INTRODUCTION TO CHEMISTRY

Chemistry is a branch of science that deals with the study of structure, composition and properties of matter. Chemistry as an experimental subject relies on:

- Careful handling of apparatus
- Carrying out chemical tests or chemical analysis
- Making critical observation
- Critical reporting of observations
- Drawing appropriate conclusions from observations

Chemistry as a practical science tries to answer the following question.
What are materials made up of?
How is matter formed?
How does matter behave?
Why does matter behave the way it does?

There are quite a number of branches of chemistry such as Organic chemistry, Physical chemistry, Inorganic Chemistry and Analytical chemistry.

Importance of chemistry

- The knowledge of chemistry makes us to understand the properties of substances and handle these substances with great care. Some may be poisonous, corrosive, toxic etc.
- Makes us to understand how to extract substances from the earth and use them e.g. gold, copper etc.
- Chemistry knowledge opens up way to science professional courses such as medicine, education and pharmacy.
- The knowledge of chemistry makes us to know the effects of chemicals on the environment and subsequently puts us in position to protect our environment.
- Broad knowledge in chemistry puts us in better positions to contribute towards more advancement in science and technology for better and quality human life.
- The knowledge of chemistry is relevant in many ways such as making of food supplements, distillation of fuel, making of plastics, making cosmetics and dental

creams, manufacture of soap and detergents, making insecticides and herbicides.etc.

LABORATORY

A laboratory is a special place with special equipments where scientific investigations or experiments are carried out. The laboratory is a place of adventure and discovery; in fact some of the most exciting events in the history of science have taken place in laboratories e.g. the discovery of oxygen.
When experiments are carried out in the laboratory, chemical materials and are used.

Chemicals are substances usually in solid or liquid forms that are consumed during reactions. They get usedup in reactions during the course of an experiment. E.g. Sodium hydroxide and Hydrochloric acid.

Apparatus are scientific tools that are used in carrying out scientific experiments. They are not usedup or consumed and can be used repeatedly. E.g. Beaker.

Laboratory safety rules

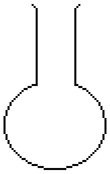
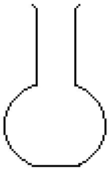
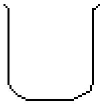
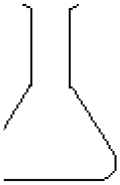
- Do not enter the laboratory without permission.
- Perform only the experiments assigned by the teacher.
- Learn at once the location and operation of fire extinguishers and other first aid materials.
- No equipment should be used until proper instructions are received from the teacher and you have proven proficiency.
- Avoid unnecessary movements in the laboratory.
- Do not play about with electrical or gas devices.
- Handle apparatus with great care.
- In case of accidents such as burns, cuts or splash of chemicals, wash with plenty of water and report to the teacher or laboratory technician.
- Do not taste or smell gases continuously.
- Do not eat or drink anything in the laboratory.
- When heating using boiling tubes or test tubes, do not turn the open end towards yourself or any one.
- Whenever in the laboratory, use clean and dry apparatus.

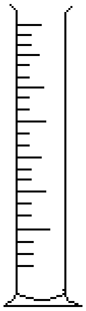
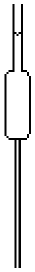
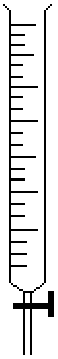
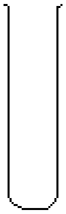
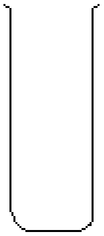
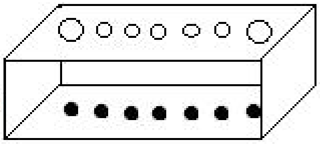
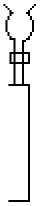
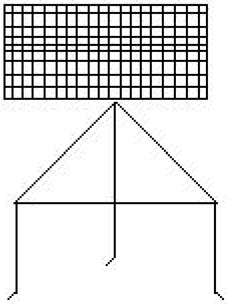
- Never use faulty equipments. Check and make sure that the devices you are given are not damaged.
- Never dispose off solid material into sinks instead throw them in solid waste containers or pits.
- Always make sure that you follow instructions that are given and read labels on bottles making sure that you use correct reagents.
- Do not leave reagent bottles open. As soon as you finish using a bottle, put its stopper and return it to its right position.
- Do not keep your work space too congested.

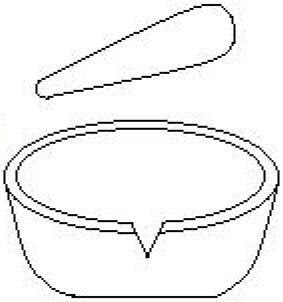
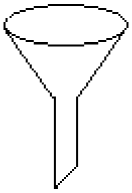
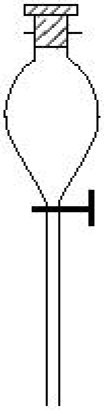
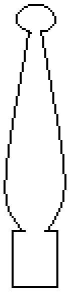
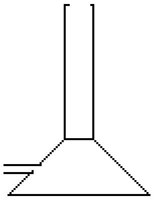
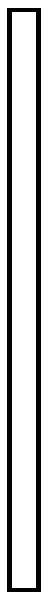
Simple laboratory equipments or apparatus

In order to carry out experiments in the laboratory, we need special tools such as measuring device. Such a tool is referred to as an apparatus.

Table of different laboratory apparatus and their functions.

Round bottomed flask Use -gas preparation when heating is involved -holding liquids 	**Flat bottomed flask** Use -gas preparation at room Temperature -holding liquids and mixing solutions. 
Beaker Use -holding reagents -mixing solutions and sometimes heating 	**Conical flask** Use -mixing reagents 

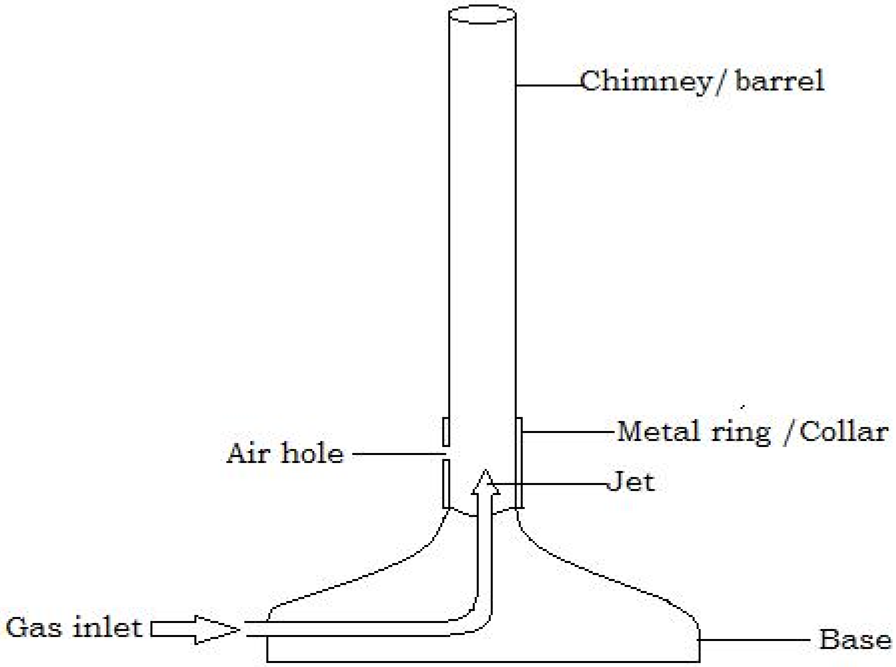
Measuring cylinder Use -measuring volumes of liquids 	**Pipette** Use -measuring a fixed volume of a liquid 
Burette Use -measuring exact volume of a liquid during titration 	**Test tube** Use -mixing small volumes of reagents -heating small volumes of liquid 
Boiling tube Use -heating small quantities of substances -holding small volumes of substance 	**Test tube rack** Use -Keeping test tubes 
Test tube holder Use -holding test tubes or boiling tube when heating 	**Tripod stand and wire gauge** Use -Supports apparatus during heating 

Mortar and pestle Use -Crushing solids to paste and powder 	**Filter funnel** Use -filtering i.e. separating liquid from solids -directing liquids into containers 
Separating funnel Use -separating immiscible liquids 	**Spatula** Use -scooping small amounts of substance 
Bunsen burner Use -heating substances 	**Glass rod** Use -stirring mixtures 

The Bunsen Burner

This is a piece of apparatus that is used for heating substances in the laboratory.

Structure



Functions of the labeled parts

Chimney/barrel: is a long metallic pipe through which the gas passes to be lit.

Collar/metal ring: is a metallic ring with a hole in it that regulates the amount of air entering the air hole.

Jet: is a tiny hole through which the gas enters the chimney. It is narrow so that the gas enters the chimney at high pressure.

Gas inlet: it connects the Bunsen burner to the gas supply.

Base: it is a heavy metal and used as a stand for the Bunsen burner i.e. it supports the Bunsen burner in an upright position.

How to light a Bunsen burner

1. Connect the Bunsen burner to the gas tap.
2. Close the air hole.
3. Open the gas on fully.
4. Light the Bunsen burner at the top of the chimney.
5. Open the air hole to get a hot non luminous flame used for heating.

