FORMULA

The diagram below shows a chromatogram for a mixture of amino acids.



The solvent was a mixture of water, butanol, and ethanoic acid.

(a) Using the table of R_f values on the table above, identify the two amino acids.

(b) Which of them is less soluble in the solvent?

(c) How will the R_f values change if the solvent travels only 6 cm?

SOLUTIONS

(a) R_f for Q

$$R_f = \frac{7.2}{12} = 0.6$$

Q is valine.

R_f for R

$$R_f = \frac{5.2}{12} = 0.43$$

R is proline

(b) **Proline**

(c) $R_f = \frac{7.2}{6} = 1.2$ for Q, $R_f = \frac{5.2}{6} = 0.6$ for R

Uses of chromatography;

1. Used to separate and identify monomers in foods.
2. Used to detect and identify poison in blood.
3. Used to detect and identify various inks used in forged documents.
4. Used to determine the purity of substances.

THE COMPOSITION OF MATTER

- Matter is anything that occupies space and has mass.
- According to kinetic theory of matter, all matter is made up of tiny particles that are in continuous motion. These tiny particles are called basic units of matter.

❖ These includes:

1. Atoms,
2. Ions, and
3. Molecules.

- Most atoms can combine with other atoms to form many different substances which make up matter. There are over 100 different types of atoms.

- The important groups of substances which make up matter are;

☐ Elements,
☐ Molecules,
☐ Compounds, and
☐ Mixtures.

ATOMS

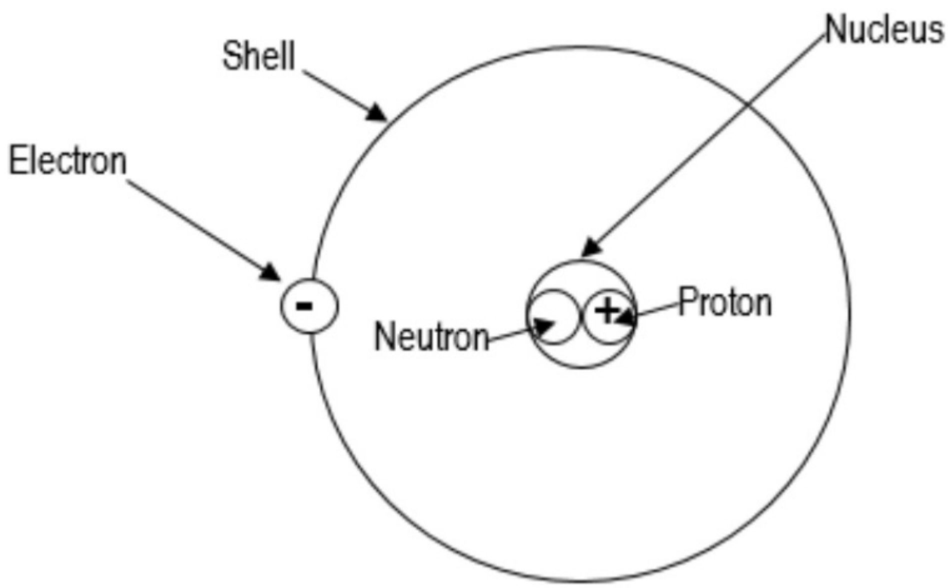
- An atom is the smallest particle of an element that can take part in a chemical reaction and remain unchanged.
- Atoms are the smallest particles of matter that we cannot break down further by chemical means.

Composition of atoms

The structure of an atom

- Atoms are made up of three kinds of fundamental particles or sub-atomic particles namely **electrons**, **protons** and **neutrons**.

- An atom has two separate regions where electrons, protons and neutrons are found. These regions are called the **nucleus** and the **energy levels** or **shells**.
- The nucleus is the central part of an atom while the shell is the area around the nucleus.



- **Protons** and **neutrons** are found in the nucleus while electrons are found in shells where they orbit (or move around) the nucleus.

- The circular paths surrounding the nucleus are called energy levels or shells.
- The electrons are held within shells of the atom by an **electrostatic force** of attraction between themselves and the positive charge of protons in the nucleus.
- The nucleus occupies only a very **small volume** of the atom but is very dense.

Properties of sub-atomic particles

Some of the particles contained in an atom are electrically charged. These are protons and electrons.

- Protons are positively charged (+1) while electrons are negatively charged (-1).
- Neutrons have no charge; they are neutral hence their name.
- Every atom has an equal number of protons and electrons. So atoms have no overall charge. This is because the electrical charge of the protons in the nucleus is balanced by the opposite charge of the electrons in the shells.
- The approximate masses of protons, neutrons and electrons together with their charges are shown in the table on the next page.

Particle	Symbol	Relative mass/amu	Relative charge
Proton	p	1	+1
Neutron	n	1	0
Electron	e	1/1837	−1

- The mass of an atom is taken to be the mass of the nucleus. The mass contributed by the electrons is negligible.
- As shown by the table, the mass of one electron is far too small compared to the mass of either the proton or the neutron. Therefore, for all practical purposes the mass of an atom is taken to be the mass of the nucleus.
- When determining the mass of an atom, we simply add the neutrons and the protons. i.e. if the atom has 11 protons and 10 neutrons then its mass is 21.
- The volume of an atom is taken to be the space occupied by the electrons around the nucleus. The volume of the nucleus is so small that it contributes almost nothing to the total volume of an atom.

ATOMIC OR NUCLIDE NOTATION

- All the information needed about an element is found inside the nucleus of an atom of that particular element, therefore, we use the nucleus to provide all the information about an atom of a particular element.

☐ The short way of providing information about an atom of an element using the information found in the nucleus of that atom is called **atomic or nuclide notation**.

- We use an atom's symbol, atomic number and mass number in nuclide notation.
- The nuclide notation of atoms of elements is represented as shown below

:

MASS NUMBER

X is the **atomic or chemical symbol**, derived from the name of the element. It can be one or two letters from an element's name.

$$^{A}_{Z}X$$

ATOMIC OR PROTON NUMBER

MASS NUMBER (A)

- **Mass number** is represented by letter **A** in nuclide notation.
- Mass number is the total number of protons and neutrons in an atom. It is also known as the **nucleon number**.

ATOMIC NUMBER (Z)

- **Atomic number** is represented by letter **Z** in nuclide notation.
- Atomic number is the number of protons in the nucleus of an atom. Atomic number is also **called proton number**.
- All atoms of a particular element have the same number of protons, and an element contains only one kind of atom.
- In a neutral atom the number of protons is equal to the number of electrons. If the number of protons (atomic number) is 11 then the number of electrons is also 11.

☐ In nuclide notation, a sodium atom is written as shown below:

$^{23}_{11}Na$, **Na** is the chemical symbol for sodium element, **11** is the **atomic number** or **proton number** for sodium atom and **23** is the **mass number** or **nucleon number** for sodium atom.

☐ All the atoms of elements are written in nuclide notation on the periodic table. See the periodic table on the last page.

THE RELATIONSHIP BETWEEN MASS NUMBER, ATOMIC NUMBER AND NUMBER OF NEUTRONS

- The mass number (A), atomic number (Z) and number of neutrons are related by the following equation:

Mass number (A) = atomic number (Z) + number of neutrons (N)

A = Z + N

- We use this equation to calculate the number of neutrons in an atom. For example, we can calculate the number of neutrons in a lithium atom given below.

$$^{7}_{3}Li$$

In this atom, mass number (A) = 7, atomic number (Z) = 3, number of neutron (N) can be calculated as follows:

A = Z + N

7 = 3 + N

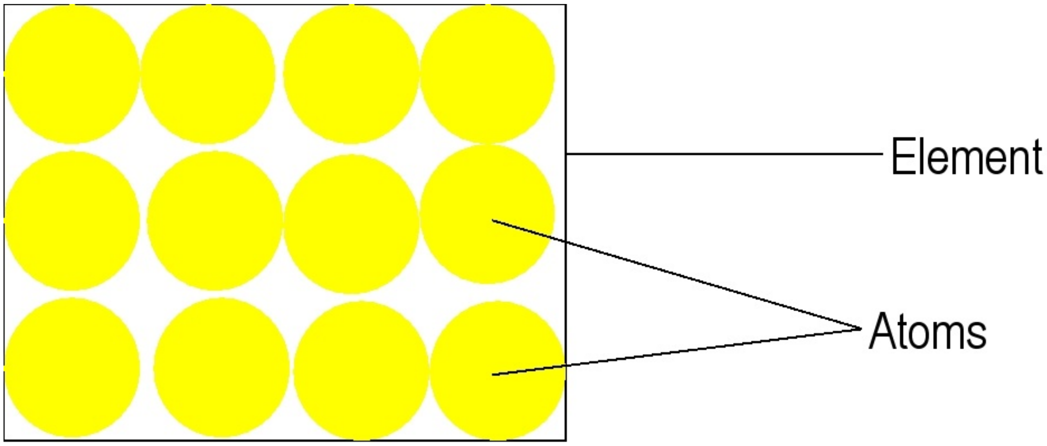
7 – 3 = N

4 = N $\therefore$ **N = 4**. Thus there are 4 neutrons in lithium atom.

ELEMENTS

- An element is a pure substance that cannot be broken down into other substances through chemical means.
- An element is made up of atoms with the same atomic number. The smallest unit of an element is the atom. There are 109 known elements.

Structure of an element



- All the elements we know today are recorded in the Periodic Table of the Elements in **nuclide notation.**

Chemical symbols of common elements

- Symbols of elements found on the periodic table were derived from the first letter or first two letters of the element’s name. Two letters were used to write an element’s symbol for elements whose names were in Latin to avoid the use of the same symbol.
- An element’s symbol is written in capital letter but in situation where an element has two letters in its symbols, only the first letter is written in capital. The element iron, for example, has the chemical symbol, Fe. This is because the elements were originally given Latin names. Iron has the symbol Fe because its Latin name is ’ferrum’. In the same way, sodium’s Latin name is ’natrium’ (Na) and gold’s is ’aurum’ (Au).
- The table below shows the first twenty elements and their symbols together with copper and iron.

						Physical

Element	Symbol	Proton number	Electrons	Neutrons	mass number	state at room temperature and pressure
Hydrogen	H	1	1	0	1	gas
Helium	He	2	2	2	4	gas
Lithium	Li	3	3	4	7	gas
Beryllium	Be	4	4	5	9	gas
Boron	B	5	5	6	11	gas
Carbon	C	6	6	6	12	solid
Nitrogen	N	7	7	7	14	gas
Oxygen	O	8	8	8	16	gas
Fluorine	F	9	9	10	19	gas
Neon	Ne	10	10	10	20	gas
Sodium	Na	11	11	12	23	solid
Magnesium	Mg	12	12	12	24	solid
Aluminium	Al	13	13	14	27	solid
Silicon	Si	14	14	14	28	solid
Phosphorus	P	15	15	16	31	solid
Sulphur	S	16	16	16	32	solid
Chlorine	Cl	17	17	18	35	gas
Argon	Ar	18	18	22	40	gas
Potassium	K	19	19	20	39	solid
Calcium	Ca	20	20	20	40	solid

BRIEF HISTORY OF THE DEVELOPMENT OF THE PERIODIC TABLE

- The need to classify elements of similar physical and chemical properties in the same groups motivated the development of the periodic table.
- **Johann W. Dobereiner** tried to arrange elements with similar properties into sets called triad using atomic mass. Examples of the triad groups he came up with are chlorine, bromine, iodine and

lithium, sodium, potassium.

- Another chemist by the name of **John Newlands** arranged the 56 known elements in order of increasing atomic weight**.** He realised when he did this that every eighth element in the series was similar.

H **Li** Be B C N O F **Na** Mg Al Si P S Cl **K**.

- He likened this to music and called it the **'Law of Octaves'**. It fell down, however, because some of the weights were inaccurate on some elements that had not been discovered then.
- Another chemist by the name of **Mendeleev** classified all the 63 known elements in order of increasing atomic weight but in such a way that elements with similar properties were in the same vertical column. He called the vertical columns **groups** and the horizontal rows **periods** as show by the table below.

Groups

Periods

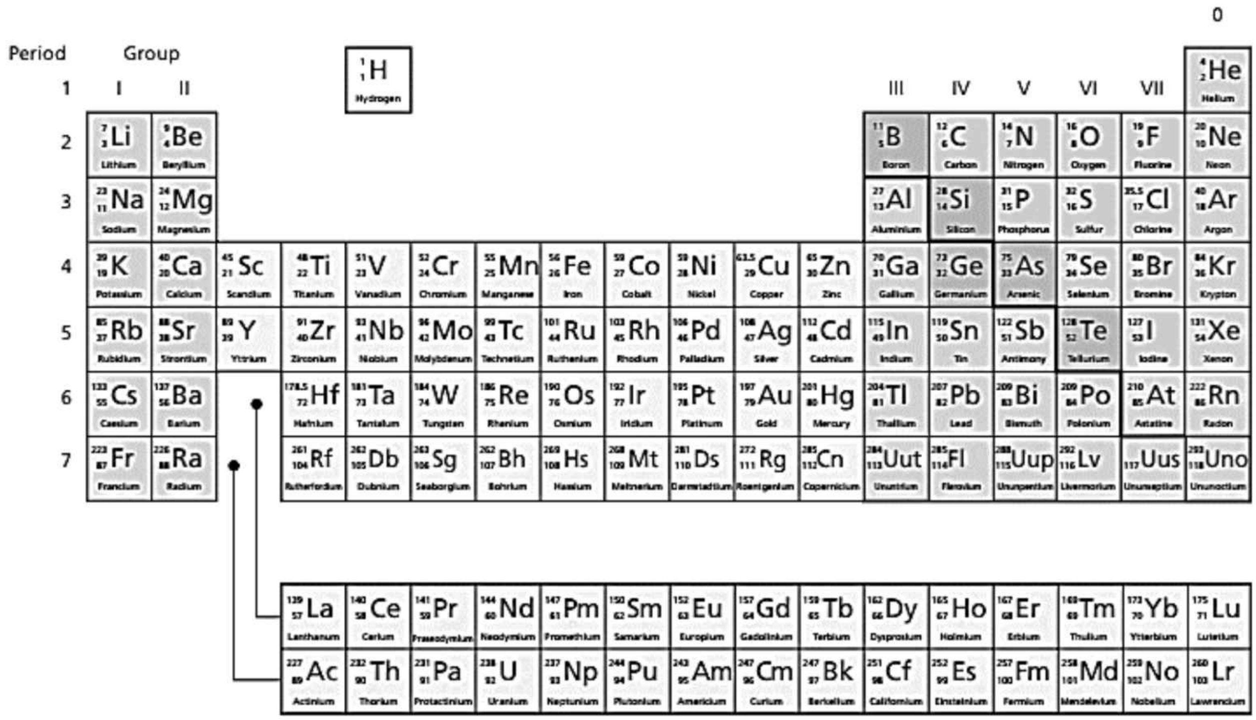
	1	2	3	4	5	6	7	8
1	H							
2	Li	Be	B	C	N	O	F	
3	Na	Mg	Al	Si	P	S	Cl	
4	K Cu	Ca Zn	* *	Ti †	V As	Cr Se	Mn Br	Fe Co Ni

- He left gaps in the table for elements that had not been discovered then. As a scientific idea, Mendeleev's periodic table was tested by making predictions about elements that were unknown at that time but could possibly fill the gaps. Three of these gaps are shown by the symbols * and † in the table. As new elements were discovered, they were found to fit easily into the classification. For example, Mendeleev predicted the colour, density and melting point as well as atomic weight of the missing element ‘eka-silicon’ (†).

- The success of Mendeleev’s predictions showed that his ideas were probably correct. His periodic table was quickly accepted by scientists as an important summary of the properties of the elements.
- Rutherford and Moseley modified Mendeleev’s periodic table into the modern periodic table by arranging the elements in order of increasing **proton number** or **atomic number**. In the modern periodic table, those elements with similar chemical properties are found in the same columns or groups.

THE MODERN PERIODIC TABLE

- A modern periodic table is an arrangement of elements according to increasing atomic numbers.
- By arranging elements using atomic numbers, a more orderly table is obtained with vertical columns containing elements with similar chemical properties and horizontal rows containing elements with the same number of shells in their atoms.
- The vertical columns of elements are called groups while the horizontal rows of elements are called periods.
- There are eight groups and seven periods in the modern periodic table as shown below. The modern Periodic Table contains 118 known elements.



ISOTOPES

- ✓ At the beginning of understanding of atoms of elements, it was assumed that atoms of the same element were similar in every way, i.e. that they have the same atomic number and mass number. However, this was disproved by experiments carried out in an instrument called **mass spectrometer**.
- ✓ It was shown that atoms of the same elements were not similar in every aspect, instead they differed in terms of **mass numbers** due to variation in the number of **neutrons** in the nucleus and such atoms were called **isotopes**.
- ✓ **Isotopes are atoms of the same element that have the same atomic number but different mass numbers or same number of protons but different numbers of neutrons.**
- ✓ In reality most atoms of elements found on the periodic table exist as two or more isotopes but they are represented by one atom carrying one mass number called **relative atomic mass**.

RELATIVE ATOMIC MASS (Ar)

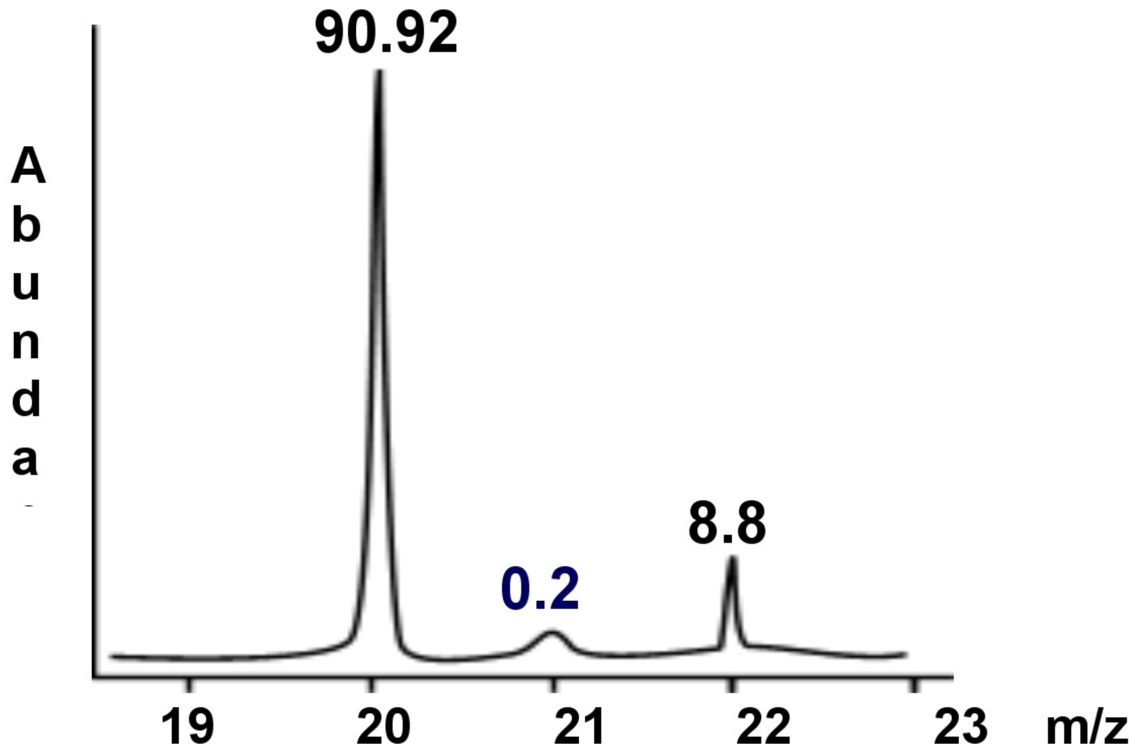
- ✓ Relative atomic mass of an element is the mass of naturally occurring isotopes of an element relative to the mass of an atom of carbon-12 isotope.
- ✓ It can be defined mathematically as:

$$A_r = \frac{\text{average mass per atom of an element}}{\text{mass of one atom of carbon} - 12} \times 12$$

- ✓ Relative atomic masses of atoms of elements were determined by an instrument called **mass spectroscopy**.
- ✓ A sample of an isotopic element whose relative atomic mass number is to be determined is placed in a mass spectrometer and the abundance of each isotope present in the sample produced as percentages. The following examples show calculations involved in determining relative atomic mass numbers.