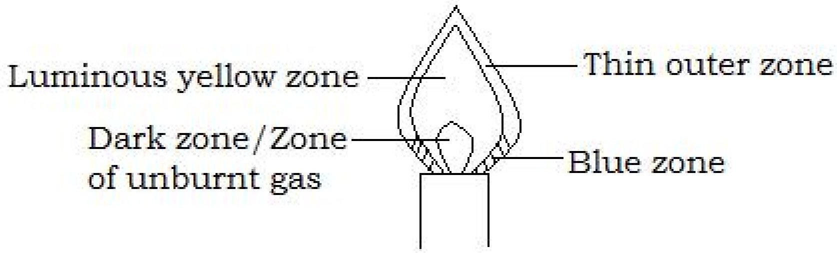
Bunsen burner flames

A flame is a burning gas that produces heat and light. A Bunsen burner produces two types of flames i.e. luminous and non luminous flames.

Luminous flame

This type of flame is produced when the air hole of the Bunsen burner in closed. The flame is large, bright, yellow and not very hot. Examples of luminous flame include; candle flame, bush fire flames and lantern flame

Structure



Zones of a luminous flame

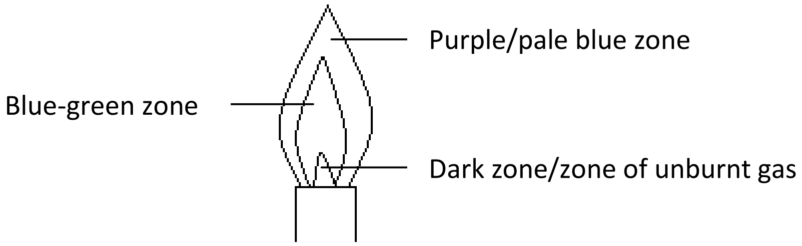
The luminous flame has four zones namely:

1. **Dark zone/ zone of un burnt gas;** burning of the gas does not take place here, it is dark and cool.
2. **Luminous yellow zone;** the gas burns in this zone but not completely because of limited air. Tiny particles glow in this region to give out light.
3. **Thin outer zone;** the zone is colourless and gas burns completely due to plenty of air.
4. **Blue zone;** this zone receives plenty of air and burning is more complete than in luminous zone.

Non luminous flame

This is the flame produced when the air hole of the Bunsen burner is open. Air enters through the air hole and mixes with the gas. This makes the gas to burn completely producing a very hot flame used for heating. The flame does not produce soot. Examples of non luminous flame include; gas cooker fame and stove flame.

Structure



This type of flame has three zones namely:

1. **Dark zone/zone of unburnt gas;** the zone is dark and the gas does not burn. It is therefore cool.
2. **Blue-green zone;** the gas here burns but not completely due to insufficient supply of air.
3. **Purple/pale blue zone;** burning of the gas take place completely as there is plenty of air, making this region very hot.

Differences between luminous and non luminous flames

Luminous flame	**Non luminous flame**
- Has four zone	- Has three zones
- Not very hot	- Very hot
- Larger is size	- Smaller in size
- Produces soot	- Does not produce soot
- Produces much light	- Produces less light
- Burns silently	- Burns with a lot of noise
- The flame is yellow and unsteady	- The flame is stable and steady

Similarities between luminous and non luminous flames

- Both flames have the dark zone (zone of un burnt gas)

- Both flames have luminous zone where the gas do not burn completely. In this zone, the carbon particles glow to give out light.
- Both flames have the non luminous zone where gas burns completely.

Strike back

Strike back occurs when the gas burns at the inner point jet or at the air hole instead of burning at the top of the chimney. It occurs normally when the air hole is left open when lighting the Bunsen burner. Strike back can cause explosion leading to destruction of property and lives.

MATTER

Matter is anything that occupies space and has weight. Matter is made up of very tiny particles known as ions, atoms or molecules.

An atom is the smallest electrically neutral particles of an element that takes part in a chemical reaction.

A molecule is the smallest electrically neutral particle of a compound or element that can exist on its own.

An ion is an electrically charged particle of an atom or a group of chemically combined atoms after losing or gaining electrons.

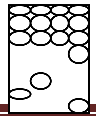
Matter exists in three state i.e. liquid, solid and gas.

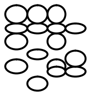
Properties of the states of matter

Solids

- Consist of particles that are fixed together i.e. particles that are not mobile.
- The particles are held by very strong forces of attraction.
- Solids have fixed shape.
- Solids have very high density.
- Solids are practically incompressible.

Arrangement of particles in a solid

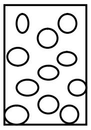




Liquids

- A liquid has a definite volume but no definite shape i.e. it takes up the shape of the container in which it is placed.
- Particles in a liquid are held together by weak intermolecular forces of attraction making the particles to move freely to some extent around each other.
- Liquids are slightly compressible
- Liquids are less dense compared to solids.
- Particles in a solid are relatively spread apart.

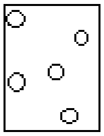
Arrangement of particles in a liquid



Gases

- Particles in a gas are far apart from each other.
- Particles in a gas are free to move randomly as they have negligible forces of attraction.
- Gases are easily compressible
- Gases have no definite shape and volume.
- Gases are very light.

Arrangement of particles in a gas



KINETIC THEORY OF MATTER

Kinetic theory of matter states that;

Particles that make up matter have kinetic energy and they are always in motion. The extent of the movement of the particles depends on the amount of kinetic energy the particles have.

In solids, the particles are held together by very strong intermolecular forces of attraction; the particles do not have enough kinetic energy to make them move from one place to another but they can vibrate in their mean position. When a solid is heated, the kinetic energy of the particles increase as they absorb the heat energy which weakens the forces of attraction between the particles. When the melting point is reached, the molecules break free and the solid changes to liquid.

In liquids, the particles are held together by weak forces of attraction. However, they have enough kinetic energy to enable the particles move from one place to another within the liquid. When a liquid is heated, the forces of attraction between the particles are weakened further until when they are completely broken and at this point, the liquid changes to a gas. This is the boiling point of the liquid.

In a gas, the particles are free to move randomly as they posses much kinetic energy and the particles are not held together by any particular forces of attraction (or negligible forces of attraction)

CHANGE OF STATES

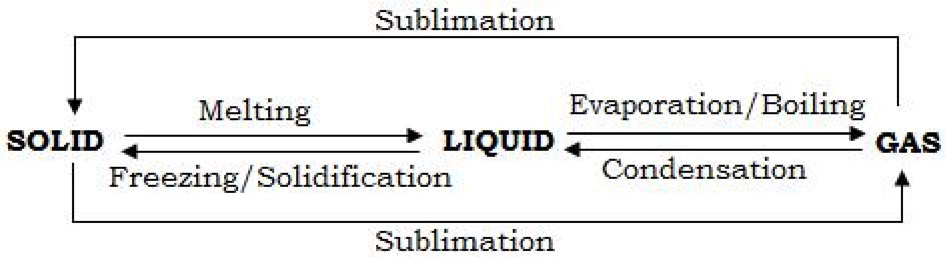
The process by which a solid changes into a liquid is the melting. The constant temperature at which a solid changes into a liquid is called melting point

The process by which a liquid changes into a gas is called boiling or evaporation. The constant temperature at which a liquid changes into a gas is the **boiling point**.

The process by which a liquid changes to a solid is freezing or solidification.
The constant temperature at which a liquid changes into a solid is referred to as **freezing point**.

The process by which a gas changes into a liquid is called condensation.
The process by which a gas changes directly to a solid is referred to as sublimation and vice versa.

Summary of the changes of states



Experiments to demonstrate that particles in liquids and gases move

a) Brownian motion

This is the continuous random movement/motion of solid particles in liquids and gases.

1. Demonstration of Brownian motion in Liquids

When pollen grains are poured in water, they are seen to be moving in a continuous random zigzag manner. The movement of the pollen grain is due to bombardment of the particles by the moving particles of water.

2. Demonstration of Brownian motion in gases

When smoke particles are trapped in a glass cell and observed under a microscope, the particles are seen to be moving in a random zigzag manner. The movement of the smoke particles is due to bombardment by the moving gas particles.

Or

When a beam of light is directed into a dark room, the dust particles are seen to be moving in a continuous random manner. The movement of the dust particles is due to the bombardment of these particles by gas particles.

b) Diffusion

This is the spreading of particles or molecules from a region of high concentration to a region of low concentration.

1. **Demonstration of diffusion in liquids**

   Place a crystal of potassium permanganate in a beaker of water and watch. After sometimes, the water turns pink due to the particles of potassium permanganate spreading throughout the water.

2. **Demonstration of diffusion in gas**

   Open a bottle of concentrated ammonia solution and place the bottle at the corner of a room. After sometimes, ammonia smell will spread through out the room due to diffusion of the ammonia particles.

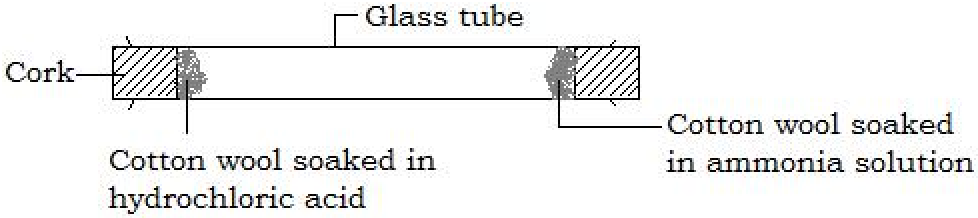
Rate of diffusion of gases

The rate of diffusion of a gas depends on;

- Density of the gas. The lighter the gas, the higher the rate of diffusion.
- Density of the diffusion medium. The lighter the density of the diffusion medium, the faster is the rate of diffusion.
- Concentration gradient. The steeper the concentration gradient, the higher the rate of diffusion.
- Surface area. The smaller the surface area of diffusion, the higher the rate of diffusion.
- Temperature of the diffusion medium. The higher the temperature, the faster is the rate of diffusion of the particles.
- Size of the particles. Smaller particles diffuse faster the larger particles.
- Distance through which diffusion occurs. The smaller the diffusion distance, the faster is the diffusion.

Expt. to compare the rates of diffusion of ammonia and hydrogen chloride gas

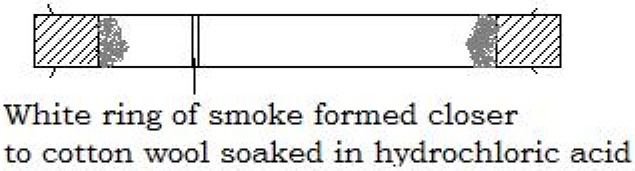
Diagram



Procedure

- Soak a piece of cotton wool in concentrated hydrochloric acid and another piece of cotton wool in concentrated ammonia.
- Insert the cotton wools with concentrated hydrochloric acid and concentrated ammonia at both ends of the glass tube simultaneously as shown above.
- Close both ends of the tube with cork.

Observation



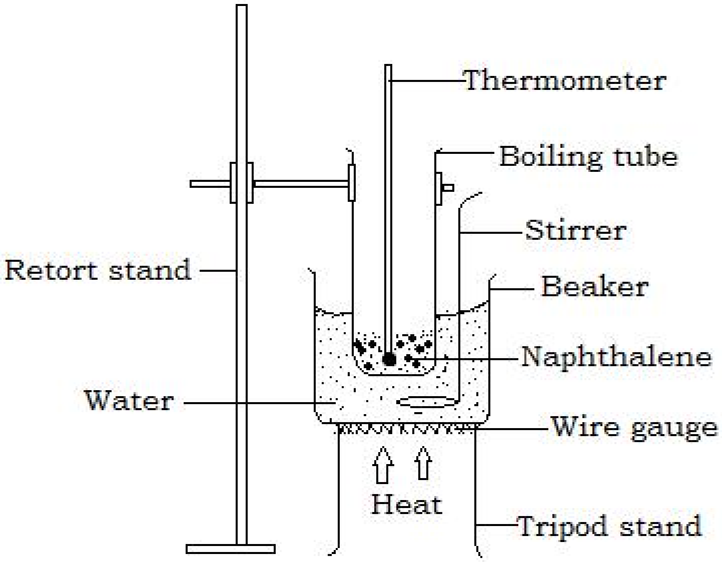
After some times, a white ring of smoke of ammonium chloride is formed in the tube, indicating the ammonia and hydrogen chloride gas diffused through some distance before they met. The white ring was formed closer to the end of the tube with cotton wool containing hydrochloric acid indicating that ammonia diffused through a longer distance than hydrogen chloride.

Conclusion

Ammonia gas diffuses faster than hydrogen chloride gas.

Determination of the melting point of naphthalene

Drawing of the setup



Procedure

- Place some naphthalene in a boiling tube and insert a thermometer in it.
- Place the boiling tube in a beaker of water and arrange the apparatus as shown above.
- Heat gently the water in the beaker while stirring.

When the naphthalene begins to melt, stir with the thermometer and note the steady temperature at which the solid naphthalene melts. This is taken to be the melting point of naphthalene.

Note

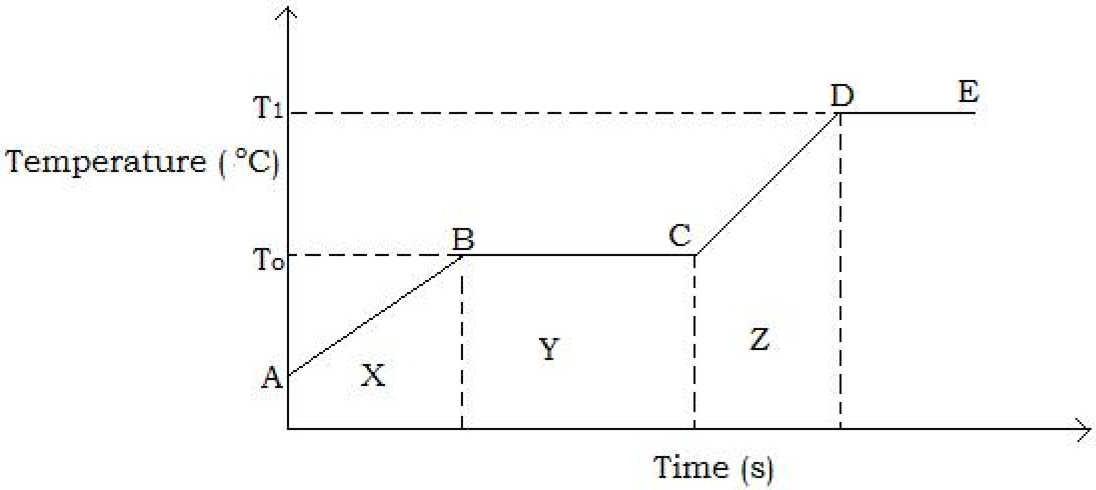
1. Water bath with a stirrer is used so as heat is distributed uniformly to all parts of naphthalene
2. The water is heated gently to prevent violent boiling of the water.

Factors that affect melting point

- **Impurity.** The presence of an impurity lowers the melting point of a substance.e,g. pure naphthalene melts at 80 °C but impure naphthalene containing benzoic acid melts at a temperature less than 80 °C. A mixture of ice and common salt cools faster than pure ice because common salt acts as an impurity which lowers the freezing point of ice.
- **Pressure.** Pressure affects the melting points of substances differently. Substances that decrease in volume when melted e.g. water, their melting points are decreased with increase in pressure. But substances that increase in volume on melting e.g. tin, paraffin wax, their melting points are increased with increase in pressure.

Graph of temperature against time for heating naphthalene

When solid naphthalene is heated and its temperature recorded after a certain time interval, a graph of temperature against time as below is obtained.



Points and regions on the graph

A is a point at which heating of naphthalene starts

B is the point at which melting of naphthalene starts

C is the point at which melting of naphthalene stops

D is the point at which boiling of naphthalene starts.

X is a region where naphthalene exists in solid states.

Y is a region in which both solid and liquid naphthalene exists.

Z is a region where liquid naphthalene exists.

Along AB, the temperature of the solid naphthalene increases steadily as the solid naphthalene is heated.

Along BC, the melting point is reached and the temperature remains constant. The corresponding temperature, T_o is the melting point of naphthalene. The heat absorbed is used to break the intermolecular forces of attraction in solid naphthalene turning it into a liquid.

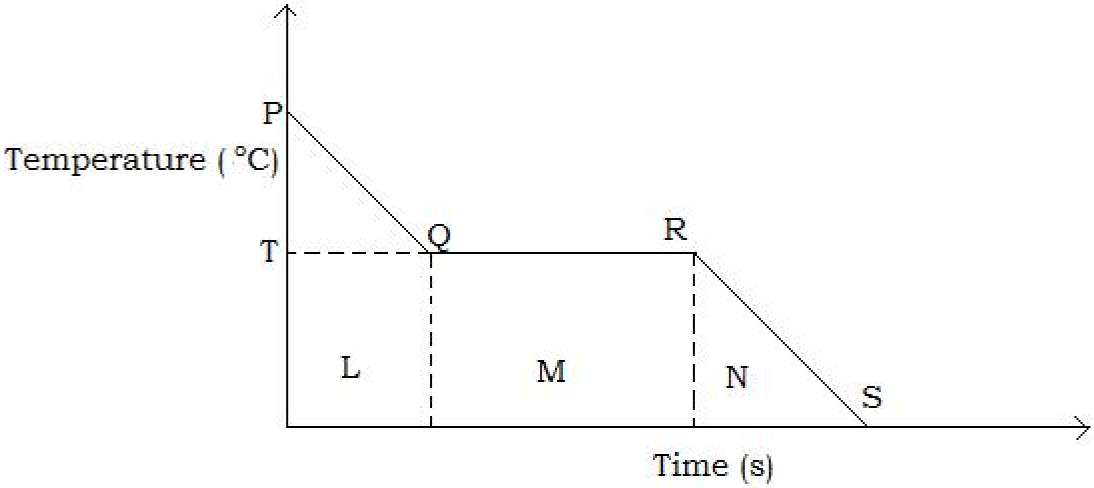
Along CD, the temperature of the liquid naphthalene rises steadily.

Along DE, the boiling point of the naphthalene is reached and the temperature remains constant. The corresponding temperature T_1 is the boiling point of naphthalene.

Freezing point of naphthalene

When solid naphthalene has fully melted, allow the naphthalene to cool slowly in air while stirring with a thermometer. As the liquid cools it becomes cloudy and temperature remains constant as the liquid changes to solid. Record this temperature as the freezing point of naphthalene.

A graph of temperature against time for cooing of a liquid naphthalene



Points and regions of the graph

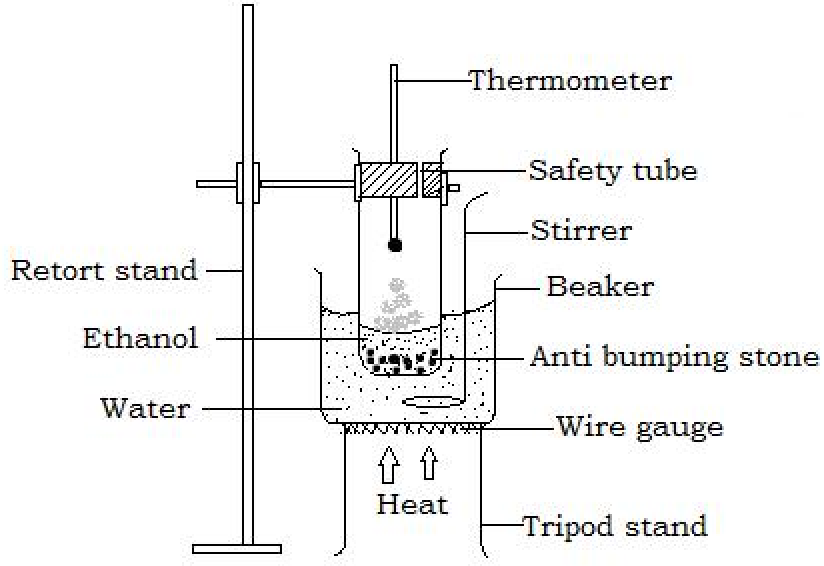
P is a point at which cooling of liquid naphthalene starts.
Q is the point at which freezing of naphthalene starts.
R is the point at which freezing of naphthalene stops.
S is the point at which cooling of solid naphthalene stops.