Example 1

An early application of mass spectroscopy was the demonstration by Aston, (Nobel Prize, 1922), that naturally occurring neon consisted of 3 isotopes... ^{20}Ne ^{21}Ne ^{22}Ne.



Calculate the average relative atomic mass of neon using the above information.

Answer:

Relative atomic mass number is equal to the sum of the products of the percentages and their mass numbers of the isotopes.

$$\mathbf{A_r = \Sigma\, f_1 \times A_1 + \Sigma\, f_2 + A_2 + \ldots\ldots}$$

Where:

A_r = relative atomic mass number,

Σ = summation,

F = fraction abundance of isotopes, and

A = mass number of isotopes.

Thus relative atomic muss number of neon is

A_r = 90.92% × 20 + 0.2% × 21 + 8.8% × 22

$$A_r = \frac{90.92 \times 20}{100} + \frac{0.2 \times 21}{100} + \frac{8.8 \times 22}{100}$$

A_r = 20.18

Example 2

There are two common isotopes of chlorine as shown in the table below. Calculate the relative atomic mass of chlorine atom.

	Protons	Neutrons	%
$^{35}_{17}Cl$	17	18	75
$^{37}_{17}Cl$	17	20	25

Answer:

$$A_r = \frac{75 \times 35}{100} + \frac{25 \times 37}{100}$$

Ar = 35.5

Assignment

Naturally occurring potassium consists of potassium-39 and potassium-41. Calculate the percentage of each isotope present if the average is 39.1.

USES OF RADIOISOTOPES

- ✓ **Radioisotopes** are unstable isotopes that undergo disintegration to produce new stable atoms with the emission of radiation in the form of rays and particles and large amount of energy. These radioisotopes have many applications such as the ones given below:

1. **Treatment of cancer and tumours:**

- ✓ Radioisotopes produce radiations that can destroy cells and because of this, controlled radiation from radioisotope sources is focused on tumours and cancer cells destroying them in the process. In this way tumours and cancer cells are treated or cured. This kind of treatment is called **radiotherapy**.

2. **Detection of leakages and flaws in materials:**

- ✓ Radioisotopes sources are added to oil or water in pipes and a detector of radiation called **Geiger counter** detects leaking radiation outside the pipes indicating leaks in pipes. Radioisotopes used in his way are called **tracers**.

3. **Control of thickness in the manufacture of shits:**

- ✓ A material whose to be made into a shit with desired thickness is placed on rollers in between a radioisotope source and radioactive counter. The counter detects any minor change in radiation caused by change in thickness of the material and rollers are adjusted to the desired thickness of the shit.

4. **Sterilisation: This is the process of purifying equipment.**

- ✓ Radioisotopes producing gamma rays are used to sterilise syringes and other disposable medical equipment because gamma rays kill germs and bacteria. So medical equipment that needs recycling are treated with a dose of gamma rays from cobalt-60 and cesium-137 to kill germs and bacteria.

5. **Preservation of food:**

- ✓ In many countries food stuffs are treated with a low dose of radiation from cobalt-60 and cesium-137 to kill germs and bacteria that cause decay.

6. **Radio – carbon dating:**

- ✓ This is the process of determining the age of fossils using radioisotope carbon – 14. This isotope (^{14}C) is taken by plants from the atmosphere and is present in animals that eat plants. The Carbon-14 within a living organism is continually decaying, but as the organism is continuously absorbing Carbon -14 throughout its life the ratio of Carbon-14 to Carbon-12 atoms in the organism is the same as the ratio in the atmosphere. Once an organism dies it stops taking in Carbon in any form and carbon – 14 begins to decay. By measuring the ratio of Carbon-14 in a sample and comparing it to the amount in a recently deceased sample its date can be determined. The present age of the earth, using this method, has been estimated to be about 4 billion years.

7. **Detection of tumours, internal bleeding and monitoring uptake of mineral salts in plants:**

✓ The method is also used in medicine to detect brain tumours and internal bleeding and in agriculture to study the uptake of fertilisers by plants. Radioactive materials are injected into a patient's body and radiation is detected outside the body to help doctors study the systems such as the circulatory system.

ELECTRONIC STRUCTURES OR ELECTRONIC CONFIGURATION

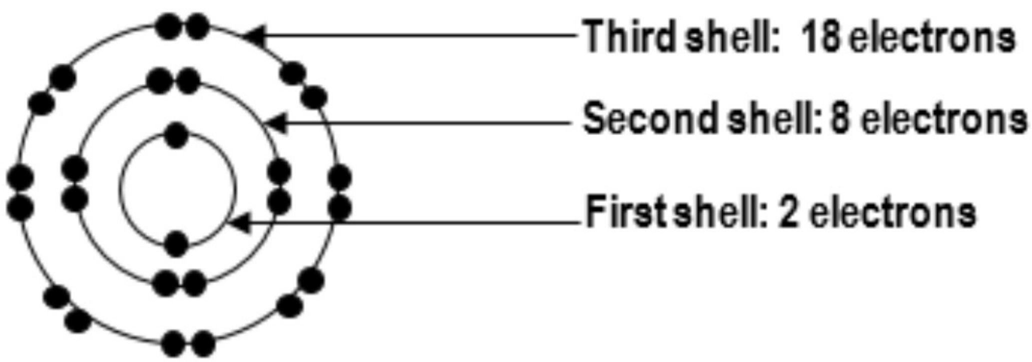
☐ **This is the way in which electrons are distributed or filled up in shells or energy levels of atoms.**

✓ Electrons can only fill up definite energy levels and that they cannot exist between them. Each of the energy levels can hold only a certain number of electrons.

✓ The formula which relates the number of electrons a shell can hold is $\mathbf{e = 2n^2}$ where e is number of electrons and n is number of shells.

- For first energy level, n =1, e = $2n^2 = 2(1)^2 = 2$, hence first shell holds up to **2** electrons.
- For second energy level, n = 2, e = $2(2)^2 = 8$, hence second shell holds up to **8** electrons.
- For third energy level, n = 3, e = $2(3)^2 = 18$, hence third shell holds up to **18** electrons.

✓



The electrons fill the energy levels starting from the energy level nearest to the nucleus, which has the lowest energy. When the first is full (with two electrons) the next electron goes into the second energy level. When second energy level is full with eight electrons, then the electrons begin to fill the third and fourth energy levels as shown above.

- ❖ It should be noted that a third shell has two possible numbers of electrons to hold, maximum of 18 if the number of electrons is large and minimum of 8 if the number of electrons left cannot go up to 18. A fourth shell has three possible numbers electrons to hold, maximum of 32 if the number of electrons is large, minimum of 18 if the number electrons left cannot up to 32 or minimum of 8 if the number of electrons left cannot go up to 18.
- ❖ If the shell does not have the minimum number of electrons needed, the atom is said to be **unstable**.
- ❖ We use a cross (x) or dot (●) to represent electrons in shells when drawing electronic structures.

DRAWING ELECTRONIC STRUCTURES OF THE FIRST 20 ATOMS OF ELEMENTS

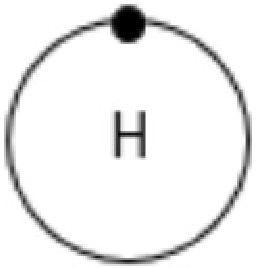
- ✓ **Electronic structures** are diagrams drawn to show arrangements of electrons in shells of atoms.
- ✓ Before drawing electronic structures of atoms of particular elements, you need to know the number of electrons contained by the atom. This information is obtained from the **nuclide notation** of the atom in the

periodic table.

✓ We insert the element's symbol in the centre of the electronic shell to represent the nucleus of the atom.

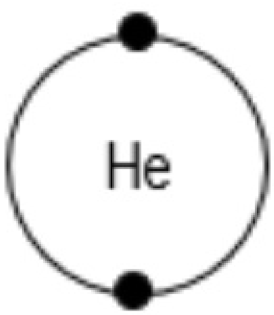
ELECTRONIC STRUCTURE OF HYDROGEN ATOM ($^{1}_{1}H$)

Since the atomic number of hydrogen is 1, hydrogen has one electron. Hence, hydrogen will only have one shell holding one electron as shown.



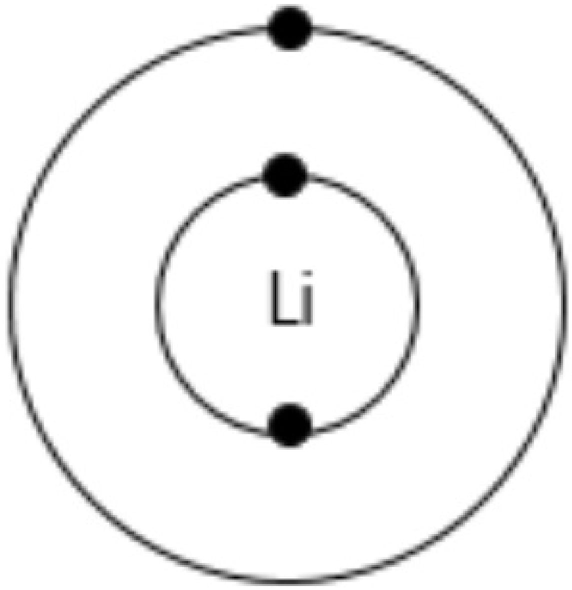
ELECTRONIC STRUCTURE OF HELIUM ATOM ($^{4}_{2}He$)

Helium atom has 2 electrons since it atomic number is 2. These two will go into the first shell, hence helium atom has only one electronic shell.



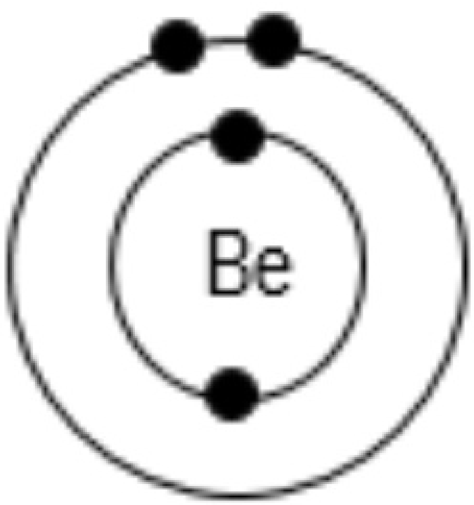
ELECTRONIC STRUCTURE OF LITHIUM ATOM ($^{7}_{3}Li$)

Lithium has 3 electrons, 2 electrons will go in the first shell and one electron in the other shell. (2: 1). So lithium atom has 2 electronic shells as shown below.



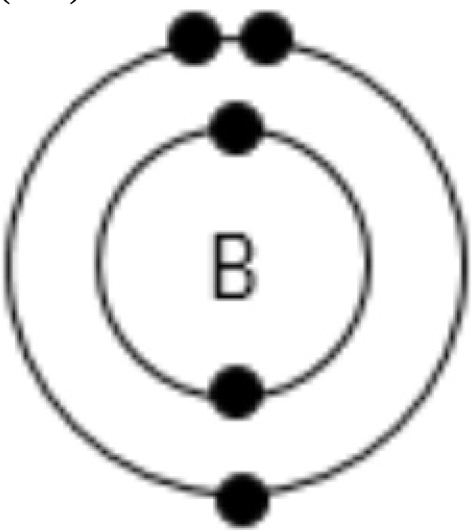
ELECTRONIC STRUCTURE OF BERYLLIUM ATOM ($^{9}_{4}Be$)

Beryllium has 4 electrons, 2 will fill up the first shell and the remaining 2 will fill up the other shell. (2: 2). So beryllium has 2 electronic shells as shown below.



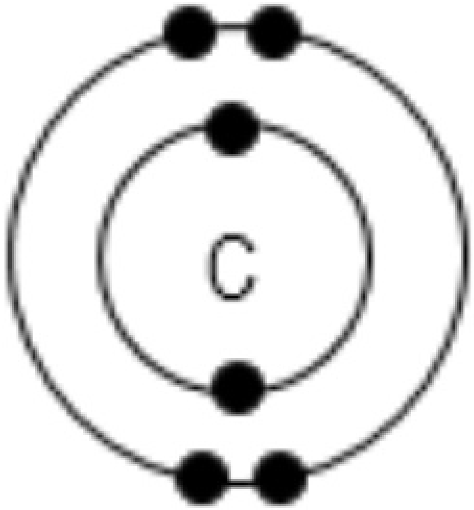
ELECTRONIC STRUCTURE FOR BORON ATOM ($^{11}_{5}B$)

Boron has 5 electrons, 2 will occupy the first shell and 3 the other shell. (2: 3). So boron has 2 electronic shells as shown below.

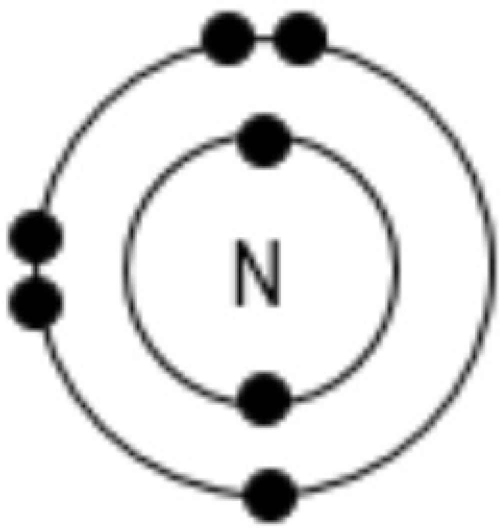


ELECTRONIC STRUCTURES FOR CARBON ATOM ($^{12}_{6}C$)

The nuclide notation of carbon atom shows that carbon has 6 electrons, 2 occupies the first shell and 4 occupy the second shell. (2: 4). So carbon has 2 electron shells as shown below.

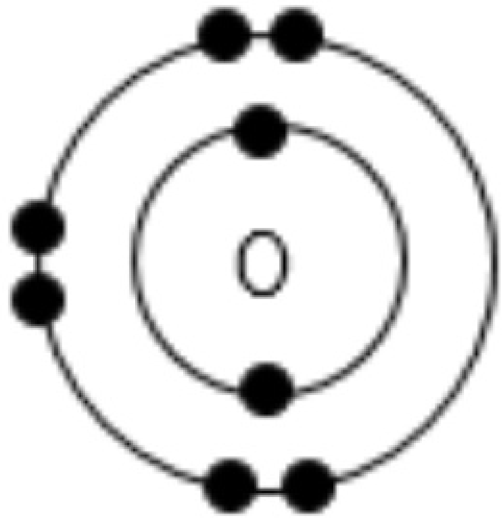


ELECTRONIC STRUCTURE OF NITROGEN ATOM ($^{14}_{7}N$)



Nitrogen has 7 electrons, 2 electrons occupy the first shell and 5 occupy the second shell. (2: 5). So nitrogen has 2 electronic shells as shown below.

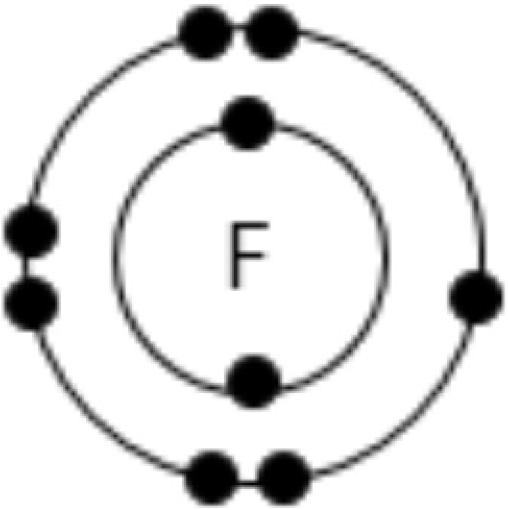
ELECTRONIC STRUCTURE FOR OXYGEN ATOM ($^{16}_{8}O$)



Oxygen has 8 electrons, 2 electrons occupy the first shell, 6 electrons occupy the second shell. (2: 6). So oxygen has 2 electronic shells as shown below.

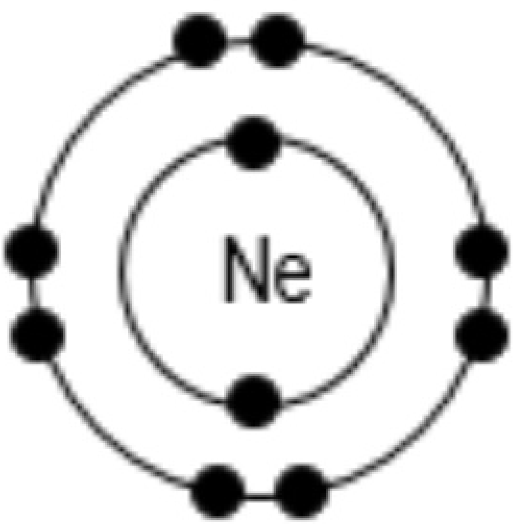
ELECTRONIC STRUCTURE FOR FLUORIDE ATOM ($^{19}_{9}F$)

Fluorine has 9 electrons; 2 electrons occupy the first shell while 7 electrons occupy the second shell. (2: 7). So fluorine has 2 electronic shells as shown below.



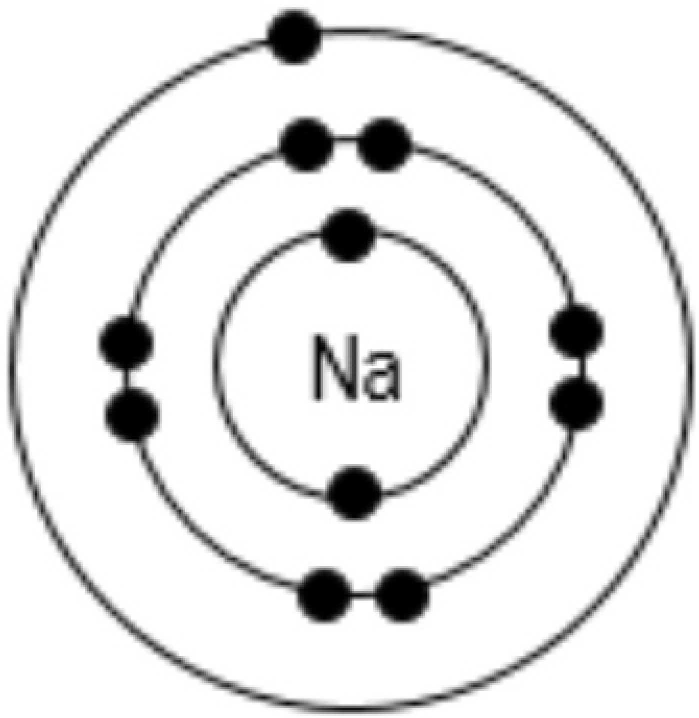
ELECTRONIC STRUCTURE FOR NEON ATOM ($^{20}_{10}Ne$)

Neon has 10 electrons, 2 electrons occupy the first shell and 8 electrons occupy the second shell. (2: 8). So neon has 2 electronic shells as shown below. In neon atom, the first shell and second shell hold the maximum number of electrons required. Hence an atom neon is said to be stable.



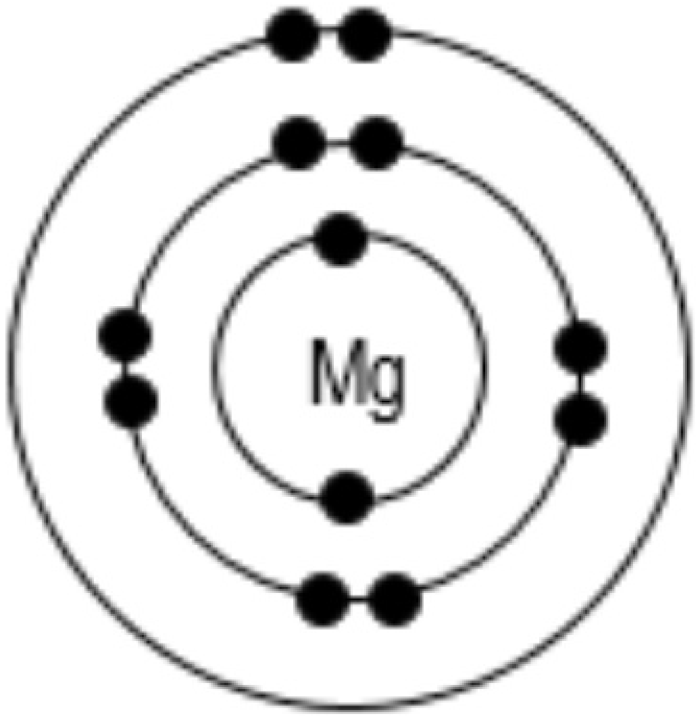
ELECTRONIC STRUCTURE FOR SODIUN ATOM ($^{23}_{11}Na$)

Sodium has 11 electrons, 2 electrons occupy the first shell, 8 electrons the second shell and 1 electron occupy the third shell. (2: 8: 1). So sodium has 3 electronic shells as shown below. Since the third shell of sodium atom does not hold 18 or 8 electrons, the atom of sodium is said to be unstable.



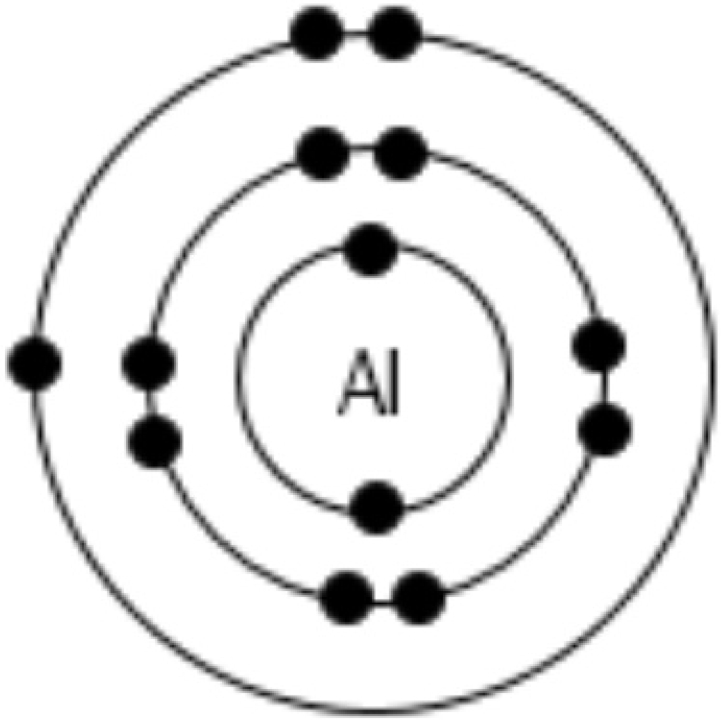
ELECTRONIC STRUCTURE FOR MAGNESIUM ATOM ($^{24}_{12}Mg$)

Magnesium has 12 electrons, 2 electrons occupy the first shell, 8 electrons occupy the second shell and 2 electrons occupy the third shell. (2: 8: 2). So magnesium has 3 electronic shells as shown below.



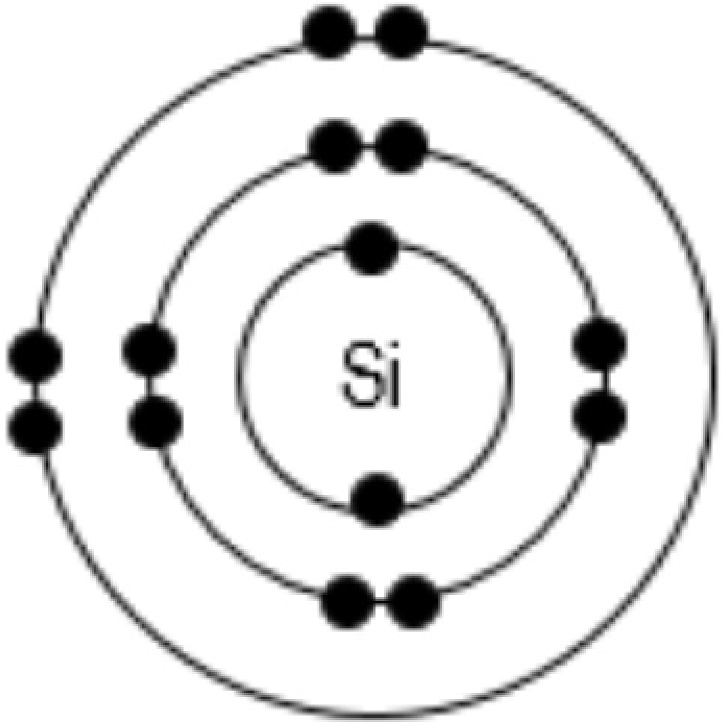
ELECTRONIC STRUCTURE FOR ALUMINIUM ATOM ($^{27}_{13}\mathbf{Al}$)

Aluminium atom has 13 electrons, 2 electrons occupy the first shell, 8 electrons hold the second shell and 3 electrons occupy the third shell. (2: 8: 3). So aluminium has 3 electronic shells as shown below.



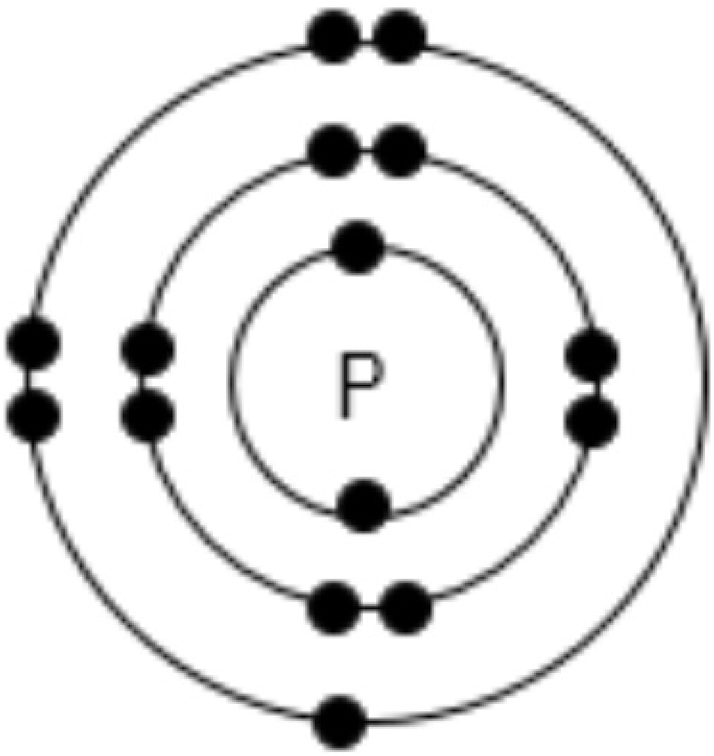
ELECTRONIC STRUCTURE FOR SILICON ATOM ($^{28}_{14}Si$)

Silicon atom has 14 electrons, 2 electrons occupy the first shell, 8 electrons occupy the second shell and 4 electrons occupy the third shell. (2: 8: 4). So silicon atom ha 3 electronic shells as shown below.



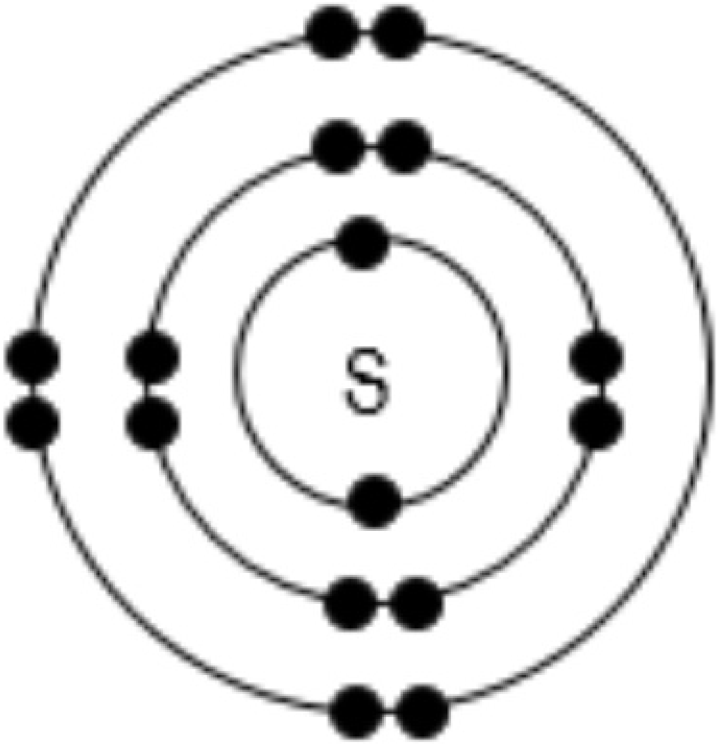
ELECTRONIC STRUCTURE FOR PHOSPHORUS ($^{31}_{15}P$)

Phosphorus atom has 15 electrons, 2 electrons in the first shell, 8 electrons in the second shell and 5 electrons in the third shell. (2: 8: 5). So phosphorus has 3 electronic shells.



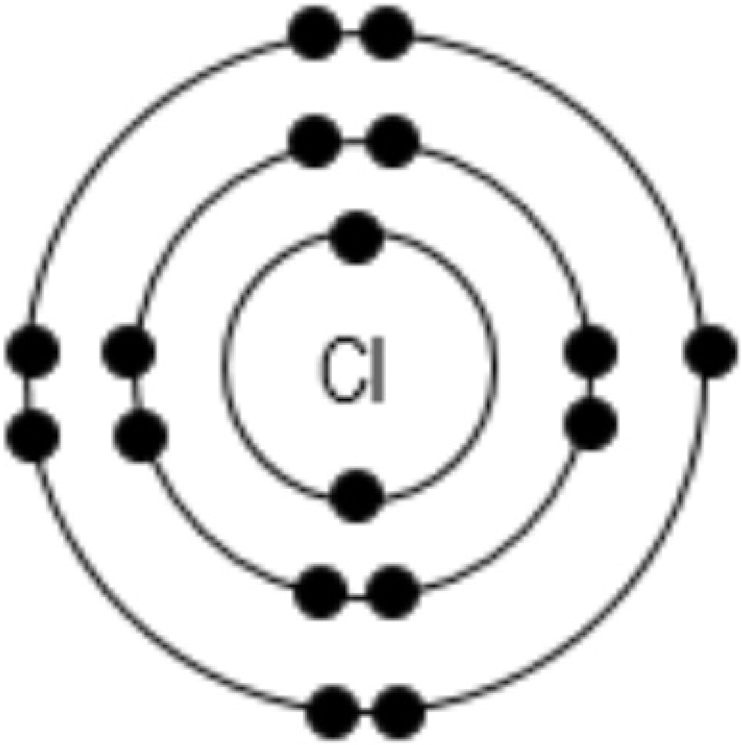
ELECTRONIC STRUCTURE FOR SULPHUR ATOM ($^{32}_{16}S$)

Sulphur atom has 16 electrons, 2 electrons hold the first shell, 8 electrons hold the second shell and 6 electrons hold the third shell. (2: 8: 6). So sulphur atom has 3 electronic shells as shown below. (2: 8: 6)



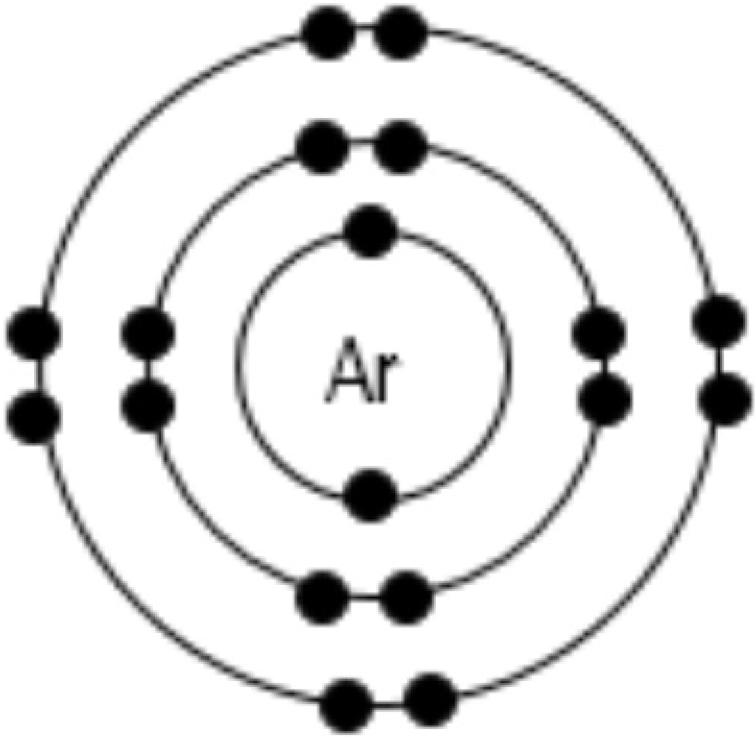
ELECTRONIC STRUCTURE FOR CHLORINE ATOM ($^{35.5}_{17}Cl$)

Chlorine atom has 17 electrons, 2 electrons hold the first shell, 8 electrons hold the second shell and 7 electrons hold the third electrons. (2: 8: 7). So chlorine atom has 3 electronic shells as shown below.



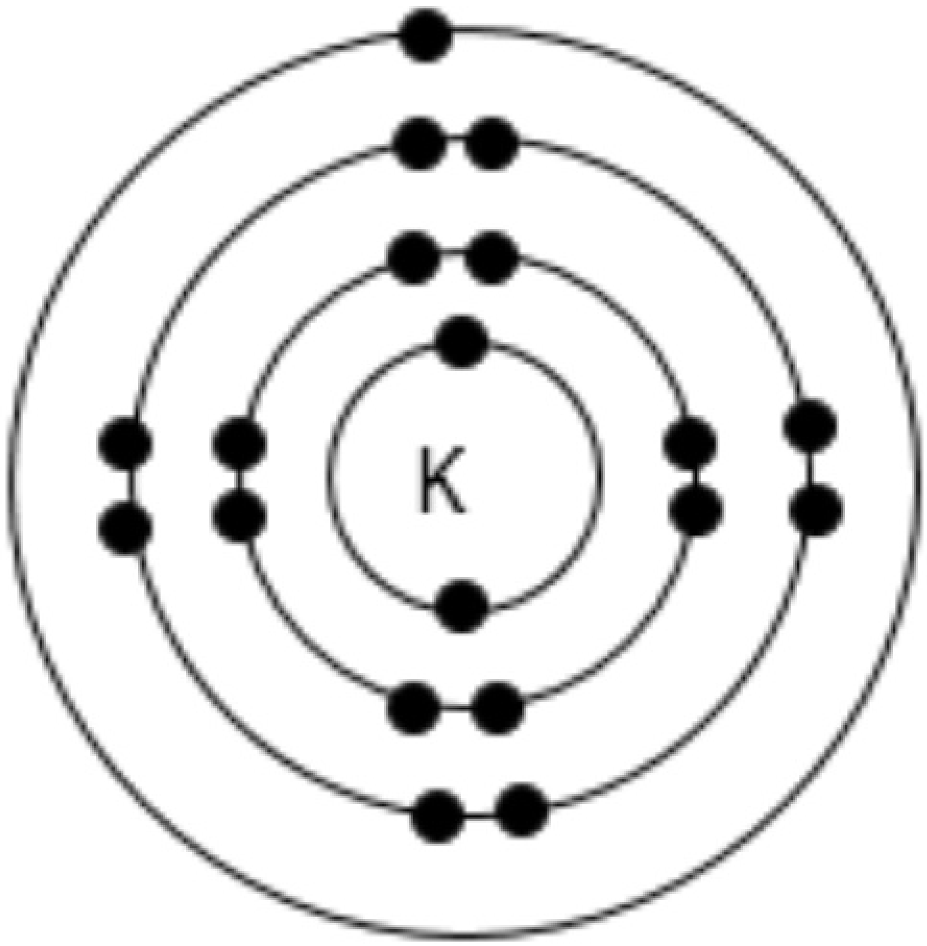
ELECTRONIC STRUCTURE FOR ARGON ($^{40}_{18}Ar$)

Argon atom has 18 electrons, 2 electrons hold the first shell, 8 electrons hold the second shell and 8 hold the third shell. Since the third shell of argon atoms has a minimum number of electrons in the third shell, argon is said to be stable. (2: 8: 8). It has 3 electronic shell as shown below.

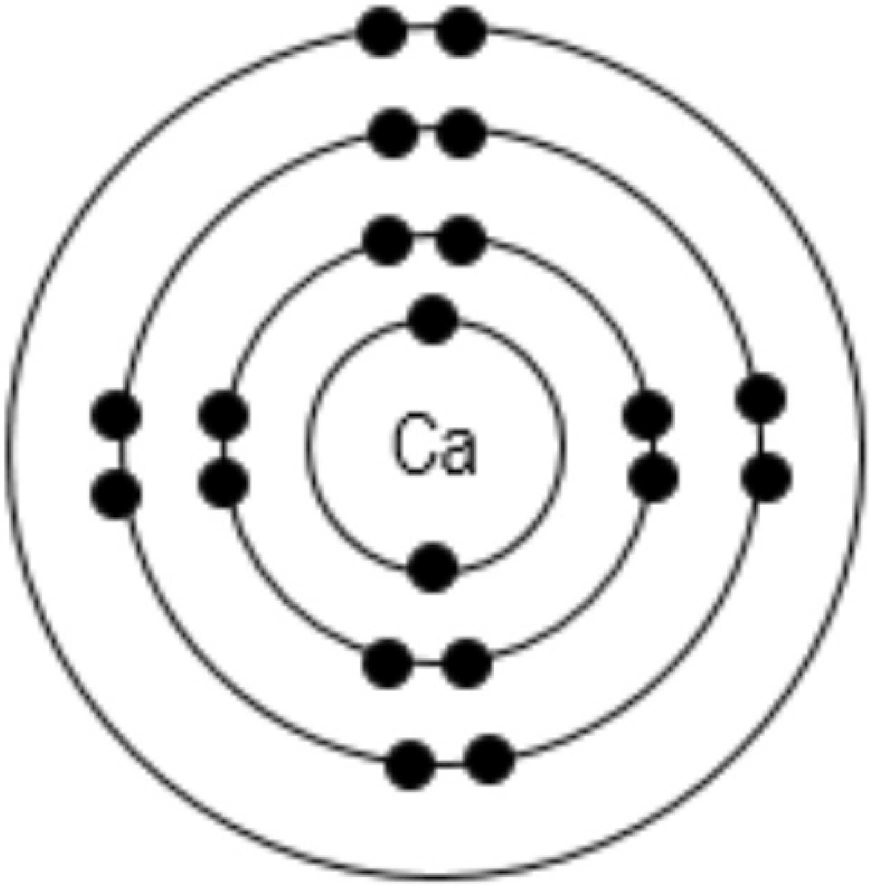


ELECTRONIC STRUCTURE FOR POTASSIUM ATOM ($^{39}_{19}K$)

Potassium atom has 19 electrons, 2 electrons hold the first shell, 8 electrons hold the second shell, 8 electrons hold the third shell and 1 electron hold the fourth electron. So potassium has 4 electronic shells as shown below. (2: 8: 8:1).



ELECTRONIC STRUCTURE FOR CALCIUM ATOM ($^{40}_{20}Ca$)



Calcium atom has 20 electrons, 2 electrons hold the first shell, 8 electrons hold the second shell, 8 electrons hold the third shell and 2 electrons hold the fourth shell. (2: 8: 8: 2). So calcium atom has 4 electronic structures.