Along PQ, temperature of the liquid naphthalene falls.
Along QR, the freezing point of naphthalene is reached and temperature remains constant as liquid changes to solid. The corresponding temperature, T is the freezing point of naphthalene,
Along RS, the temperature of the solid naphthalene falls steadily.

L is a region where liquid naphthalene exists
M is a region where there is both solid and liquid naphthalene.
N is a region with solid naphthalene.

Determination of boiling point of ethanol

Drawing of the set up



Procedure

- Place about $5cm^3$ of ethanol in a boiling tube and add some anti bumping stones.
- Suspend a thermometer to about 2cm above the surface of ethanol and place the boiling tube in a beaker of water as shown above.
- Heat the beaker gently and record the temperature at which ethanol boils, this occurs when ethanol vapor condenses at the bulb the thermometer and drops back. This is recorded as the boiling point of ethanol.

N.B.

1. If the thermometer is dipped into the ethanol, it will be measuring the internal temperature of ethanol instead of the boiling point.
2. The safety tubes allow excess ethanol vapor to escape,
3. Water bath in used for heating ethanol, because;
    - Ethanol is highly flammable (catches fire easily) and cannot be heated directly
    - The water bath distributes heat evenly (uniformly).

Factors that affect boiling pot of liquids

1. Atmospheric pressure. The higher the atmospheric pressure, the higher he boiling point, this is because boiling takes place then the vapor pressure of the liquid is equal to the atmospheric pressure. Atmospheric pressure varies with altitude. The lower the atmospheric pressure, the lower the boiling point.

2. Dissolved impurity. The presence of dissolved impurity increases the boiling point of the liquid e.g. water boils at 100 ^0C at one atmospheric pressure but water containing dissolved salts boils at a temperature higher than 100 ^0C at one atmosphere.

PHYSICAL (NON PERMANENT) AND CHEMICAL (PERMANENT) CHANGES

a) CHEMICAL (PERMANENT) CHANGE

This is a change in which a new substance is formed and the change is irreversible i.e the substance formed cannot be turned back to its original form.

Characteristics of chemical changes

- Great heat changes occur i.e. heat is either absorbed or released.
- Changes in mass occur i.e. there is increase or decrease in the mass of the product formed.
- New substances are formed during chemical changes.
- The reaction is not reversible (irreversible).

Examples of chemical changes include: burning of charcoal; rusting of iron; burning of substances in air e.g. magnesium; addition of water to calcium oxide; explosion of natural gas or hydrogen with air.

b) PHYSICAL (NON PERMANENT) CHANGE

This is a change in which no new substance is formed and the change is reversible.i.e the substance formed can be converted back to its original form.

Characteristics of physical changes

- The reaction is not accompanied by great heat changes(except latent heat changes accompanying the reaction)
- No new substance is formed
- No changes in mass occurs i.e. the mass of the product is the same as that of the reactant.

- The reaction is reversible.

Examples of physical changes include: melting of ice; vaporization of a liquid; magnetization of iron; sublimation of iodine.

Differences between physical and chemical changes

Chemical change	**Physical change**
- A new substance is formed	No new substance is formed
- The reaction is not reversible	The reaction is reversible
- Changes in mass occur	No changes in mass occur
- Considerable heat changes occur	No considerable heat changes occur

Effects of heat on some substances

a) Magnesium

When magnesium is burnt, it burns with a brilliant white flame producing white fumes and white ash.

b) Iodine

When the grey shinny solids of iodine are heated in a test tube, it forms a purple vapor which settles as grey shinny solids at the cooler parts of the test tube.

c) Candle wax

When candle wax is heated in a test tube, it forms a colorless liquid which turns back into a solid on cooling.

d) Sulphur

When sulphur is heated gently in a test tube, it melts into a brown liquid and on cooling, it turn to yellow solids. However, when heated strongly, itburns with a blue flame producing a colorless gas with a choking smell.

e) Potassium permanganate

When potassium permanganate is heated in a boiling tube, it decomposes with a cracking sound producing black fumes and black powder.

f) Copper (II) sulphate crystals (Hydrated)

When the blue copper (II) sulphate crystals are heated in a test tube, a colorless vapor condenses at the cooler part of the test tube forming a colorless liquid and the blue crystals turn to white powder. However when water is added to the white powder, blue crystals of copper (II) sulphate are formed back.

g) Zinc oxide

When white powder of zinc oxide is heated in a boiling tube, it turns to yellow and on cooling in turns back to white.

h) Nicrome wire

On heating nicrome wire, it glows white hot and on cooling turns back into its original state.

i) Lead (II)nitrate

When the lead (II) nitrate is heated in a test tube, it burns with a cracking sound and melts into brown liquid producing fumes of brown gas with an irritating smell and finally yellow residue is left in the test tube.

CLASSIFICATION OF SUBSTANCES

In chemistry, substances are classified as elements, compounds and mixtures.

ELEMENTS

An element is a pure substance that cannot be split or divided into two or more simpler forms by any known chemical means. An element is made up of atoms.

Examples of elements with their symbols (first twenty elements)

Name of element	Symbol of element	Name of element	Symbol of element
1. Hydrogen	H	11. Sodium	Na
2. Helium	He	12. Magnesium	Mg
3. Lithium	Li	13. Aluminium	Al
4. Beryllium	Be	14. Silicon	Si
5. Boron	B	15. Phosphorus	P
6. Carbon	C	16. Sulphur	S
7. Nitrogen	N	17. Chlorine	Cl
8. Oxygen	O	18. Argon	Ar
9. Fluorine	F	19. Potassium	K
10. Neon	Ne	20. Calcium	Ca

Other examples of elements and their symbols include: Copper (Cu);Zinc (Zn);Iron (Fe);Mercury (Hg);Lead (Pb);Gold (Au);Bromine (Br)Cobalt (Co);Manganese (Mn).

Naming of elements

In naming elements, the 1st letter or the 1st two letters are used as symbols for the elements.

Examples

Element	Symbol	Element	Symbol
Hydrogen	H	Carbon	C
Helium	He	Argon	Ar
Beryllium	Be	Calcium	Ca

Fluorine	F	Oxygen	O

However, there are elements whose symbols are different from the 1st letters of their names. This is because; the symbols used are of old latin names. Examples of these elements include the following.

Element	**Latin name**	**Symbol**
Sodium	Natrium	Na
Potassium	Kalium	Ka
Copper	Cuprum	Cu
Iron	Ferrum	Fe
Lead	Plumbum	Pb
Gold	Aurum	Au
Silver	Argentum	Ag
Mercury	Hydrargyrum	Hg

COMPOUNDS

A compound is a substance made up of two of more elements chemically combined together. Examples of compounds include the following.

Name of compound	**Formula**	**Elements contained**
Water	H_2O	Hydrogen and oxygen
Ammonia	NH_3	Nitrogen and hydrogen
Carbondioxide	CO_2	Carbon and oxygen
Sodium chloride	$NaCl$	Sodium and chlorine
Copper(II)sulphate	$CuSO_4$	Copper, sulphur and oxygen
Potassium permanganate	$KMnO_4$	Potassium, manganese and oxygen
Sulphuric acid	H_2SO_4	Hydrogen, sulphur and oxygen
Calcium carbonate	$CaCO_3$	Calcium, carbon and oxygen
Lead(II)nitrate	$Pb(NO_3)_2$	Lead, nitrogen and oxygen
Oxygen gas	O_2	Oxygen

During the formation of compounds (chemicals combination of elements), atoms of elements combine in definite proportions and energy in the form of heat, light or electrical form is either released or absorbed.

MIXTURES

A mixture is a substance made up of two or more elements or compounds physically combined together.

Classes of mixture

- Solid-solid mixture e.g. copper and tin, salt and sugar, sand and salt.
- Solid-liquid mixture e.g. sand and water, chalk and water.
- Liquid-liquid mixture e.g. ethanol and water, oil and water.
- Liquid-gas mixture e.g. water and air.
- Gas-gas mixture e.g. carbondioxide and oxygen, nitrogen and oxygen.

SEPARATION OF MIXTURES

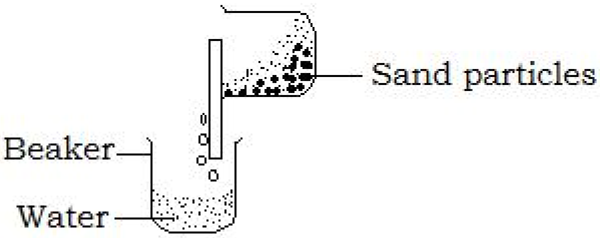
Mixtures can be separated by a number of physical methods depending on the properties of the components such as solubility, boiling point and magnetic properties. Some common methods of separating mixtures include: decantation, filtration, centrifuge method, sublimation method, use of a separating funnel, evaporation, simple distillation, fractional distillation, separation by a magnet and chromatography.

a) **Separation of insoluble solids from liquids** e.g. sand from water. The methods used are: Decantation, centrifuging and filtration.