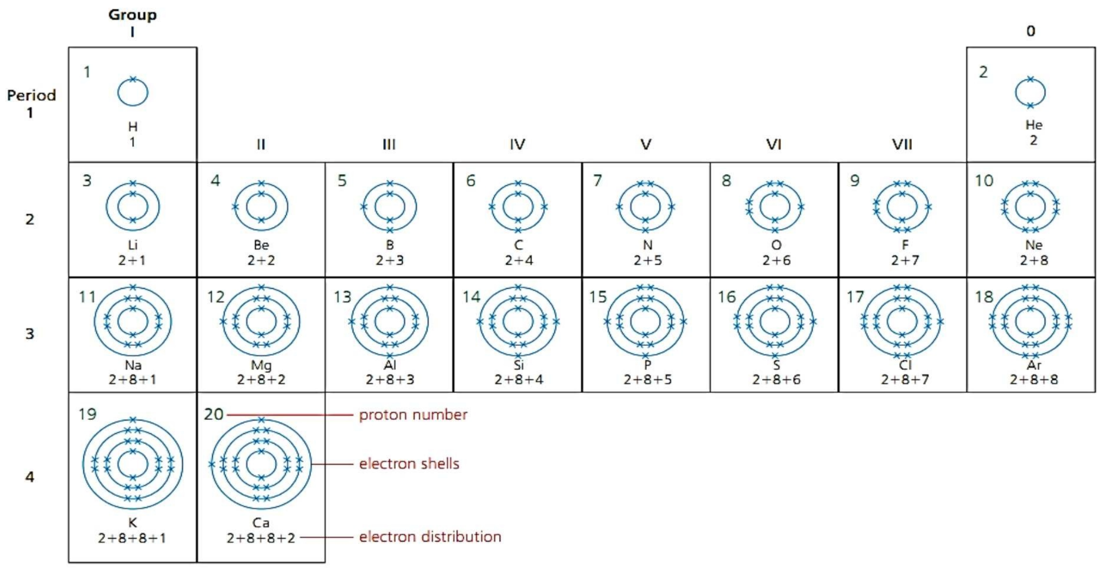
ELECTRONIC STRUCTURE AND THE PERIODIC TABLE

- ✓ From electronic structures of atoms drawn above, it can be noted that the **number of electrons found in the outermost shell** determines the group in which the atom belongs whereas the **number of shells contained in the atom determines the period in which the atom belongs**. Therefore, we can conclude that electronic structures provide the bases for determining **groups** and **periods** in which atoms of elements belong in the periodic table.
- ✓ The number of electrons found in the outermost shell of atoms is called **valence electron**. Valence electrons are equal to group number.
- ✓ Atoms with the same number of valence electrons belong to the same

group while atoms with same number of shells belong to the same period.

- ✓ The valence electrons dictate how an element reacts. Atoms with the same number of valence electrons react in the similar manner.

- ✓



As the number of electrons increases, the group number as well as the period increases.

CHEMICAL BONDING

- **Bonding is the chemical combination of two or more atoms of elements to form a compound.**
- Atoms of elements bond in order to acquire stable filled outermost shell.
- Atoms bond by either losing, gaining or sharing outermost electrons (valence electrons).
- During bonding, nearly empty shells lose electrons while nearly filled shells gain or share electrons.

Types of bonding

There three types of bonding namely;

- Ionic (electrovalent) bonding.
- Covalent bonding.
- Metallic bonding.

Ionic bonding

- **Ionic bonding is an electrostatic force of attraction between positively charged atoms from metals and negatively charged atoms from non-metals.**
- Ionic bonding involves the transfer of electrons from metallic atoms to non-metallic atoms to form a compound.
- Metallic atoms lose electrons to form stable positively charged ions called **cations** while non-metallic atoms gain electrons to form stable negatively charged atoms called **anions**.
- The compounds formed by ionic bonding are called **ionic compounds**. They are made of positively and negatively charged **ions**. **Ions** are positively charged atoms.

STRUCTURES OF IONIC BONDING

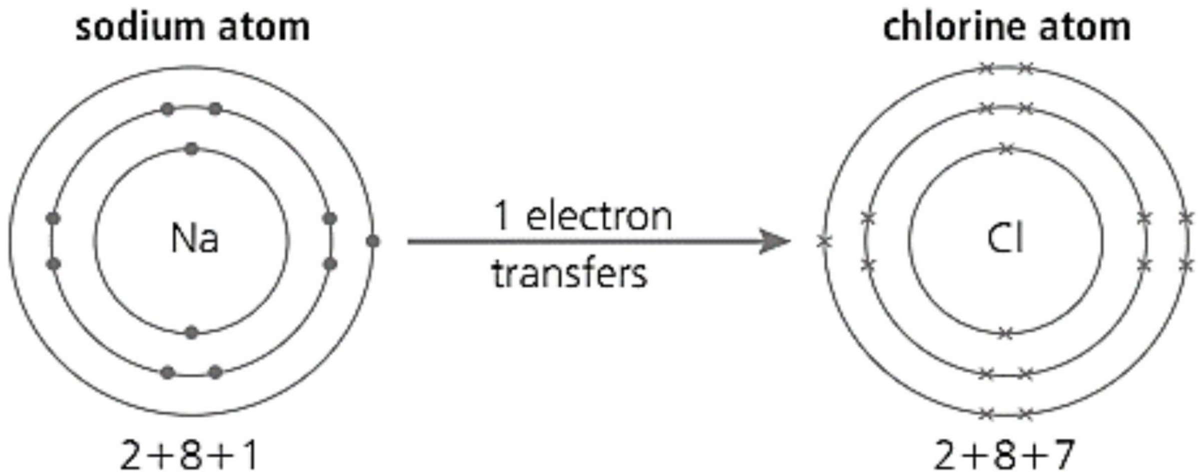
SODIUM CHLORIDE:

- Sodium chloride is a compound formed by ionic bonding between sodium metal atom and chlorine non-metal atom.

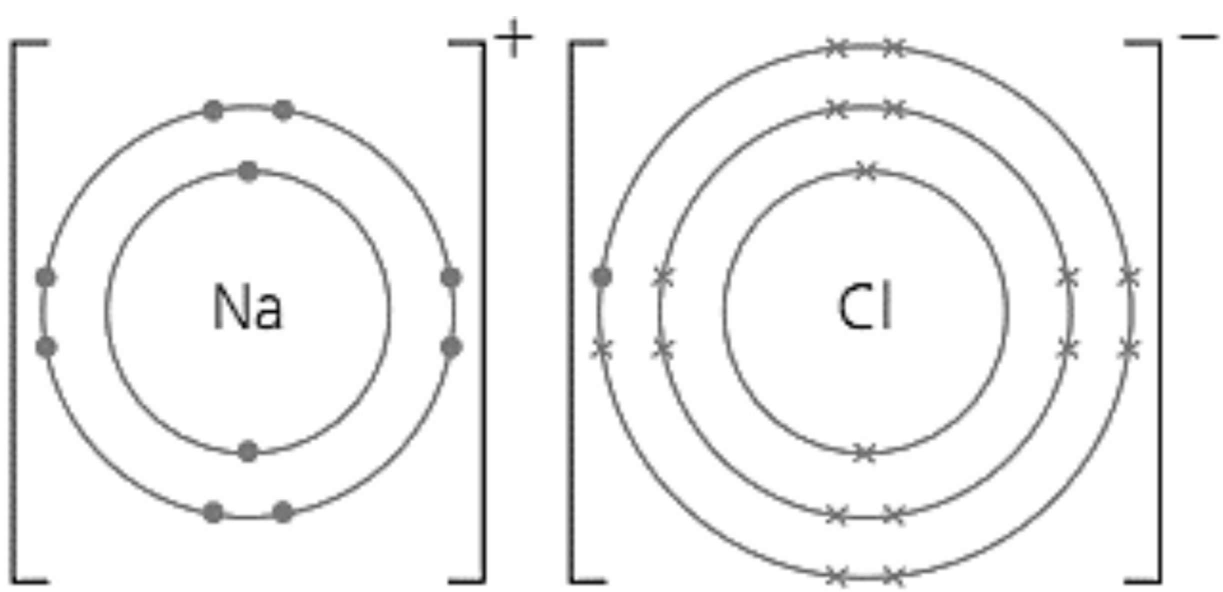
Chlorine has 17 electrons arranged in shells as shown below. Since the last shell is nearly filled, one electron is gained from sodium atom to form negatively chloride ion (Cl^-).

Sodium has 11 electrons arranged in shells as shown below. Since the last shell is nearly empty, the electron in the last shell is lost to chlorine atom and form a stable positively charged sodium ion (Na^+).

Structure of sodium atom ($^{23}_{11}Na$): Structure of chlorine atom ($^{35.5}_{17}Cl$):



- The two ions have opposite charges, so they attract each other. The force of attraction between them is strong. It is called an **ionic bond**. Below is the electronic structure of sodium chloride.



- Since it is made of ions, sodium chloride is called an ionic compound. It contains one Na^+ ion for each Cl^- ion, so its formula is **NaCl**.

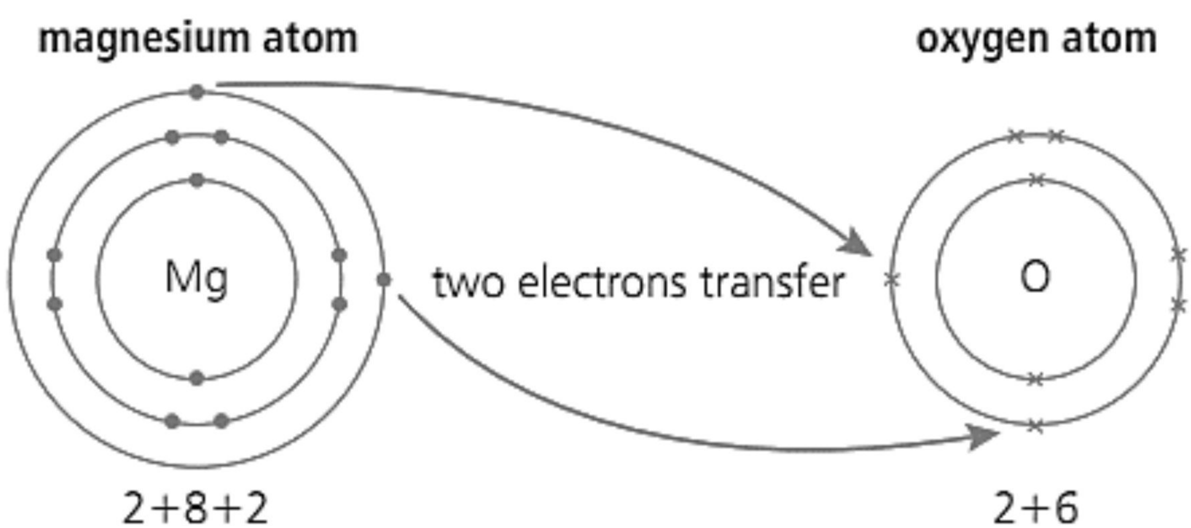
MAGNESIUM OXIDE:

- Magnesium oxide is a compound formed by ionic bonding between magnesium metal atom and oxygen non-metal atom.

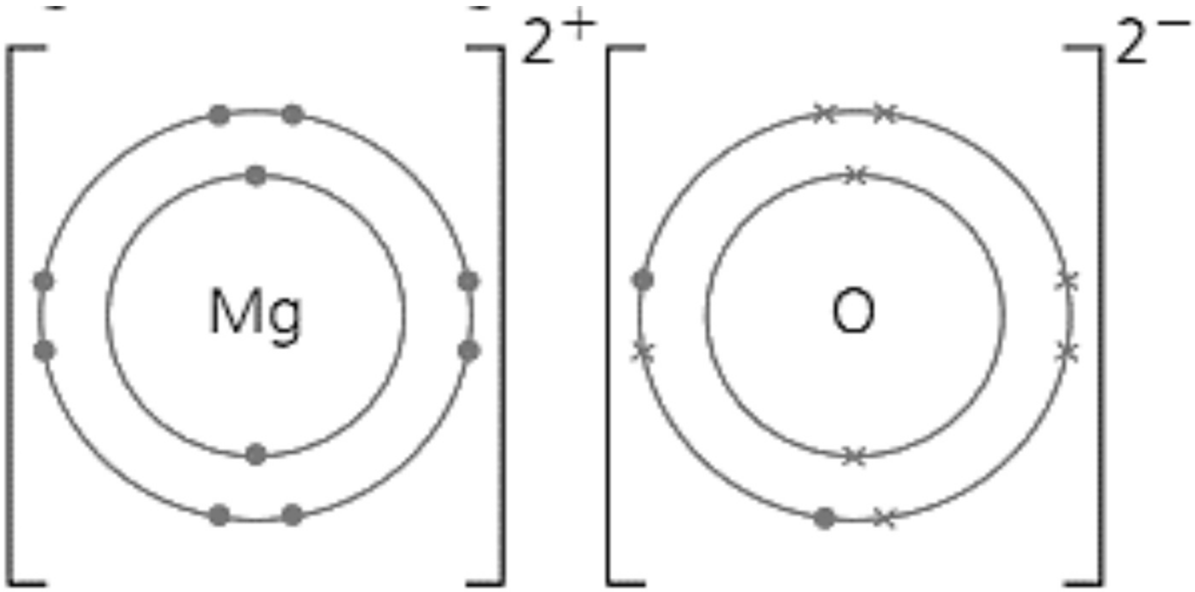
Structure of magnesium atom **(**$^{24}_{12}\mathbf{Mg}$**)**: Structure of oxygen atom ($^{16}_{8}\mathbf{O}$ **):**

Oxygen atom has 8 electrons arranged in shells as shown below. Since the last shell is nearly filled, two electrons are gained from magnesium atom to form negatively oxygen ion (O^{2-}).

Magnesium atom has 12 electrons arranged in shells as shown below. Since the last shell is nearly empty, the two electrons in the last shell are lost to chlorine atom and form a stable positively charged magnesium ion (Mg^{2+}).



- The diagram below is the electronic structure of sodium chloride.



- The resulting compound is called **magnesium oxide.** It has one magnesium ion for each oxide ion, so its formula is MgO.

MAGNESIUM CHLORIDE:

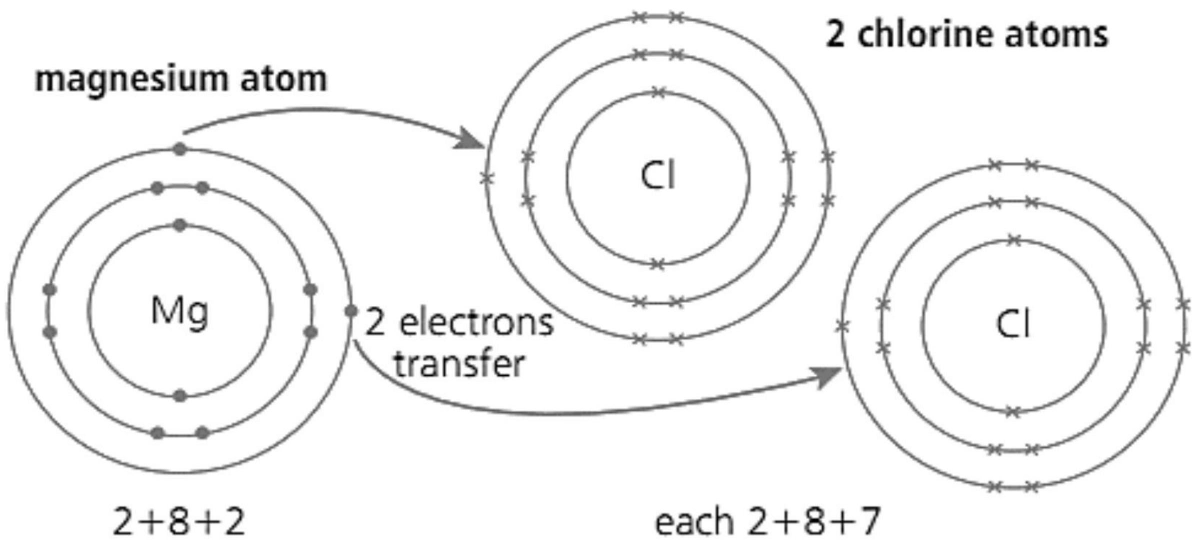
- Magnesium chloride is a compound formed by ionic bonding between magnesium metal atom and chlorine non-metal atom.

Structure of magnesium atom **(**$^{24}_{12}\mathbf{Mg}$**)**: Structure of chlorine atom **(** $^{35.5}_{17}\mathbf{Cl}$ **)**:

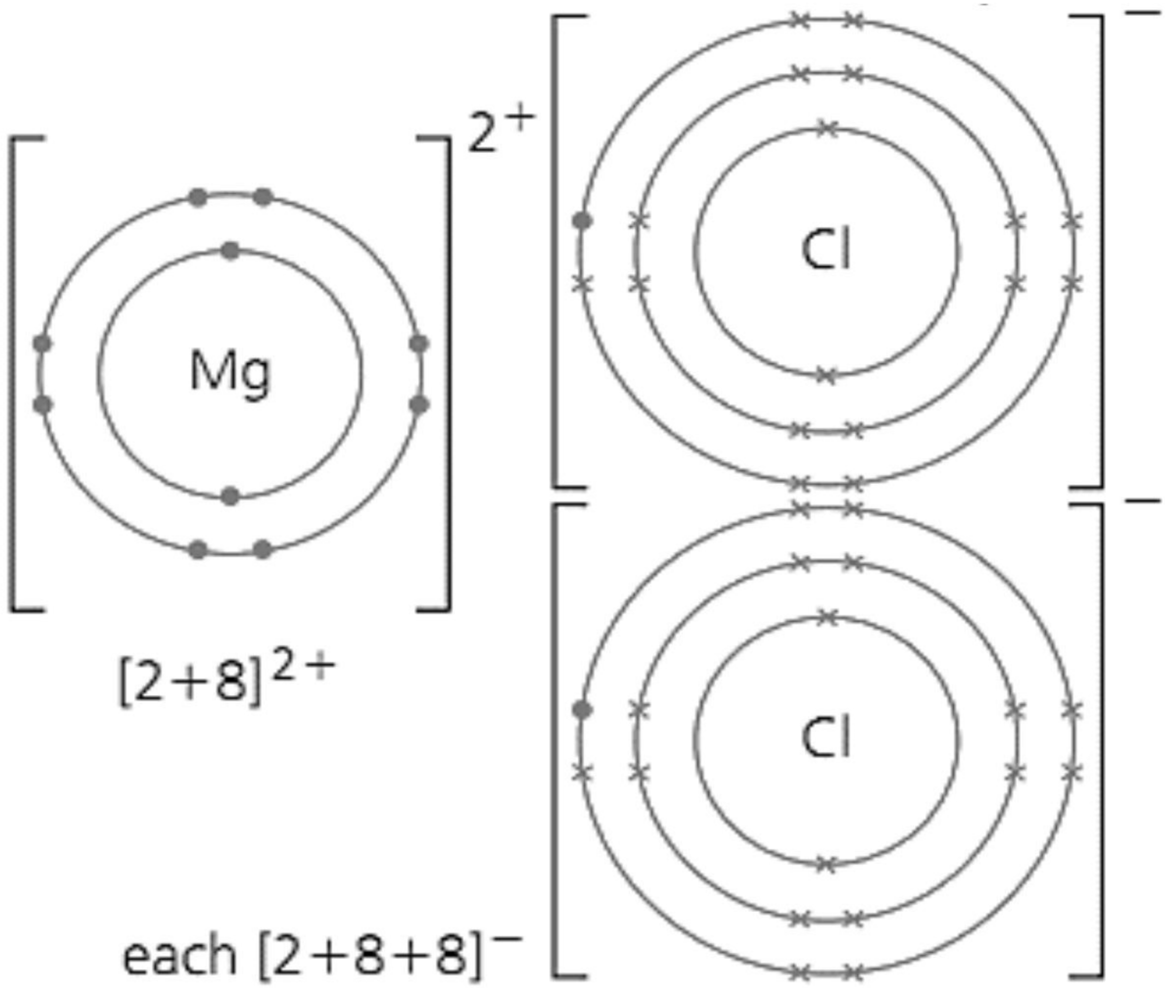
Chlorine has 17 electrons arranged in shells as shown below. Since the last shell is nearly filled, one electron is gained from sodium atom to form negatively chloride ion (Cl^-).

Magnesium atom has 12 electrons arranged in shells as shown below. Since the last shell is nearly empty, the two electrons in the last shell are lost to chlorine atom and form a stable positively charged magnesium ion (Mg^{2+}).

- Since magnesium atom loses two electrons and a chlorine atom only needs to gain one electron, then two chlorine atoms are needed to gain the two electrons as shown below.



- The diagram below is the structure of magnesium chloride.

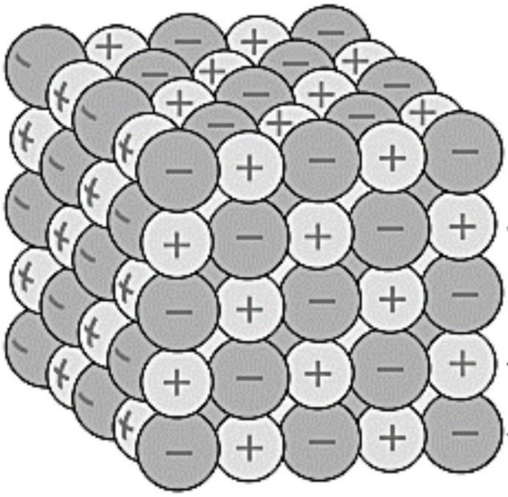


Mg
2+
$[2+8]^{2+}$
Cl
Cl
each $[2+8+8]^-$

- The resulting compound is called **magnesium chloride.** It has two chloride ions for each magnesium ion, so its formula is $MgCl_2$.