1. Decantation

The mixture of the solid (sand) and liquid (water) is poured in a beaker and allowed to stand for a while. The solids (sand) being denser will settle at the bottom of the beaker while the liquid (water) remains on top as it is less denser than sand. The liquid (water) is carefully poured into another beaker leaving sand in the beaker that had the mixture.

Drawing to illustrate

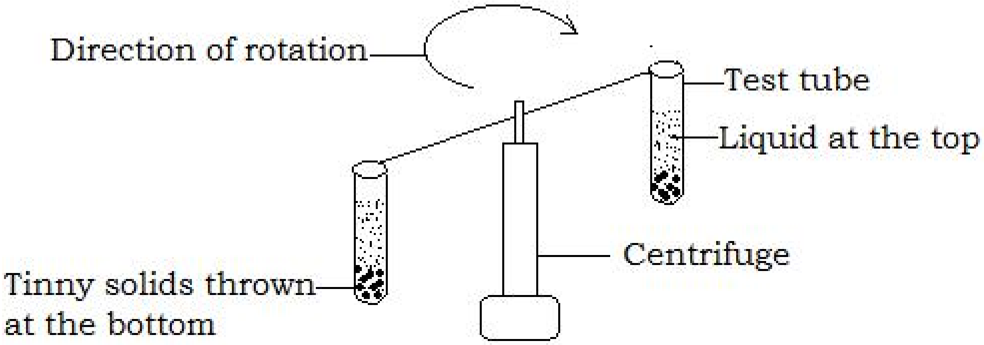


Note:

Sometimes, the solid particles may be very tinny and may take time to settle at the bottom therefore separation by decantation may be very difficult. For such a mixture, the solids can be made to settle faster at the bottom of the container by using a **centrifuge**. A centrifuge is a machine that speeds up the process of sedimentation (settling of solids).

During centrifuging, test tubes containing the mixture are spun at a very high speed and the solids are thrown at the bottom of the test tubes. Separation by decantation can then be possible.

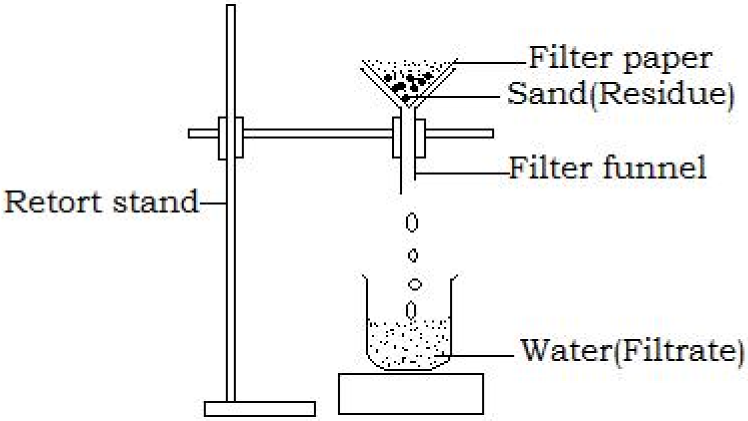
Drawing to illustrate



2. Filtration

During filtration, the mixture of the solid (sand) and liquid(water) is poured through a filter funnel containing a filter paper. The solid particles remain trapped on the filter paper and is called **residue** while the liquid passes through the filter paper and is collected in another beaker as the **filtrate**.

Drawing to illustrate



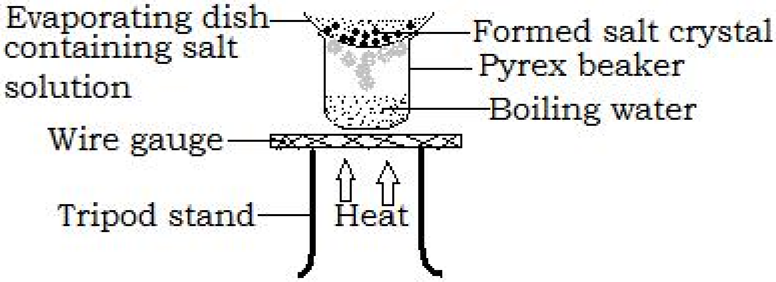
b) **Separation of a solute from a solvent** e.g. sugar (salt) from water. Methods used include: evaporation and simple distillation.

A **solute** (e.g. salt) is a substance that dissolves in a **solvent** (e.g. water) and a solvent is a substance that dissolves a solute. A mixture of the solute and solvent forms a **solution** (e.g. salt solution).

3. Evaporation method

In this method, the solution is put in an evaporating dish and heated gently over a steam bath as shown below.

Drawing to illustrate



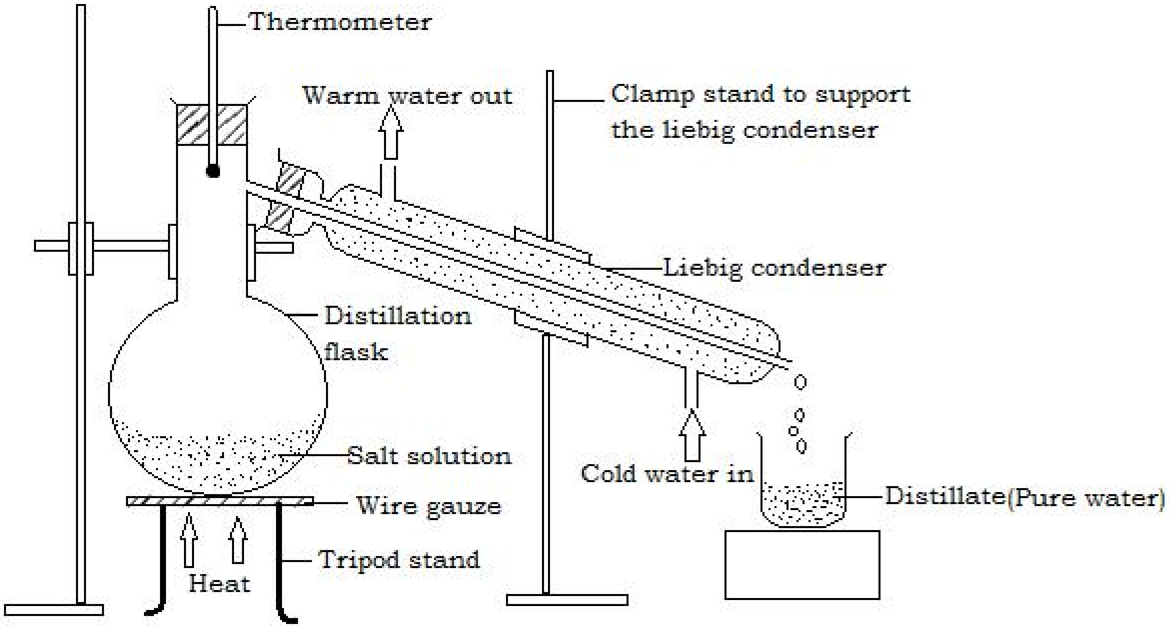
Continue heating until when all the water has evaporated. As the water evaporates, crystals of the salt or sugar are formed in the evaporating dish (crystallization occurs).

4. Simple distillation

Distillation is the process of heating a liquid to form vapor and condensing the vapor to form a more pure liquid. This method can be used to separate solutes from solvents and for separating liquid mixtures whose boiling points are wide apart(i.e. separating liquids with a big difference in boiling point.) such as ink and water.

Separating salt from water by simple distillation

The salt solution is placed in a distillation flask and the apparatus arranged as shown below.



When the mixture is heated, the solvent (water) being more volatile will escape and its vapor condensed by the Liebig condenser forming the **distillate**. The solute (salt) will remain at the bottom of the distillation flask as crystals.

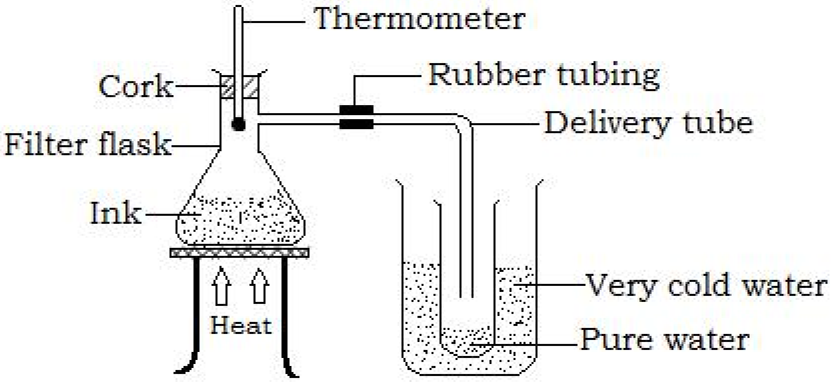
c) Separation of miscible liquids

Miscible liquids are liquids that can be mixed together in all proportion forming a uniform layer. The uniform mixtures of two or more liquids form a **homogenous mixture.** Examples of uniform mixtures are: ethanol and water, ethanol and propanol, water and ink e.t.c. the homogenous mixtures (uniform mixtures) can be separated by simple distillation (when the differences in boiling points of the liquids are wide) and fractional distillation (when the boiling point difference of the liquids is narrow).

Separation of a mixture of ink and water by simple distillation

Procedure

- Pour a mixture of ink and water into a filter flask fitted with a delivery tube to its side arm and insert a rubber bung carrying a thermometer to its top as shown below.