❖ One important reminder about ionic compounds in solid state is that they form a regular lattice made of billions of alternating cations and anions bonded together.

❖ The diagram below shows the lattice of sodium chloride compound.



COVALENT BONDING

- **Covalent bonding is an electrostatic force of attraction between the positively charged protons in the nucleus and the negative shared pair of electrons between two or more non-metal atoms.**
- Covalent bonding involves the sharing of one or more pairs of electrons in the outermost shell between non-metal atoms to form a compound. By sharing electrons, non-metal atoms become stable.
- Compounds formed by covalent bonding are **covalent compounds**. They are also called **molecules**. A molecule is a group two or more atoms of the same element covalently bonded together.
- So all molecules are covalent compounds and not all covalent compounds are molecules.
- They do not have charges.

Types of covalent bonds

(i) **Single bonds**: The type of bond formed when one pair of electrons is shared between atoms. It is represented by dash line (—). A pair of electrons has two electrons.

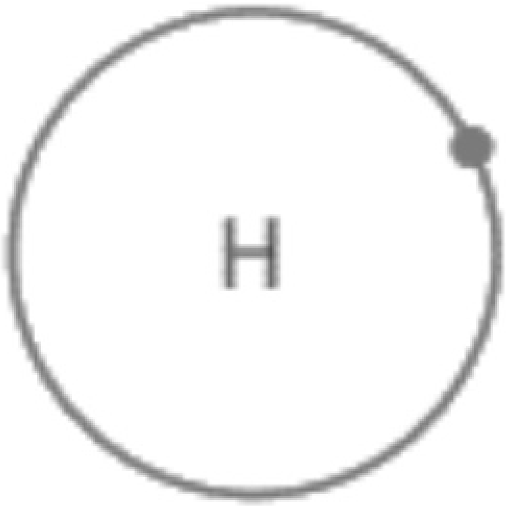
(ii) **Double bond**: The type bond formed when two pairs of electrons are shared between atoms. It is represented by double dash lines (=).

(iii) **Triple bond**: the type of bond formed when three pairs of electrons are shared between atoms. It is represented by triple dash lines (≡).

(iv) **Tetra bond**: the type of bond formed when four pairs of electrons are shared between atoms. It is represented by four dash lines.

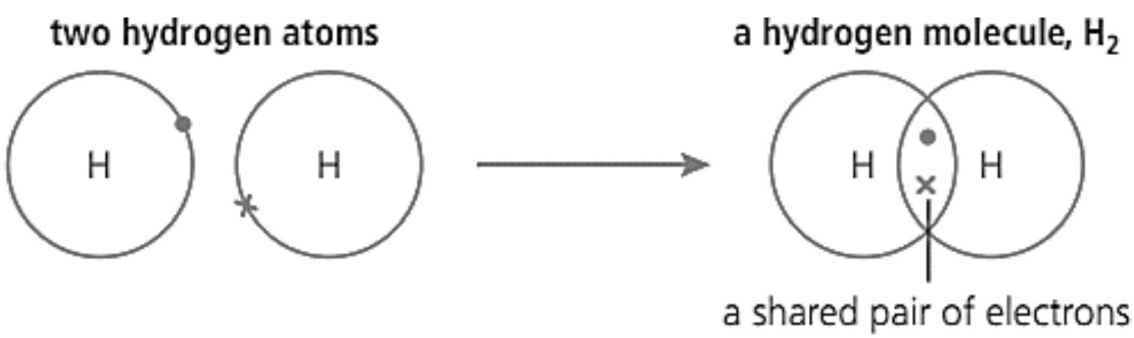
STRUCTURES OF COVALENT COMPOUNDS

HYDROGEN MOLECULE



- Below is the structure of hydrogen atom:

- There is one electron in the shell of hydrogen atom, so it needs one more electron from another hydrogen atom to be stable. So two hydrogen atoms come together and share one pair of electron by overlapping their shells as shown below.



- Each hydrogen atom has now gained a full shell of two electrons.
- Hydrogen molecule is formed from two hydrogen atoms covalently bonded together and its formula is written as H_2. Since only one pair of electrons is shared, the bond between the atoms is a **single bond** shown in a short way like this: H ─ H.

CHLORINE MOLECULE

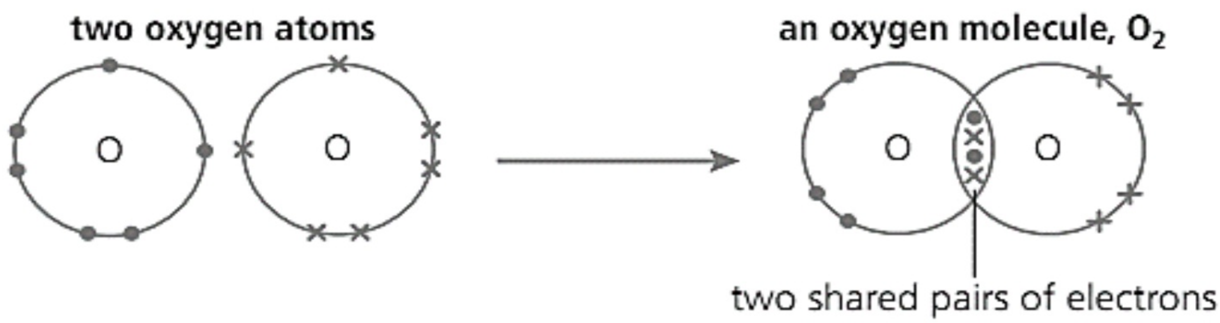
- Below is the structure of chlorine atom showing only the outer most shell. The shell needs one electron, hence this chlorine atom will share this electron with another chlorine atom by overlapping to form a shared pair as shown below.

- Each chlorine atom has now gained a full shell of eighty electrons.
- Chlorine molecule is formed from two chlorine atoms covalently bonded together and its formulas is written as **Cl_2**. Since only one pair of electrons is shared, the bond between the atoms is a **single bond** shown in a short way like this: Cl ─ Cl.

Oxygen molecule

- An oxygen atom has six outer electrons, so needs a share in two more. So two oxygen atoms share two electrons each, giving molecules with the

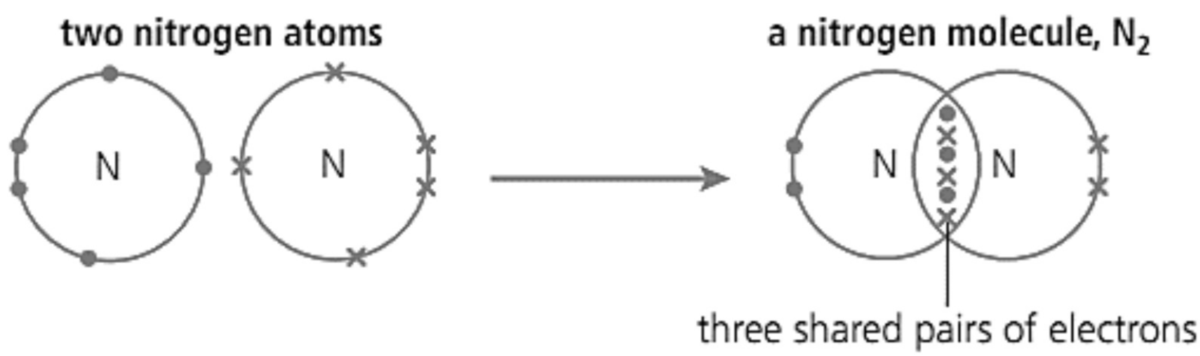
formula O_2. The diagram below shows the structure of oxygen molecule.



- Since the oxygen atoms share two pairs of electrons, the bond between them is a **double bond** shown it like this: O=O.

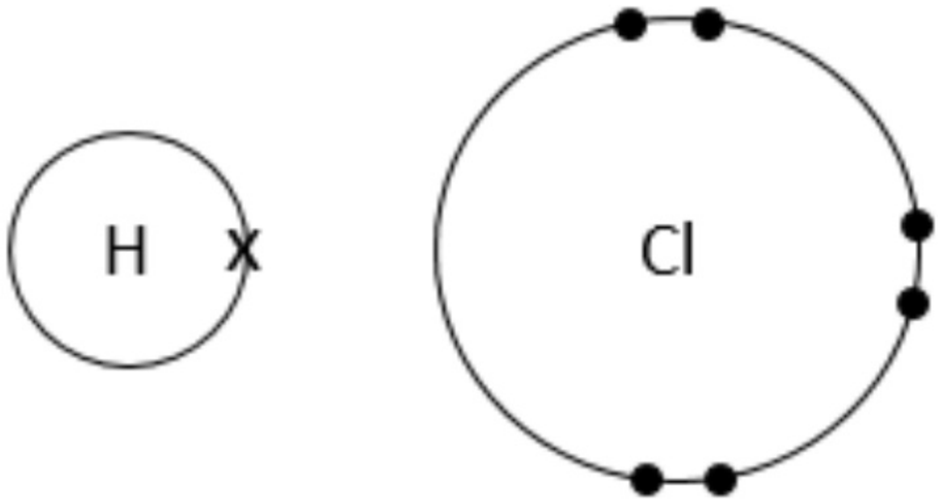
NITROGEN MOLECULE

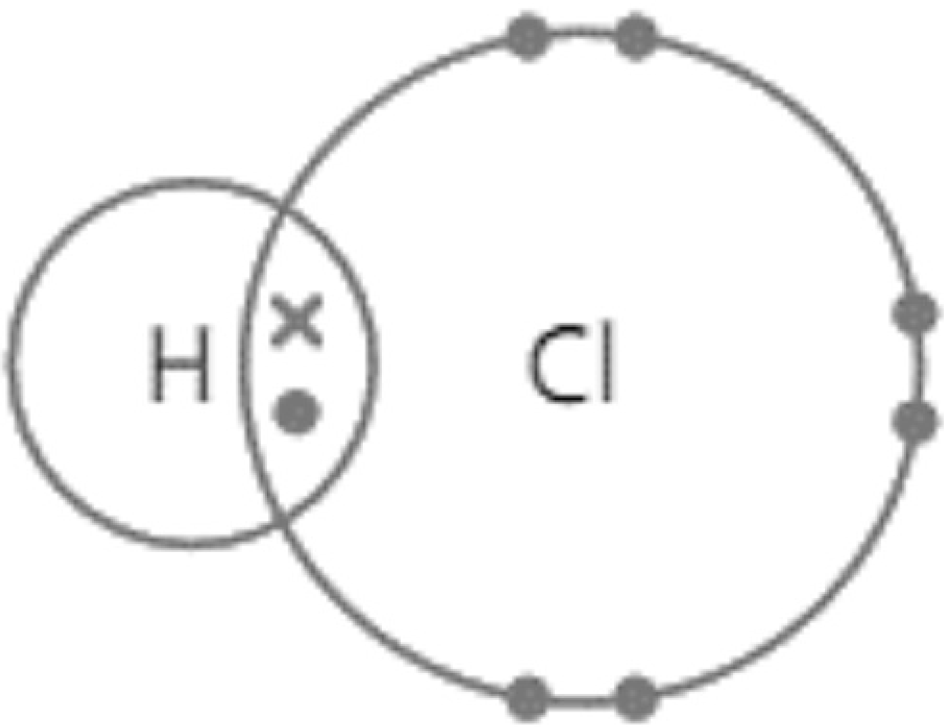
- A nitrogen atom has five outer electrons, so needs three more electrons. So two nitrogen atoms share three electrons each, giving molecules with the formula N_2. Each atom now has a stable outer shell of eight electrons:



- Since nitrogen atoms share three pairs of electrons, the bond between them is a triple bond shown in short like this: N≡N.

HYDROGEN CHLORIDE (HCl)



- 

Both chlorine and hydrogen atom need one electron in their outermost shell to become stable. They each share one electron to form hydrogen chlorine compound as shown below.

Hydrogen chloride molecule, HCl

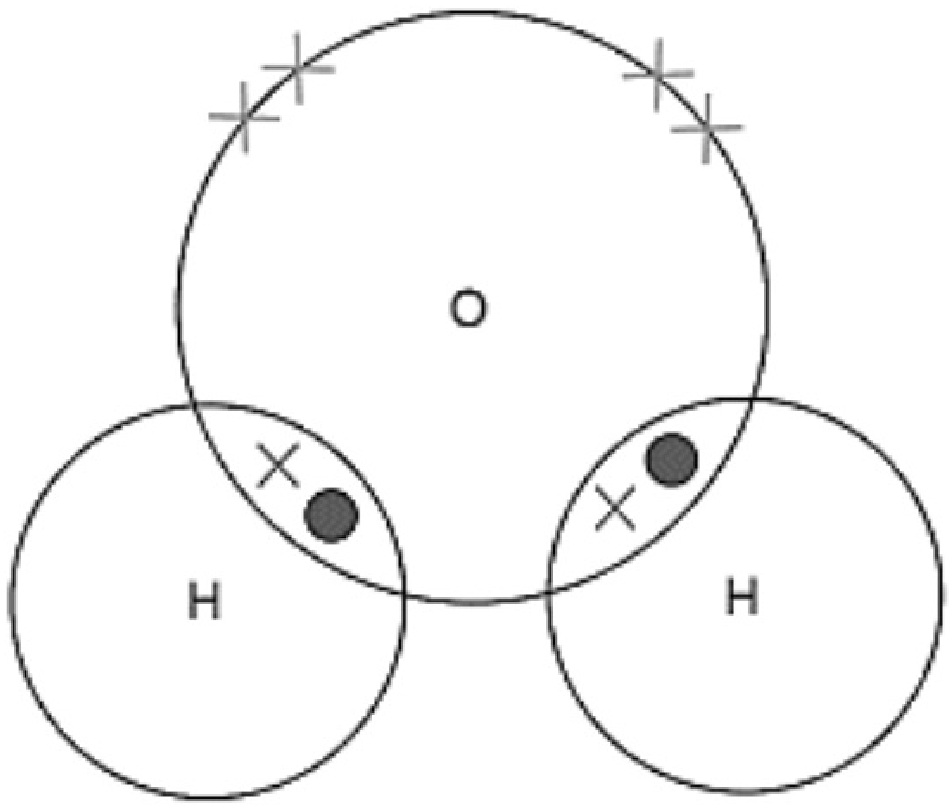
Chlorine atom

Hydrogen atom

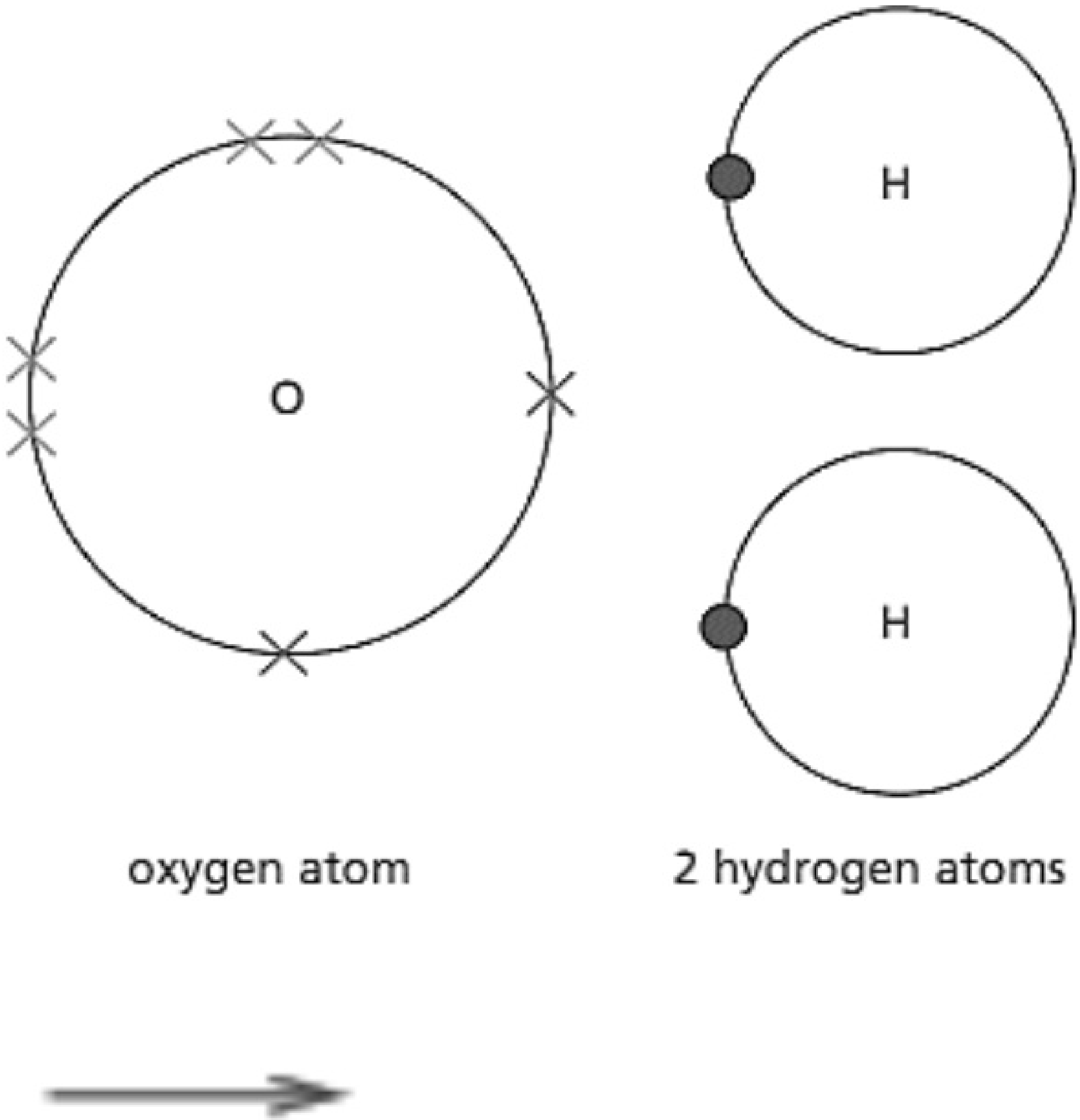
- Since there is one atom of hydrogen for one atom of chlorine, the formula for hydrogen chloride is HCl

WATER (H_2O)

- Oxygen atom needs two electrons while hydrogen atom needs one atom to become stable. So one oxygen atom shares its two electrons with two hydrogen atoms to form a compound as shown below.