- The delivery tube at the side arm is connected to an empty test tube placed in cold water (the cold water is to condense the gas vapor)
- Heat the mixture (ink solution) carefully and note the temperature at which the solution is boiling. As heating in continued, the water vapor escapes from the solution since its boiling point is lower and condenses in the test tube. The temperature remains steady and this will be recorded as the boiling point of water.
- Slightly above the steady temperature, heating is stopped as the water has been separated from the ink. The water is collected as a distillate and the ink remains in the filter flask.

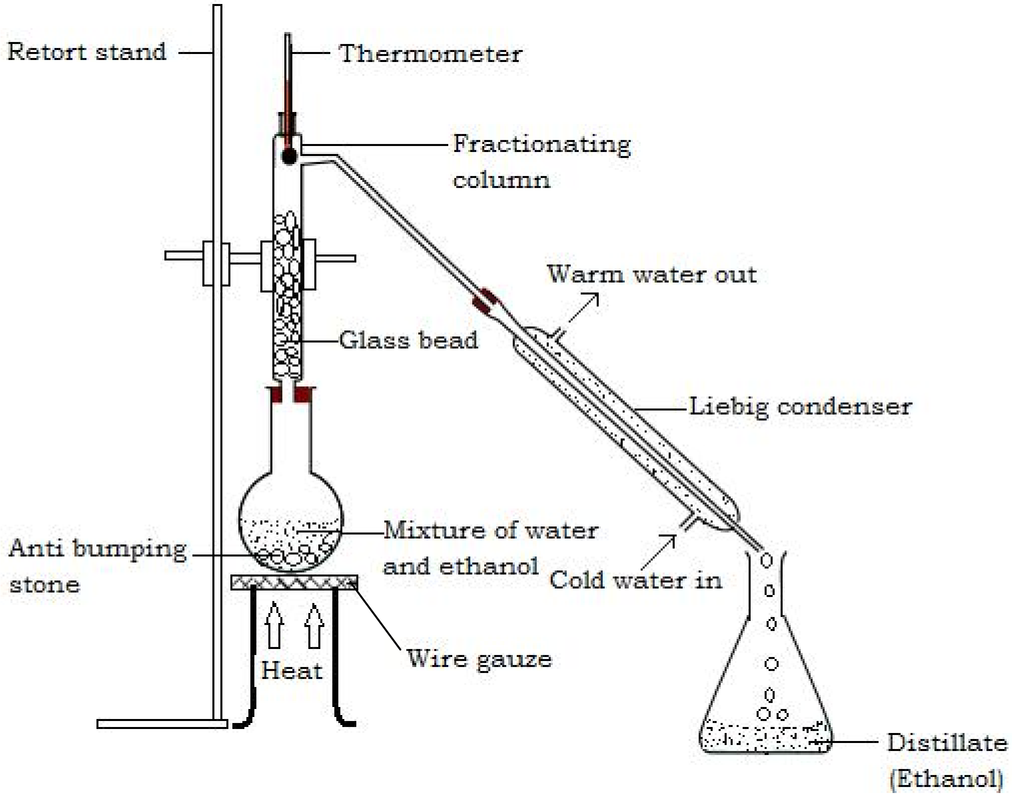
N.B. Do not continue to heat until all the liquid has been transferred from the flask to the test tube. Stop heating shortly when the temperature remains steady.

5. Fractional distillation

This is a process of separating two liquids by using their differences in boiling point. Liquids to be separated by this method must have different boiling points. This method can be used in separating ethanol and water; separating a mixture of gases; petroleum refining and brewing industries.

Separating a mixture of ethanol and water by fractional distillation

Set up of apparatus



Procedure

- Place the mixture of ethanol and and water in the distillation flask and arranged the apparatus asa shown above.
- Heat the mixture until when it begins to boil. Ethanol having a lower boiling point (78˚C) boils first leaving the water in the distillation flask. The temperature remains constant for a while as the ethanol vapor rises up the fractionating column and is condensed by the Liebig condenser forming a colorless distillate (ethanol)

N.B. 1. Water vapor may also try to escape but it cannot go up the fractionating column as it condenses and falls back into the distillation flask.

2. The beads in the fractionating column increase the surface area for condensation of the rising vapor.

- When the temperature rises slightly above the boiling point of ethanol (78˚C). Remove the distillate and this will be pure ethanol. The water remains in the distillation flask and may be collected as the second distillate if heating is continued up to 100˚C.

N.B. Anti bumping stones (broken porcelain) is added to the mixture of ethanol and water to prevent violent boiling of the mixture

Applications of fractional distillation

- It's used in the brewing industry to produce quality drinks such as beer, whisky and distilled waragi.
- It's used in the oil refining industry to separate the different components of crude oil like kerosene, diesel and petrol.
- It's used in large scale manufacture of oxygen. Liquid air can be fractionally distilled and its components separated.

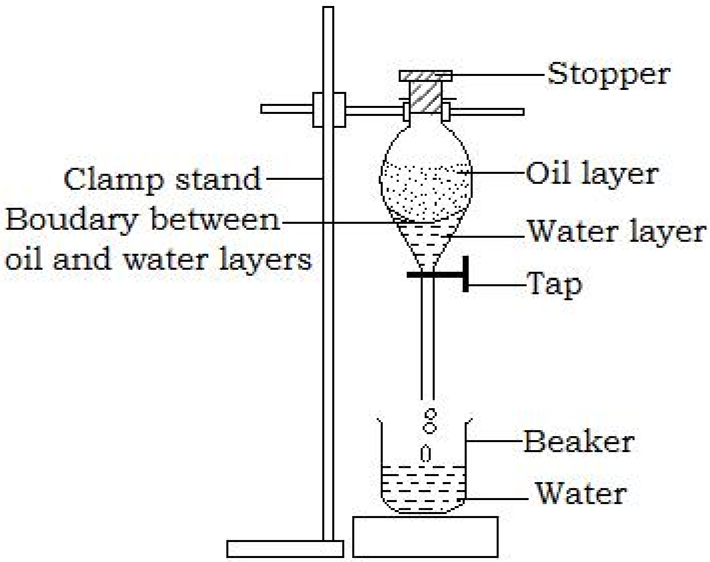
d) Separation of immiscible mixtures/liquids(Heterogeneous mixture)

Heterogeneous mixture is a non uniform mixture of two or substances. Each of the substances may form separate layers.
Immiscible liquids are liquids which do not mix in any proportion with each other. They therefore form separate layers e.g. a mixture of oil and water, paraffin and water. Such a mixture can be separated by use of a separating funnel.

6. Use of a separating funnel

A separating funnel is used to separate liquids that do not mix like oil and water.

Separation of oil and water using a separating funnel
Set up of apparatus



Procedure

- A mixture of oil and water is poured into the separating funnel when the tap is closed.
- The mixture is allowed to settle and oil being less dense than water will remain at the top and water at the bottom.
- Carefully open the tap and run the water in an empty beaker placed below the separating funnel. As soon as the water has run into the beaker, close the tap and the oil remains in the separating funnel.

e) Separation of volatile solids from non volatile solids

Solids that can sublime (volatile solids) can be separated from solids that cannot sublime by sublimation method.

7. Sublimation method

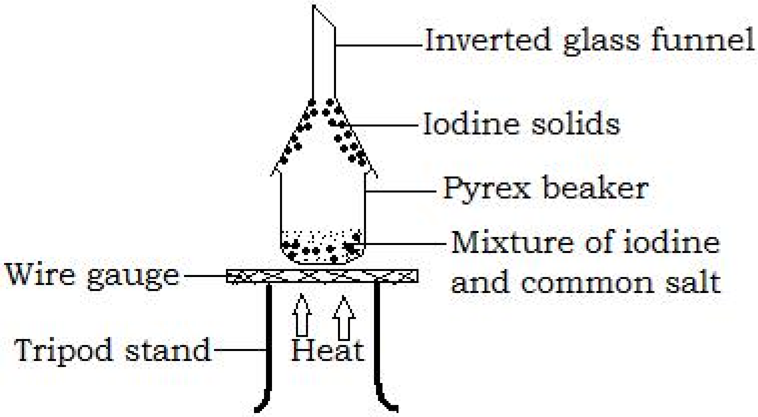
Sublimation is the direct change of state of a substance from solid to gas and from a gas to solid without going through the liquid state. Sublimation method can be used to separate:

- Volatile solids from their solution e.g. iodine solids from its solution in water.
- Volatile solids mixed with non volatile solids e.g. iodine from common salt.

Examples of substances that can sublime are: iodine, ammonium chloride, anhydrous iron(III) chloride, anhydrous aluminium chloride and benzoic acid.

Separation of iodine from common salt

Set up of apparatus



Procedure

- A mixture of equal amounts of iodine and common salt (sodium chloride) is placed in a beaker with a glass filter funnel inverted over its top as shown above.
- The beaker is then heated gently.

Observation

Iodine solids sublime forming a purple vapor that settles at the inside of the filter funnel forming grey solids of iodine. White common salt is left in the Pyrex beaker.

f) Separation of magnetic substances from non magnetic substances

Magnetic substances are attracted to a magnet while non magnetic substances are not, there fore they can be separated by magnetic method.

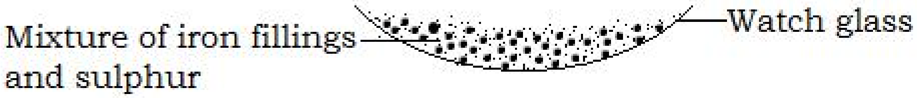
8. Magnetic separation

This method can be used to separate feromagnetic substances like iron fillings from non magnetic substances like sulphur and glass.

Separation of iron fillings from sulphur

Procedure

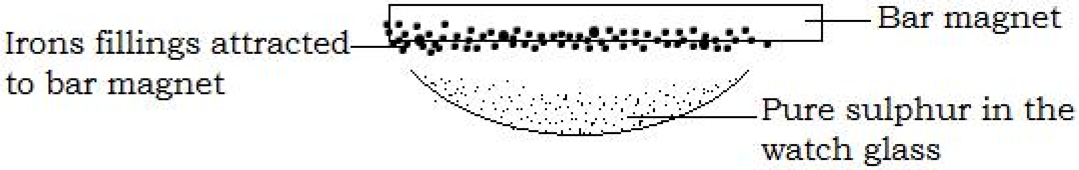
- Place a mixture of iron fillings and sulphur on a watch glass. As shown below.



- Pass a bar of magnet over the mixture.

Observation

- The iron fillings were attracted to the bar magnet and clung to it while the sulphur was left in the watch glass because they are non magnetic.



g) Separation of different solutes e.g. salts dissolved in a solvent

Two of more solutes (salts) can dissolve in a particular solvent and can be separated from each other on the basis of their differences in solubility in the same solvent. E.g. a mixture of potassium chlorate and potassium chloride can be separated from solution. The method of separation used is **fractional crystallization**. Another method that can be used to separate a mixture of two solutes in a solution is **solvent extraction**.

9. Fractional crystallization

This is a method of separating two or more soluble solutes by crystallization using their differences in solubility in the same solvent.

Crystallization is the process of forming crystals by evaporating a saturated solution. I.e. the solution is made more and more saturated by evaporating it so that the excess solutes are deposited as crystals.

In fractional crystallization, a mixture of two solutes (e.g. salts) with different solubilities is dissolved in a minimum amount of hot solvent to make a saturated solution. The solution is then allowed to cool and as it cools, the less soluble salt crystallizes out (forms crystals). It is then filtered and the more soluble salt remains in the filtrate. The crystals are again dissolved in a minimum amount of hot solvent and cooled again to form crystals. The process is repeated until pure crystals of less soluble solutes are separated from the mixture.

In other words, as the hot solution cools, the solutes with different solubilities form crystals at different temperatures.

10. Solvent extraction

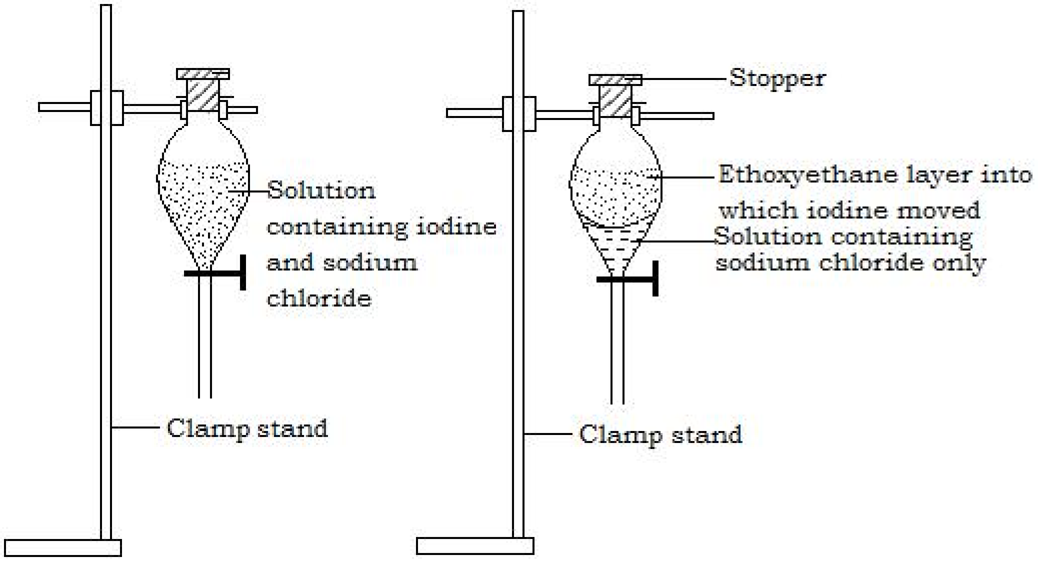
This method can be solutes from a solution in which more than one solute is dissolved. E.g. a solution containing iodine and sodium chloride dissolved in water. Solvent extraction works on two principles:

- One solute in the solution must be more soluble in the extracting solvent.
- The extracting solvent must not be miscible with the solvent in which the solutes are dissolved.

Separation of iodine from a solution containing iodine and sodium chloride using ethoxyethane as the extracting solvent

Procedure

- A mixture of iodine and sodium chloride solution is poured into a separating funnel.
- Ethoxyethane solvent is added on to the mixture and the stopper placed on top of the funnel. The funnel is then shaken and the mixture left to settle.



Results

The ethoxyethane layer is less dense and settles on top. Iodine being more soluble in the ethoxyethane solvent than in water passes into the ethoxyethane leaving sodium chloride in the water solvent.

h) Separation of colored substances (dyes)

A mixture of colored substance is separated by chromatography.

11. Chromatography

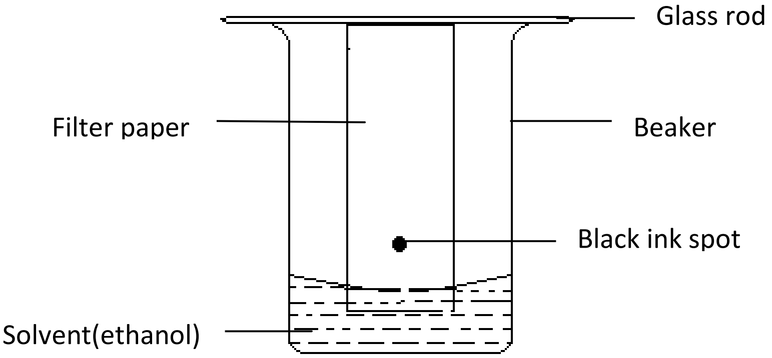
This is the process of separating a mixture of colored substances with different rates of movement in a solvent over an adsorbent medium. The adsorbent medium should be stationary and porous. This method can be used in separating different components in ink, pigments in leaves or grasses. There are different types of chromatography but the most common one is paper chromatography in which the stationary medium is a porous paper e.g. filters paper.

The dyes (colored substances) separated by this method should be **soluble** in the solvent and should be **absorbed** to some extent by the stationary medium (adsorbent medium). The pigment (dye) that is more soluble in the solvent is less absorbed (has les tendency to stick to the paper) and there fore moves faster. The dye that is less soluble has a great tendency of being absorbed and travels a very short distance.

Separation of the components in black ink

Procedure

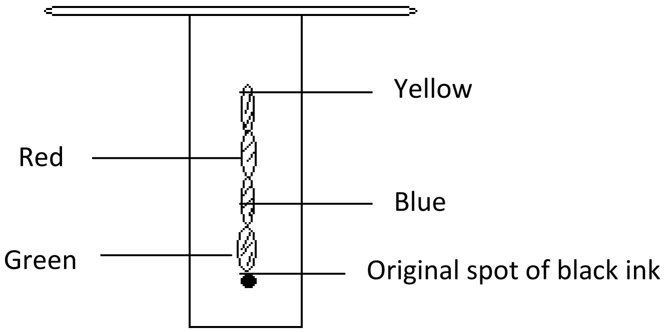
- Mark a spot on a strip of filter paper using black ink.
- Suspend the filter paper with its tip in a beaker containing ethanol as shown below.



- After some times, the filter paper was removed from the solvent (ethanol) and allowed to dry.

Results

The chromatogram of the different components of black ink is as shown below.



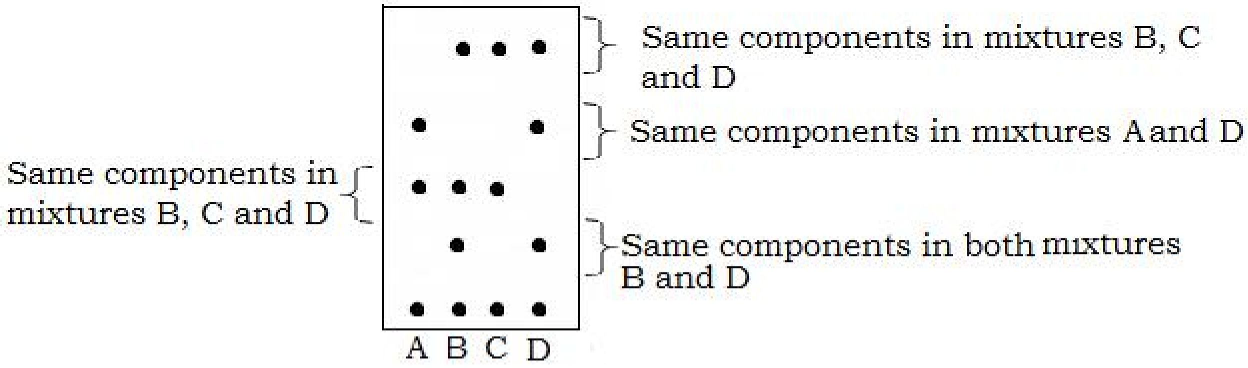
The chromatogram is the paper that shows the different components (colors) of the mixture separated from each other.

The solvent (ethanol) moves up the filter paper and dissolves the colored components in the black ink, then moves with them. The colored components moved upwards at different rates depending on their solubility in the ethanol. The most soluble component (yellow) moved the greatest distance as it was less absorbed by the filter paper and vice versa.

N.B. If two of more mixtures are separated using the same chromatographic paper, the components which are common in both mixtures travel the same distance.