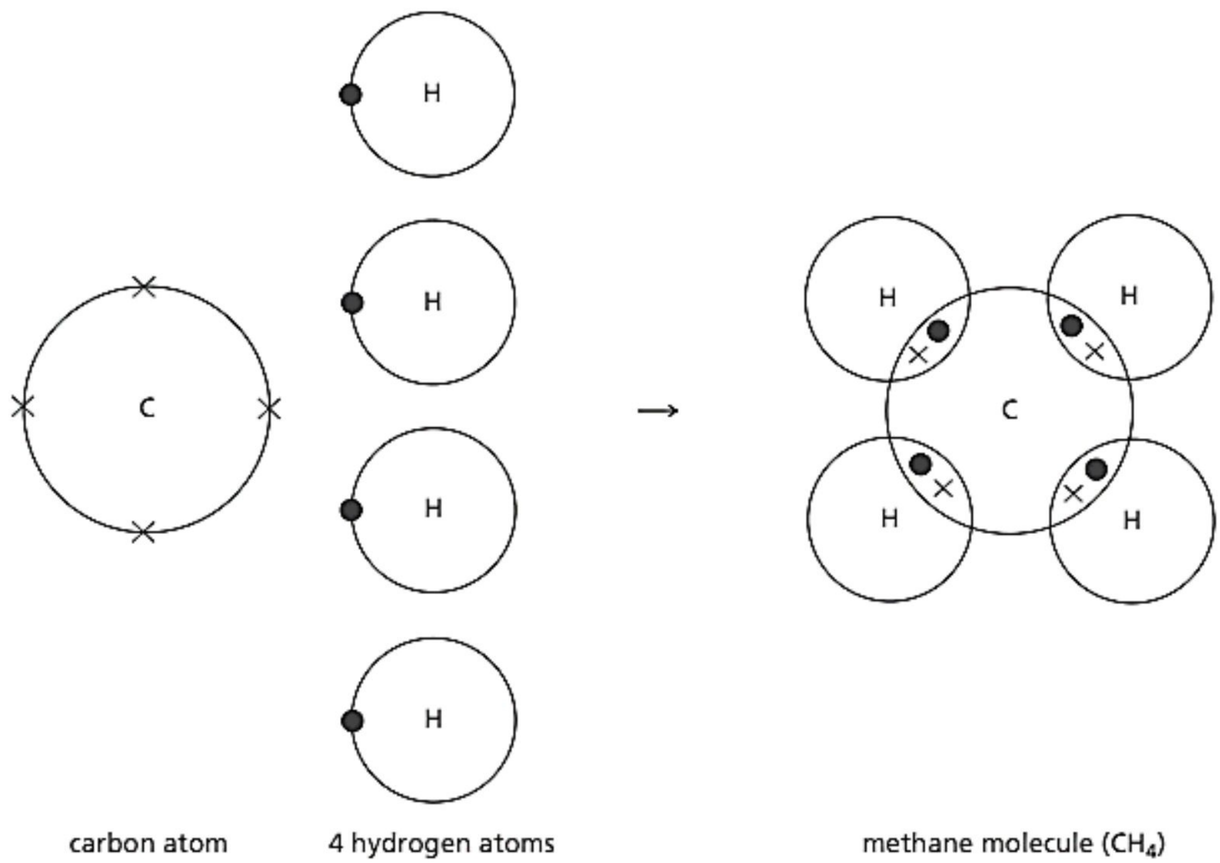
water molecule (H_2O)



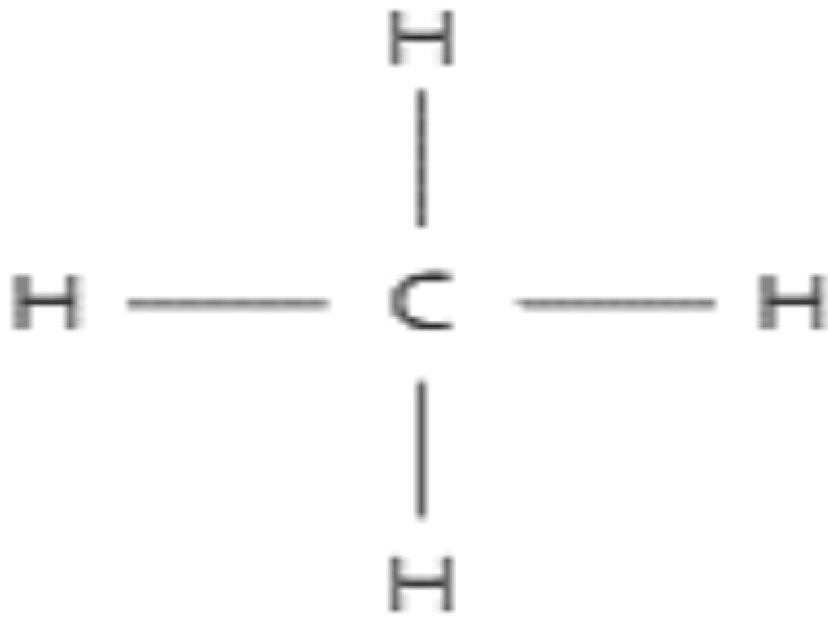
- There are two hydrogen atoms for each oxygen atom, hence its formula is H_2O.

Methane (CH_4)

- Carbon atom (C) needs four electrons while hydrogen atom (H) needs only one electron to become stable. So one carbon atom shares its four electrons with four hydrogen atoms to form a compound of methane as shown below.

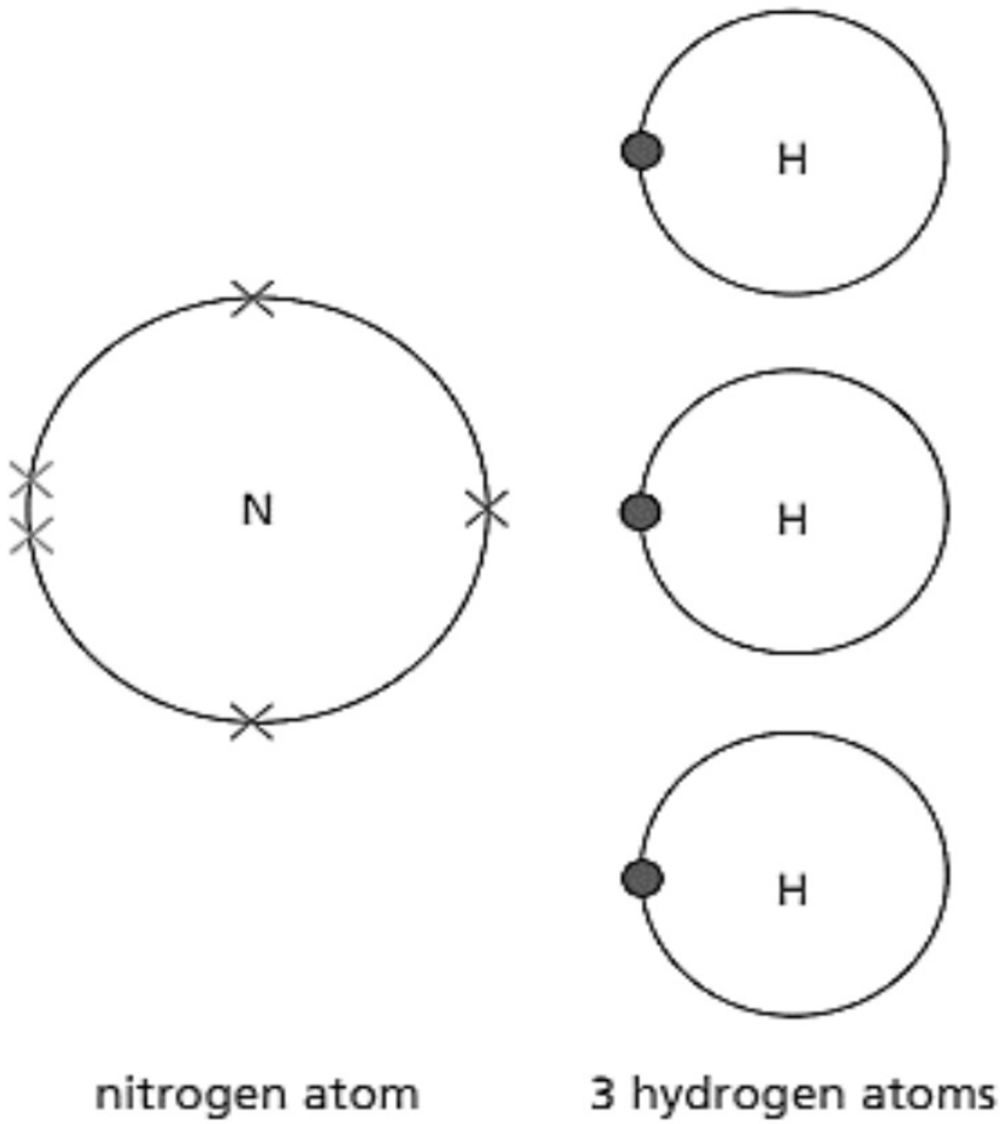


C
H
H
H
H
→
H
H
C
H
H
carbon atom
4 hydrogen atoms
methane molecule (CH_4)

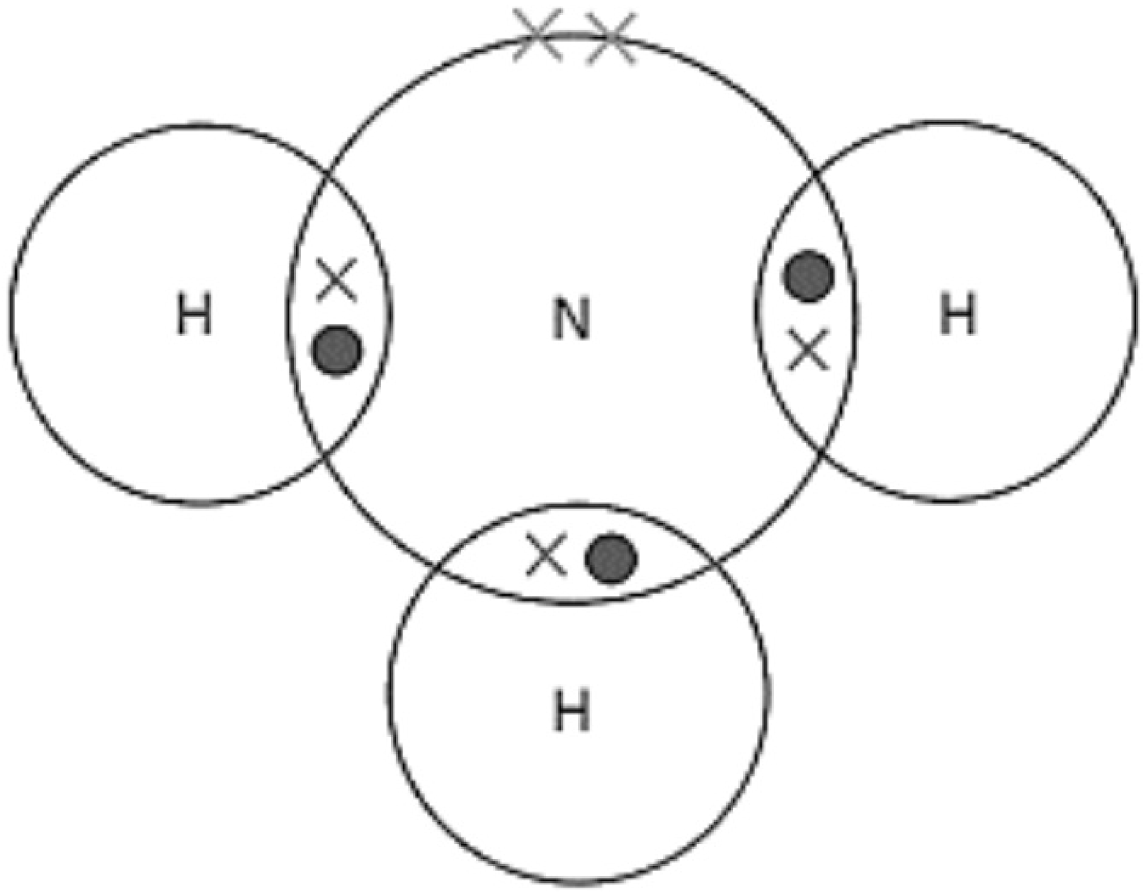


- Since one carbon atom combines with four hydrogen atoms, the formula for methane is CH_4. The single bonds in the compound of methane are written as shown below:

AMMONIA (NH_3)



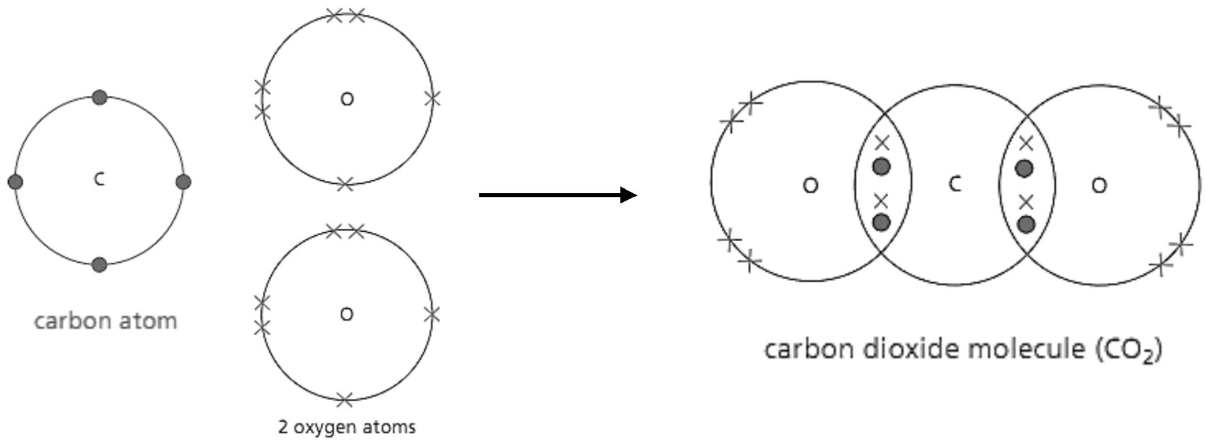
- Nitrogen atom (C) needs three electrons while hydrogen atom (H) needs only one electron to become stable. So one nitrogen atom shares its three electrons with three hydrogen atoms to form a compound of methane as shown below:



ammonia molecule (NH_3)

CARBON DIOXIDE (CO_2)

- Carbon atom needs 4 electrons in its outermost shell to become stable while oxygen atom needs two electrons in its outermost shell to become stable. So carbon atom shares its four electrons with two oxygen atom to form a compound of carbon dioxide as shown below.



- Since each oxygen atom shares two electrons with carbon atom, the bond is single represented as O=C=O. The formula for carbon dioxide is written as CO_2.

COVALENT BONDING IN MACROMOLECULES

- **Macromolecules are giant molecules formed when thousands of atoms join together by strong covalent bonds.**

❖ Carbon, silicon and oxygen are some of the atoms that bond in large numbers to form macromolecules.

- Examples of such macromolecules includes diamond, graphite, silicon(iv) oxide and plastics such as polythene.

(i) Diamond and graphite

❖ These are giant macromolecules formed when carbon atoms covalently bond together in large numbers.

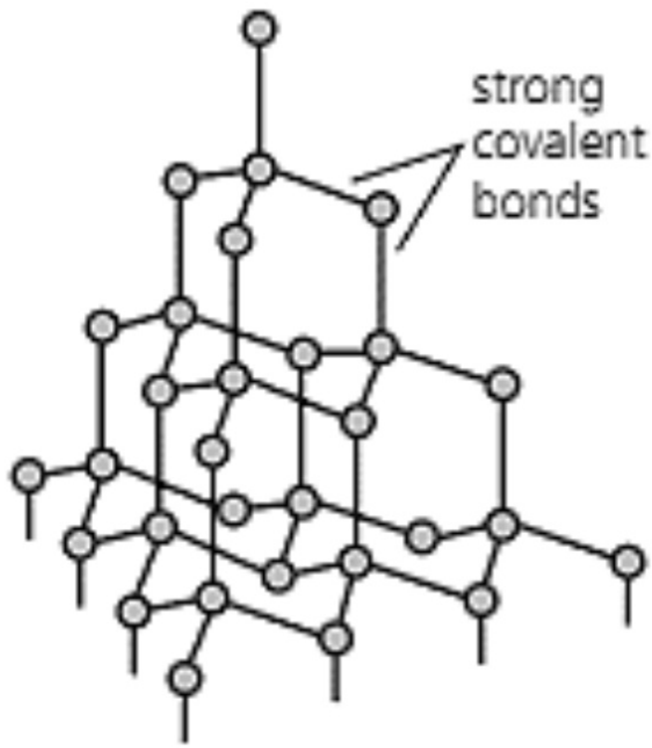
- Graphite and diamond are the two macromolecules of carbon that exist in different crystal structures in the same state. They are called

allotropes.

- **The existence of elements in more than one crystal structures in the same state is called allotropy (or polymorphism).**
- Each of the allotropes has a different structure in the same state and so the allotropes exhibit different physical properties.

Bonding in diamond structure

- In the giant structure of diamond, each of the carbon atoms is covalently bonded to four other carbon atoms as shown below.
- All the four outermost shell electrons of the carbon atoms are used to form covalent bonds, so there are no electrons available to enable diamond to conduct electricity.



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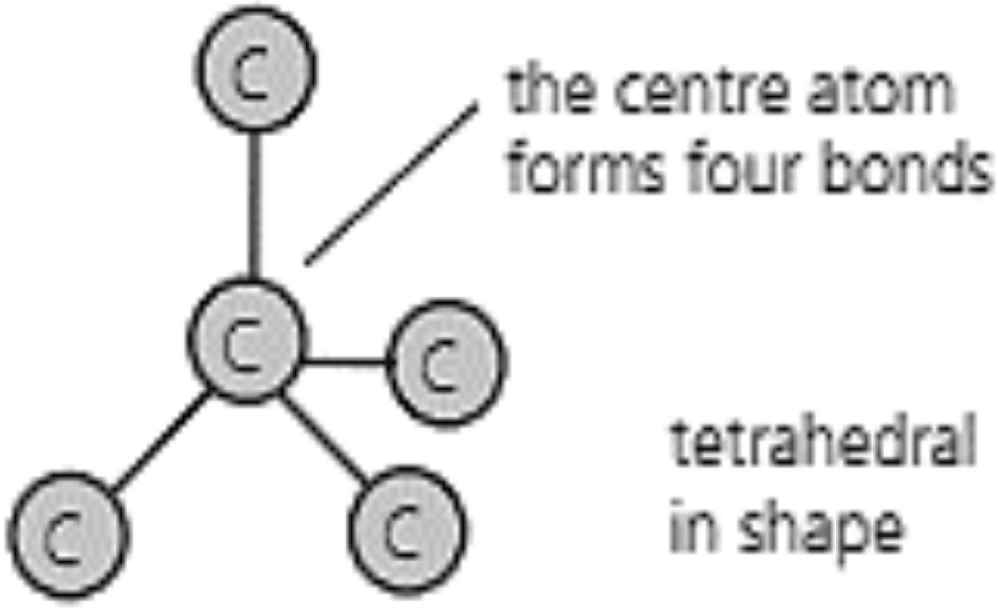


This bonding scheme gives rise to a very rigid, three-dimensional **tetrahedral structure**. Diamond is the hardest and has high melting and boiling point because of the four covalent bonds.

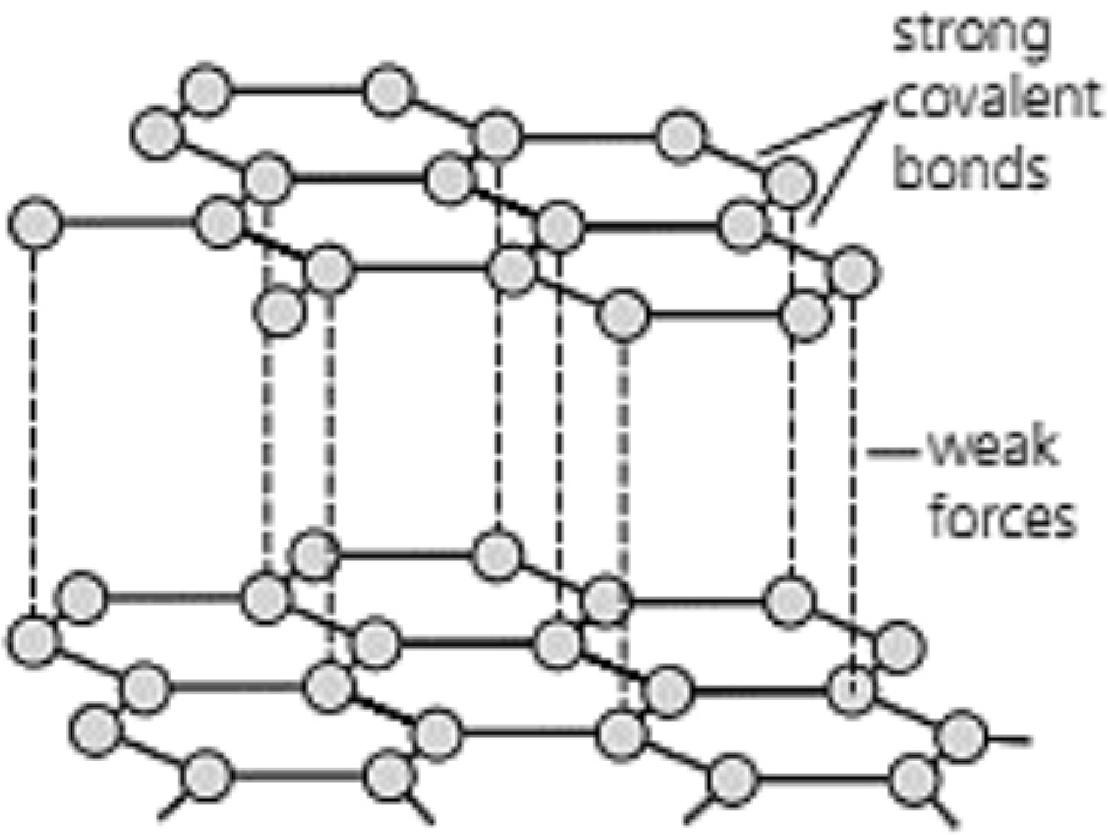
Structure of diamond

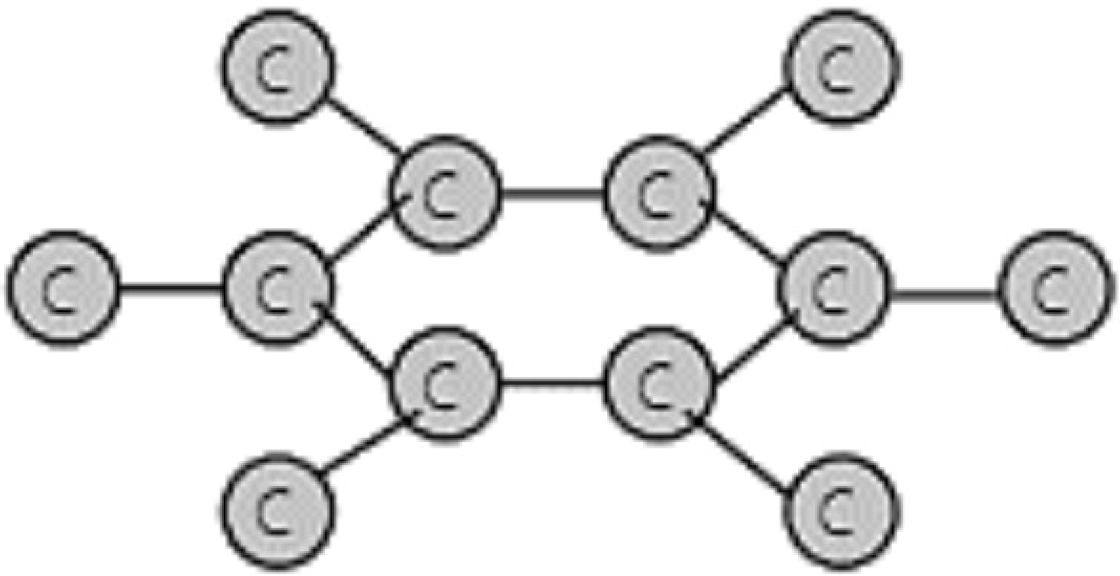
Bonding in graphite structure

- In the giant structure of graphite, each of the carbon atoms is covalently bonded to three other carbon atoms to form layers or rings of six atoms as shown below.
- The rings of atoms are held together over each in layers by weak forces called **van der Waals' forces**. The weak forces between layers makes graphite soft because the layers can slide over each other.
- The three covalent bonds in graphite causes graphite to have high melting and boiling points.
- With only three covalent bonds formed between carbon atoms

within the layers, an unbonded electron is present on each carbon atom.

- These ‘spare’ (or delocalised) electrons form electron clouds between the layers and it is because of these spare electrons that graphite conducts electricity.

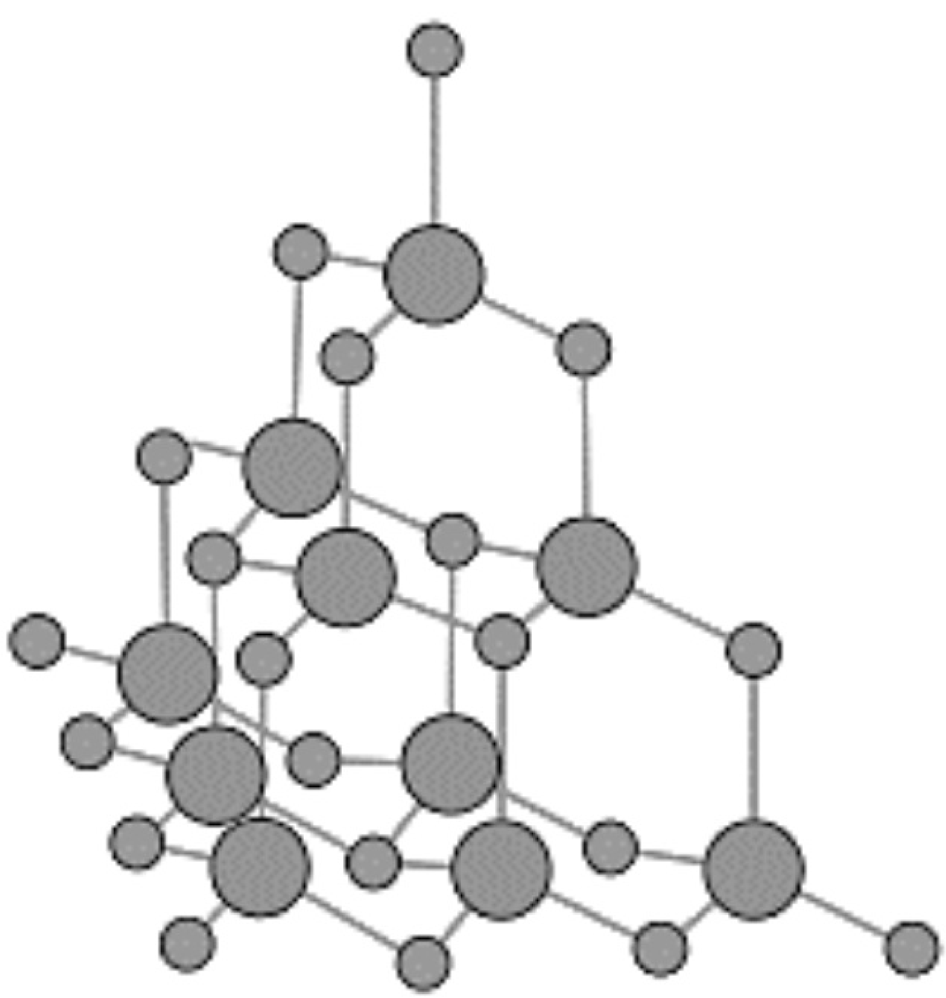




Structure of graphite

Bonding in silicon(iv) oxide structure

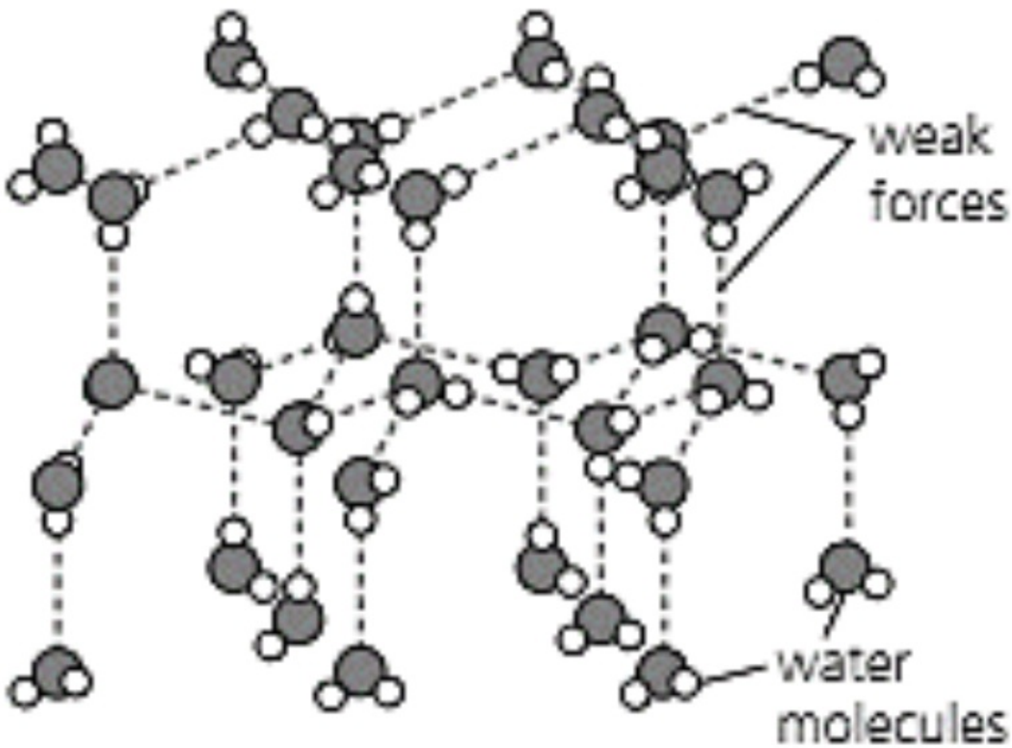
- In giant structure of silicon (iv), each of the silicon atoms is covalently bonded to four oxygen atom and each of the oxygen atoms is covalently bonded to two silicon atoms.
- The kind of bonding in silicon (iv) oxide give rise to a tetrahedral arrangement similar to that found in diamond.
- Silica (SiO_2) occurs naturally as quartz, the main mineral in sand.



- Because of the strong covalent bonds, silica is hard and has high melting and boiling points and does not conduct electricity.

Structure of silicon (iv) oxide

- One important reminder about covalent compounds in solid state is that they form a regular lattice structure made of billions molecules bonded together by **van der Waal's forces**.



- In ice, the water molecules are held in a regular lattice like this. But the forces between them are weak.

PROPERTIES OF IONIC AND COVALENT COMPOUNDS

- Ionic compounds are usually solids at room temperature due to strong electrostatic forces holding the crystal lattice together.
- Ionic compounds have high melting and boiling points due to the strong electrostatic forces holding the crystal lattice together. A lot of energy is therefore needed to separate the ions and melt the substance.

- Generally ionic compounds do not conduct electricity when solid, because the ions are not free to move.
- Usually ionic compounds conduct electricity when in the molten state or in aqueous solution because the forces of attraction between the ions are weakened and the ions are free to move to the appropriate electrode. This allows an electric current to be passed through the molten compound.
- Generally ionic compounds are soluble in water but insoluble in organic substances because water molecules are able to bond with both the positive and the negative ions, which breaks up the lattice and keeps the ions apart.

PROPERTIES COVALENT COMPOUNDS:

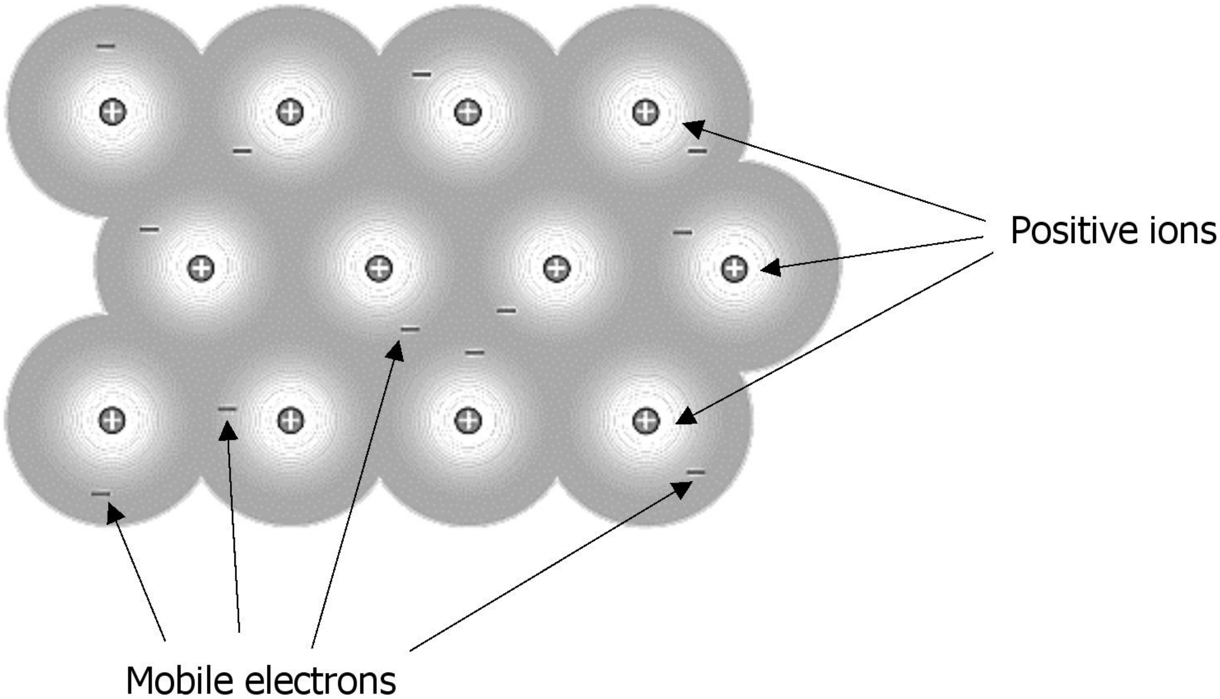
- Covalent compounds are usually gases, liquids or solids.
- Generally covalent compounds do not conduct electricity in molten or aqueous. This is because they do not contain ions. However, some molecules actually react with water to form ions. For example, hydrogen chloride gas produces aqueous hydrogen ions and chloride ions when it dissolves in water.
- Generally covalent compounds are insoluble water but soluble in organic solvents such as petrol.
- Generally simple covalent compounds have low melting and boiling points because of the weak intermolecular forces of attraction which exist between them.
- Generally giant covalent compounds such as diamond, graphite and silica have higher melting points because the whole structure is held together by strong covalent bonds.
- Generally covalent compounds are volatile at room temperature and pressure.

Metallic bonding

- Metallic bonding is an electrostatic force of attraction between the positively charged metal ions and the negatively charged free mobile (delocalized) electrons.
- The electrons in the outermost shell of the atom of a metal move

freely throughout the structure (they are delocalised) forming a mobile ‘sea’ of electrons.

- When outermost shell electrons are delocalised from the metal atoms, they form a lattice of stable positive ions. Positive metal ions then attract each other with free mobile electrons forming a metallic bond as shown below.



Properties of metals

- Metals usually have high melting points because metallic bonds are strong and requires a lot of energy to break.
- Metals are malleable and ductile. Malleable means they can be bent and pressed into shape. Ductile means they can be drawn out into wires. This is because when a force is applied across the metal, a layer can slide over each other without breaking due to the attraction between electrons and cations.
- Metals are good conductors of heat because the free electrons take in heat energy and move faster transferring the heat through the metal structure quickly.
- Metals are good conductors of electricity because the free electrons can move through the lattice carrying charge when a voltage is applied across the metal. Silver is the best conductor of all the metals. Copper is next – but it is used much more than silver because it is cheaper.

VALENCY AND VALENCE ELECTRONS

Valency electrons:

- This is the number of electrons an atom gains, loses or shares to form a compound or stable atoms.
- Nearly empty metal atoms lose electrons whereas nearly full non-metal atoms gain or share electrons to become stable.
- Valency electrons are determined by drawing electronic structures and counting the number of electrons in the outermost shell the atom gains, loses or shares.

Valence electrons:

- This is the number of electrons in the outermost shell of an atom.
- Valence electrons are determined by drawing electronic structures and then counting the number of electrons in the outermost shell.
- The number of valence electrons is equal to the group number to which an element belongs.

Examples

Determine the valency and valence electrons for the following atoms.

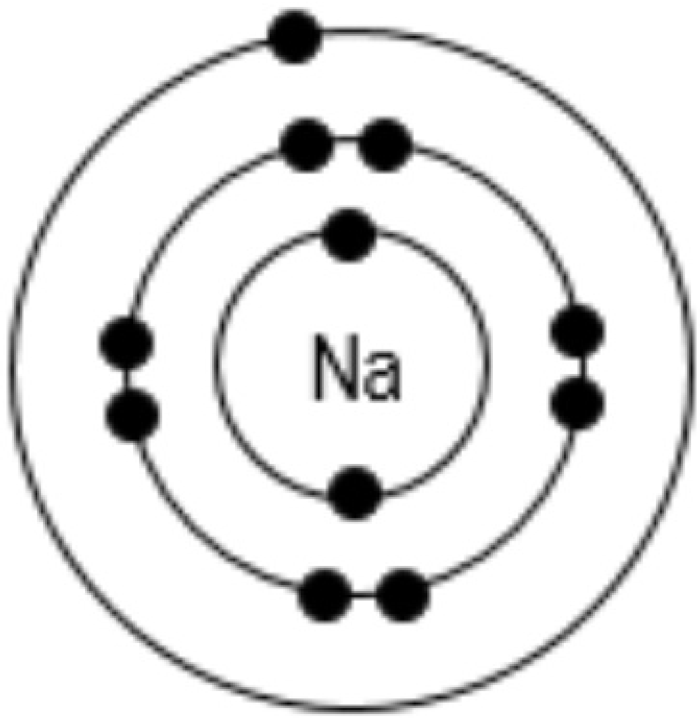
(i) Sodium atom (Na)
(ii) Magnesium atom (Mg)
(iii) Aluminium atom (Al)
(iv) Oxygen atom (O)
(v) Carbon atom (C)
(vi) Nitrogen atom (N)
(vii) Chlorine atom (Cl)
(viii) Argon atom (Ar)

Solutions

(i) Electronic structure of sodium atom (Na)

Sodium atom loses its one electron in the outermost shell to become stable, hence its valency electron is one.

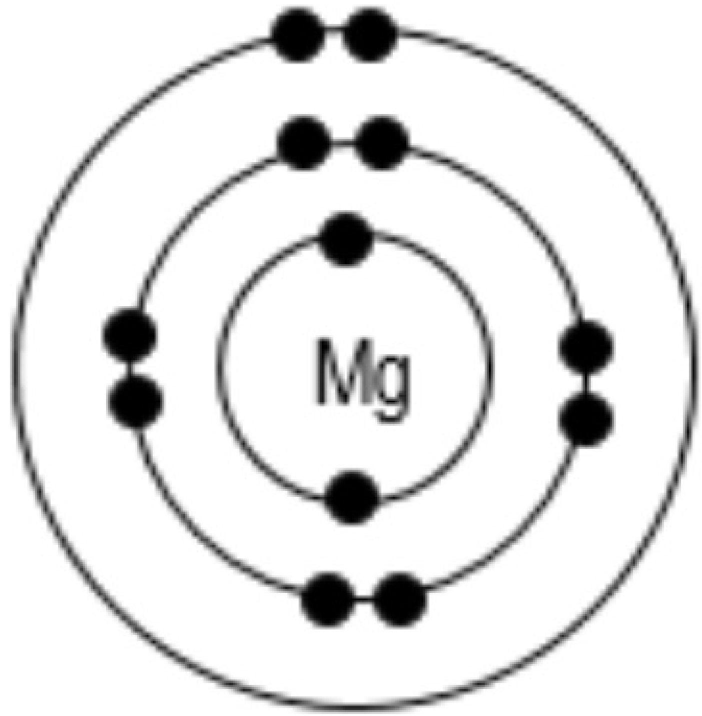
Sodium atom has one electron its outermost shell, hence its valence electrons one.



(ii) Electronic structure of magnesium atom (Mg)

Magnesium atom loses its two electrons in the outermost shell to become stable, hence its valency electron is two.

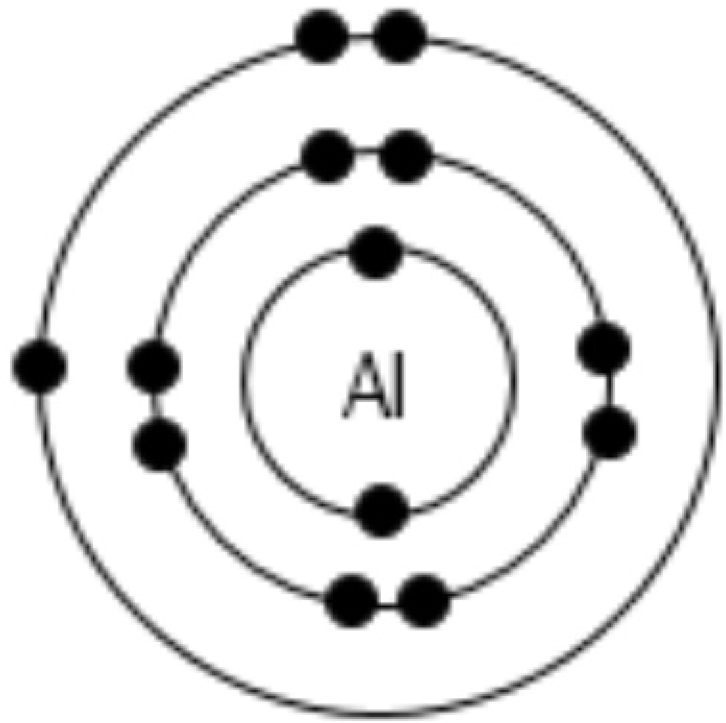
Magnesium atom has two electrons its outermost shell, hence its valence electrons two.



(iii) Electronic structure of aluminium atom (Al)

Aluminium atom loses its three electrons in the outermost shell to become stable, hence its valency electron is three.

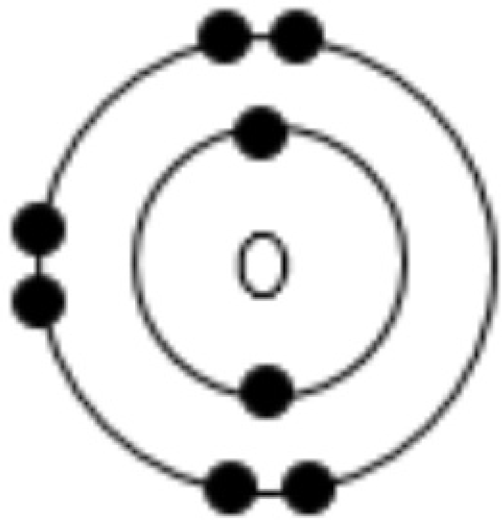
Aluminium atom has three electrons its outermost shell, hence its valence electrons three.



(iv) Electronic structure of oxygen atom (O)

Oxygen atom gains or share two electrons in the outermost shell to become stable, hence its **valency electron is two**.

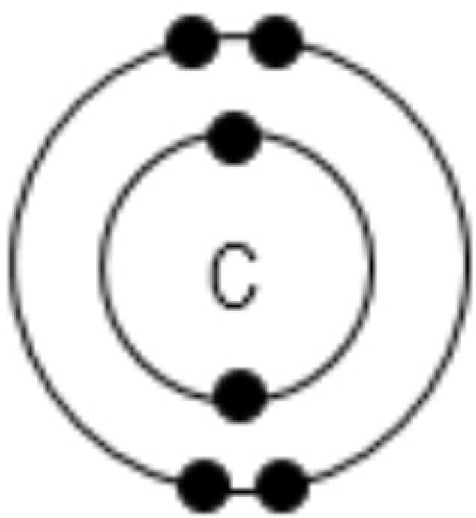
Oxygen atom has six electrons its outermost shell, hence its **valence electrons six**.



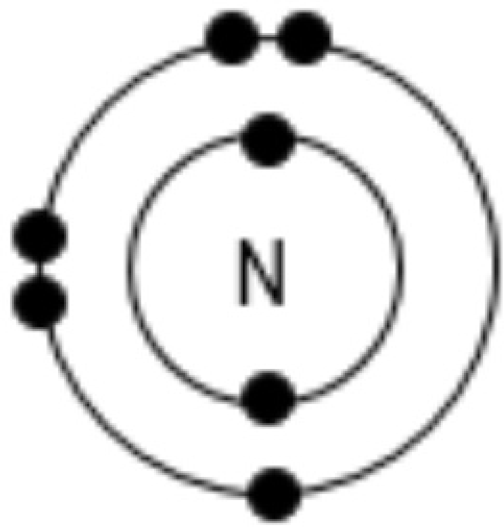
(v) Electronic structure of carbon atom (C)

carbon atom gains or share four electrons in the outermost shell to become stable, hence **its valency electron is four**.

carbon atom has six electrons its outermost shell, hence its **valence electrons four**.



(vi) Electronic structure of Nitrogen atom (N)



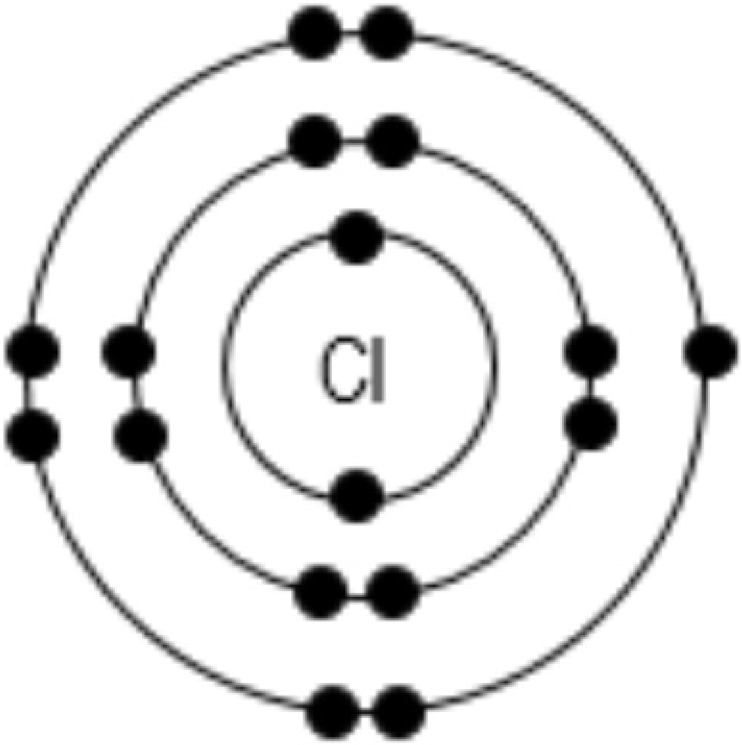
Nitrogen atom gains or share three electrons in the outermost shell to become stable, hence its **valency electron is three**.

Nitrogen atom has five electrons its outermost shell, hence its **valence electrons five**.

(vii) Electronic structure of Chlorine atom (Cl)

Chlorine atom gains or shares one electrons in the outermost shell to become stable, hence its **valency electron is one**.

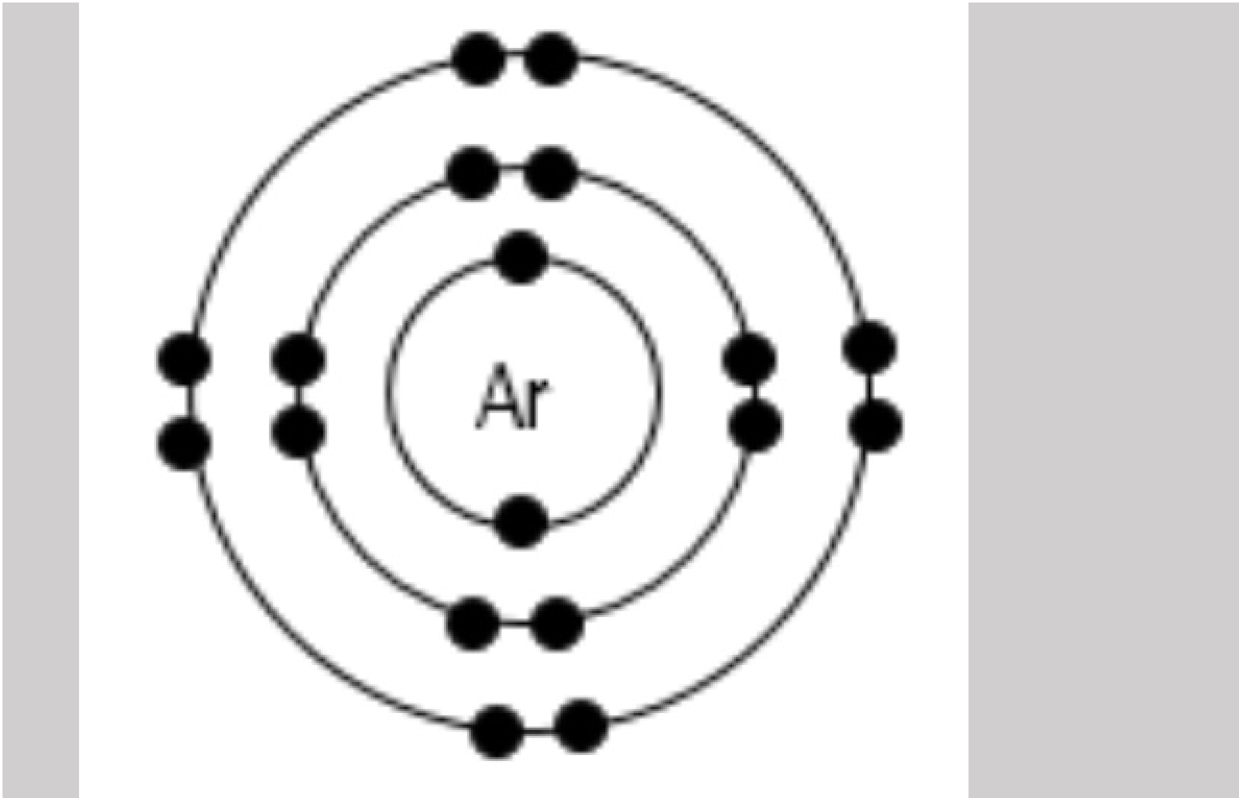
Chlorine atom has seven electrons its outermost shell, hence its **valence electrons seven**.



(viii) Electronic structure of Argon atom (Ar)

Argon atom is stable because it has eighty electrons in its outermost shell to become, hence its **valency electron zero**.

Argon atom has eighty electrons its outermost shell, hence its **valence electrons eighty**.


Ar

CHEMICAL FORMULAE OF COMPOUNDS

- A chemical formula is a short way of writing names of compounds using chemical symbols of atoms of elements present in the compound.
- The following steps are followed when writing chemical formulae of elements and compounds.

(i) Writing down the symbols of atoms of elements present in the compound.
(ii) Determining the valency electrons of atoms of elements present in the compound.
(iii) Writing the number of valency electrons of the atom of the element at the right top corner of the symbol of the atom.
(iv) Swapping the number of valency electrons so that the number of valency electron at the top right corner of one symbol takes a new position at the bottom left of another symbol.
(v) Reducing the resulting numbers to lowest ratio.

Examples

Determine chemical formulae of the following compounds.

(i) Sodium chloride
(ii) Sodium oxide
(iii) Aluminium chloride
(iv) Aluminium oxide
(v) Magnesium oxide
(vi) Carbon dioxide

Solutions

(i) Sodium chloride

Na Cl

Na^1 Cl^1

NaCl is the formula for sodium chloride.

(ii) Sodium oxide

Na O

Na^1 O^2

Na_2O is the formula for sodium oxide.

(iii) Aluminium chloride

Al Cl

Al^3 Cl^1

$AlCl_3$ is the formula for aluminium chloride.

(iv) Aluminium oxide

Al O

Al^3 O^2

Al_2O_3 is the formula for aluminium oxide.

(v) Magnesium oxide

Mg O

Mg^2 O^2

$Mg_{\cancel{2}}O_{\cancel{2}}$

MgO is the formula for magnesium oxide.

(vi) Carbon dioxide

C O

C^4 O^2

$C_{\cancel{2}}O_{\cancel{4}}$

CO_2 is the formula for carbon dioxide.

VALENCY ELECTRONS FOR RADICALS

- A radical is a group of atoms present in many compounds but incapable of existing independently.
- These radicles do not exist on the periodic table hence it is difficult to determine their valency electrons using the method outlined above.
- However, the valency electrons of radicals are already determined as shown in the table below.

Name of radical	Formula of radical	Valency electrons of radical
Nitrate	NO_3^-	1
Hydroxide	OH^-	
Ammonium	NH_4^+	
Nitrite	NO_2^-	
Carbonate	CO_3^{2-}	2
Sulphate	SO_4^{2-}	
Sulphide	SO_3^{2-}	
Phosphate	PO_4^{3-}	3
Phosphide	PO_3^{3-}	

EXAMPLES

Write down the chemical formulae of the following compounds:

(i) Calcium carbonate
(ii) Zinc nitrate
(iii) Aluminium hydroxide
(iv) Ammonium hydroxide
(v) Sodium sulphate
(vi) Magnesium carbonate

Solutions

(i) Calcium carbonate

Ca CO_3^{2-}

Ca^2 $(CO_3)^2$. We introduce blankets around the carbonate radical.

$Ca_{\cancel{2}}(CO_3)_{\cancel{2}}$

$CaCO_3$ is the formula for calcium carbonate.

(ii) Zinc nitrate

Zn NO_3^-

$Zn^2\ (NO_3)^1$

$Zn(NO_3)_2$ is the formula for zinc nitrate.

(iii) Aluminium hydroxide

Al OH^-

$Al^3\ (OH)^1$

$Al(OH)_3$ is the formula for aluminium hydroxide.

(iv) Ammonium hydroxide

NH_4^+ OH^-

$(NH_4)^1\ (OH)^1$

NH_4OH is the formula for ammonium hydroxide.

(v) Sodium sulphate

Na SO_4^{2-}

$Na^1\ (SO_4)^2$

Na_2SO_4 is the formula for sodium sulphate.

(vi) Magnesium carbonate

Mg CO_3^{2-}

$Mg^2\ (CO_3)^2$

$Mg_{\cancel{2}}(CO_3)_{\cancel{2}}$

$MgCO_3$ is the formula for magnesium carbonate.

WORD EQUATIONS

- **A word equation is a statement in words that shows a chemical change of reactants to products.**
- Reactants are substances that undergo chemical change while products are substances that are formed by chemical change.
- In a word equation, reactants are separated from products by an arrow sign and reactants are written on the left side of the arrow sign while products are written on the right side of the arrow sign.

Left side Right side

Reactant Product

- If there are many reactants changing into many products, then reactants are separated from each other by a plus sign just as products are separated from each other by a plus sign.

Reactant 1 + Reactant 2 Product 1 + Product 2

EXAMPLES

Write the word equations for the following reactions or chemical changes.

(i) Reaction between carbon and oxygen to form carbon dioxide.
(ii) Magnesium reacts with oxygen to form magnesium oxide.
(iii) Potassium reacts with water to produce potassium hydroxide and hydrogen gas.
(iv) Magnesium reacts with hydrochloric acid to produce magnesium chloride and hydrogen gas.

Solutions

(i) Carbon + oxygen → carbon dioxide
(ii) Magnesium + oxygen → magnesium oxide
(iii) Potassium + water → potassium hydroxide + hydrogen
(iv) Magnesium + water → magnesium hydroxide + water

CHEMICAL EQUATION

- A chemical equation is a short way of writing a word equation using chemical symbols of elements and chemical formulae of compounds.
- Reactants and products are accompanied by state symbols which show their states. The symbols used to represent the states include (g) for gaseous state, (l) for liquid state, (s) for solid state and (aq) for aqueous or solution state.

EXAMPLES

Write the following word equations into chemical equation.

(i) Carbon + oxygen → carbon dioxide

(ii) Magnesium + oxygen → magnesium oxide

(iii) Potassium + water → potassium hydroxide + hydrogen

(iv) Magnesium + water → magnesium hydroxide + water

SOLUTIONS

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

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

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

(iv) $Mg\ (s) + H_2O\ (l) \rightarrow Mg(OH)_2\ (aq) + H_2\ (g)$

BALANCING CHEMICAL EQUATIONS

- A balanced chemical equation is a chemical equation that has the same number of each type of atoms on both sides of the chemical equation.
- Chemical equations can be balanced by counting the number of each type of atoms on both sides of the chemical equation and making them equal.
- Each kind of atoms are balanced by adding numbers in front of chemical symbols that needed balancing.

Examples

Balance the following chemical equations.

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

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

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

(iv) $Mg (s) + H_2O (l) \rightarrow Mg(OH)_2 (aq) + H_2 (g)$

Solutions

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

One carbon atom one carbon atom

Two oxygen atoms two oxygen atoms

The equation is balanced.

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

One magnesium atom one atom of magnesium

Two oxygen atoms one atom of oxygen

The equation is not balanced on oxygen atom, add 2 in front of MgO to balance oxygen as shown below.

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

One magnesium atom two atoms of magnesium

Two oxygen atoms two atom of oxygen

The equation is not balanced on magnesium, add 2 in front of Mg as shown below.

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

two magnesium atoms two atoms of magnesium

Two oxygen atoms two atoms of oxygen

The equation is now balanced

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

One potassium atom one potassium atom

Two hydrogen atoms three hydrogen atoms

One oxygen atom one oxygen atom

The equation is not balanced on hydrogen atom, add 2 in front of KOH to make the number of hydrogen atom on the right side even and four as shown below.

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

One potassium atom two potassium atoms
Two hydrogen atoms four hydrogen atoms
One oxygen atom two oxygen atom

The equation is still balanced on potassium, hydrogen and oxygen on the left side of the equation, add 2 in front of K and H_2O to balance potassium, hydrogen and oxygen as shown below.

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

Two potassium atoms two potassium atoms
Four hydrogen atoms four hydrogen atoms
Two oxygen atoms two oxygen atom

The equation is now balanced.

(iv) $Mg\ (s) + H_2O\ (l) \rightarrow Mg(OH)_2(aq) + H_2\ (g)$

One magnesium atom one magnesium atom
Two hydrogen atoms four hydrogen atoms
One oxygen atoms two oxygen atom

The equation is not balanced on hydrogen and oxygen atoms, add 2 in front of H_2O to balance hydrogen and oxygen as shown below.

$$Mg\ (s) + 2H_2O\ (l) \rightarrow Mg(OH)_2(aq) + H_2\ (g)$$

One magnesium atom one magnesium atom
Four hydrogen atoms four hydrogen atoms
Two oxygen atoms two oxygen atom

The equation is now balanced.

NAMES OF SOME COMMON ACIDS

Name of acid	**Chemical formula of acid**
Hydrochloric acid	HCl
Nitric acid	HNO_3
Sulphuric acid	H_2SO_4
Carbonic acid	H_2CO_3
Ethanoic acid	CH_3COOH

BALANCING IONIC EQUATIONS

- **An ionic equation is an equation that shows ions that take part in a reaction when all the reactants and products are written in ionic form.**
- Reactants and products that are left out in a balanced ionic equation are called spectator ions.

Steps followed when balancing ionic equations includes;

1. Balance the given chemical equation.
2. Write only reactants and products in aqueous state as ions.
3. Cancel out ions that appear on both sides of the equation.
4. Write down the resulting equation as a balanced ionic equation.

EXAMPLES

Write down the balanced ionic equations for the following equations.

1. $NaOH\ (aq) + HCl\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$
2. $K_2SO_4\ (aq) + Ba(NO_3)_2\ (aq) \rightarrow BaSO_4\ (s) + KNO_3\ (aq)$
3. $H_2SO_4\ (aq) + KOH\ (aq) \rightarrow K_2SO_4\ (aq) + H_2O\ (l)$
4. $HCl\ (aq) + NH_4OH\ (aq) \rightarrow NH_4Cl\ (aq) + H_2O\ (l)$

SOLUTIONS

1. $NaOH\ (aq) + HCl\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$

The equation is balanced

Ionically:

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

Cancel ions appearing on both sides of the equation as shown below.

$\cancel{Na}^+\ (aq) + OH^-\ (aq) + H^+(aq) + \cancel{Cl}^-(aq) \rightarrow \cancel{Na}^+(aq) + \cancel{Cl}^-\ (aq) + H_2O(l)$

$\mathbf{OH^-\ (aq) + H^+(aq) \rightarrow H_2O(l)}$ is the balanced ionic equation.

2. $K_2SO_4\ (aq) + Ba(NO_3)_2\ (aq) \rightarrow BaSO_4\ (s) + KNO_3\ (aq)$

The equation is not balanced, so it balanced as shown below.

K_2SO_4 (aq) + $Ba(NO_3)_2$ (aq) → $BaSO_4$ (s) + $2KNO_3$ (aq)

$2K^+$(aq) + SO^{2-}_4(aq) + Ba^{2+}(aq) + $2NO^-_3$(aq) → $BaSO_4$(s) + $2K^+$(aq) + $2NO^-_3$ (aq)

$\cancel{2K^+(aq)}$ + SO^{2-}_4(aq) + Ba^{2+}(aq) + $\cancel{2NO^-_3(aq)}$ → $BaSO_4$(s) + $\cancel{2K^+(aq)}$ + $\cancel{2NO^-_3\ (aq)}$

SO^{2-}_4(aq) + Ba^{2+}(aq) → $BaSO_4$(s) is the balanced ionic equation.

Exercise.

3. H_2SO_4 (aq) + KOH (aq) → K_2SO_4 (aq) + H_2O (l)
4. HCl (aq) + NH_4OH (aq) → NH_4Cl (aq) + H_2O (l)

About The Author

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