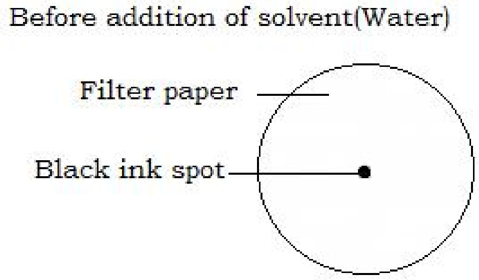
Example to illustrate

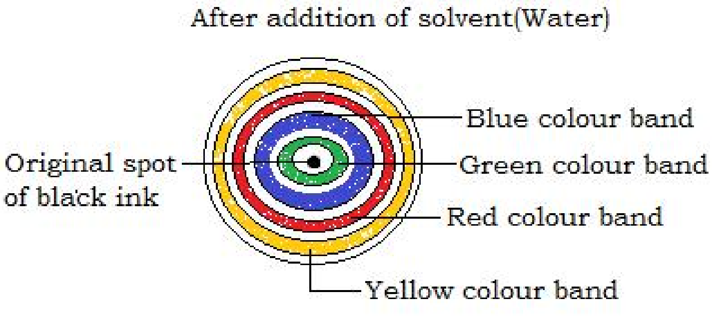
Consider the chromatogram of mixtures A,B,C and D(which are all different mixtures) shown below.



Alternative procedure for separating components in black ink

- Place a filter paper on a flat surface and a drop of black ink to the centre of the filter paper; allow the ink on the paper to dry.
- Add a drop of water (solvent) to the dried ink spot and allow it to dry before you add another drop of water i.e. for every drop of water, allow it to dry before you add another drop.
- Repeat the addition of water until the band of colored substances almost reaches the edge of the filter paper.





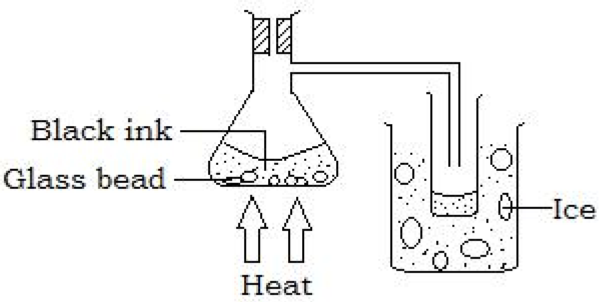
Observation

As the solvent (water) moved across the filter paper, it carried a long with it the different dyes in black in and separated them. The dyes separated because they have different solubilites in water and are absorbed to different degrees by the porous filter paper. Yellow was the most soluble in water and least absorbed by the filter paper there fore travelled the longest distance. Green was least soluble in water and most absorbed by the filter paper there fore travelled the shortest distance.

The ratio of distance moved by the solute (dye) to the distance moved by the solvent in chromatography is called, the **retention factor** (R_f).

Exercise

1. Sharron wants to separate the liquid from the dye in black ink. She uses this apparatus to distill the liquid.



Copy the diagram and add to it a thermometer in a correct position to measure the boiling point of the liquid which distilled over first.

a) Why was ice put in the beaker around the test tube.

b) The black dye in ink was thought to have a boiling point of either 54˚C, 93˚C or 150˚C

Sharron collected the first liquid in the tes tube and found that its boiling point was 100˚C. The liquid was colorless. State with a reason which of the three temperatures is the boiling point of the dye in black ink.

c) Sharron tested another black ink to see if more than one color was present in it. She made chromatograms by placing drops of the liquid to be tested on a pencil line drawn on a filter paper. She stood the paper vertically in a suitable solvent for some hours.

Figure 1 shows the chromatogram of some known dryes.

Figure 2 shows the chromatogram of the black ink.

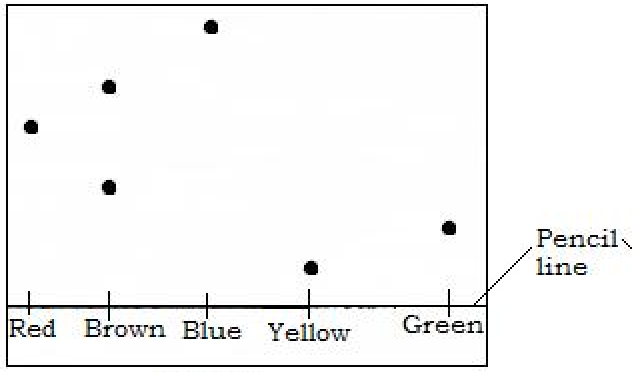


Figure 1

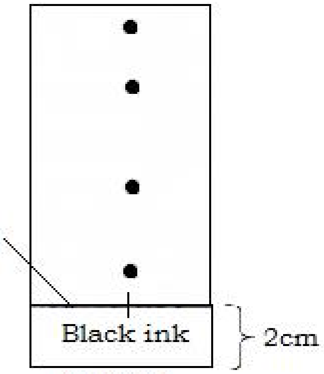


Figure 2

i) Which dyes are present in the black ink

ii) Suggest why the starting line is drawn in pencil

Alloys

An alloy is a mixture of two or more metals. Alloys are formed by thoroughly mixing molten metals. It has been found that alloying produces metallic substance with more useful properties than the original pure metal it is made from. Examples of alloys include:

Alloy	**Composition**	**Use**
Brass	Copper and zinc	Machine bearings, jewellery, electrical objects, metallic parts of door furniture, water pipes.
Bronze	Copper and tin	Machine parts, ornaments e.g. necklaces, castin.
Solder	Lead and tin	Connecting electrical wiring

Duralumin	Aluminium, copper and magnesium	Aircraft construction, bicycle parts and small boats.
Magnalium	Alluminium and magnesium	Aircraft construction, small boats
Alnico (ferromagnetic)	Aluminium, nickel and cobalt	Generators, electric motors, mass spectrometer
Pewter	Lead, tin and small amount of antimony	Plates, ornaments and drinking mugs.
Stainless steel	Iron, chromium and nickel	Cutlery, kitchen sinks, surgical instruments
Hard steel	Iron and carbon	Cutting tools, razor blades, chisels
Tungsten steel	Iron and tungsten	Edges of high speed cutting tools

Differences between mixtures and compounds

Mixture	**Compound**
It contains two or more substances	Is a single substance
Components can be separated by physical means	Components cannot be separated by physical means
The composition varies	Composition is always fixed
No chemical change takes place when mixtures are formed	Chemical changes take place when compounds are formed
Properties of a mixture are an average properties of the components	Properties of a compound are different from those of its component elements

Simple Criteria for purity

A pure substance is one which has distinct physical and chemical properties that are only unique it. These physical and chemical properties can be reproduced at ant time under the same conditions.

Physical properties such as taste, smell, and color of a substance cannot give accurate measurements of purity. For instance, sea water looks like pure water, impure naphthalene has the same smell as pure naphthalene (smell of moth balls), pure copper wire feels like a wire made of copper alloy.

However, the following properties can be used to determine the degree of purity of substances. The values are constant for a pure substance.

- Melting points of solids (is constant and the solid melts sharply)
- Density of solids and liquids
- Boiling point of liquids
- Freezing point of liquids
- Refractive index for liquids

The experimental values of the above properties are compared to the standard values and if they coincide, then the substance being investigated is pure.

N.B. The experiment for the investigation of purity must be carried out under the same conditions (of pressure and temperature) in which the standard values were obtained.

SOLUTIONS AND SUSPENSION

SOLUTIONS

A solution is a uniform mixture of two or more substances. Examples of solutions include: Air-a solution of gases; aqueous solution-a solution of any substance in water; alloy-a solution of metals.

Solutions are formed when solutes completely mix with solvents.
I.e. Solute + Solvent = Solution

A **solute** is a substance that dissolves in a solvent. Examples are salts and sugar.
A **solvent** is a substance that dissolves solutes. E.g. water.
A solute is said to be **soluble** in a given solvent if it can dissolve in the solvent and **insoluble** if it does not dissolve in that given solvent.

Depending on the amount of solute in the solvent, solutions can be classified as unsaturated, saturated and super saturated.

Unsaturated solution: This is a solution that can dissolve more solutes at a particular temperature.

Saturated solution: This is a solution that cannot dissolve any more solute at that temperature in the presence of undissolved solutes.

Super saturated solution: This is a solution that contains more solutes than it can hold at that temperature in the presence of undissolved solutes.

SUSPENSION

A suspension is a liquid containing small particles of solids which are spread throughout it and the solid particles settle on standing. Examples of suspensions include:

Paint – a suspension of colored substances in water or oil: Muddy water- a suspension of mud in water: suspension of chalk dust particles in water.

N.B. A suspension of a liquid in another liquid is called an emulsion and not a suspension.

Characteristics of a suspension

1. Tinny solid particles are visibly seen spread throughout the liquid.
2. On standing, the tinny solid particles settle at the bottom leaving a clear liquid. Centrifuging makes the solid particles to even settle faster.
3. The tinny solid particles in a suspension can be separated from the liquid by filtration.

CRYSTALS

A crystal is a solid that consists of particles (atoms, molecules or ions) arranged in an orderly and repeatitive manner resulting into a definite shape. Crystals have regular shapes with flat sides and sharp edges. Examples of crystals include: sugar crystals, blue copper(II)sulphate crystals, common salt (sodium chloride) crystals, potassium nitrate crystals and potassium aluminium nitrate(Alum). Crystals are formed by the process of crystallization.

Crystallization is the process of evaporating a solution making it more saturated with the solutes such that the excess solutes are deposited as crystals. The solution is evaporated by either heating it or exposing it to sunlight.

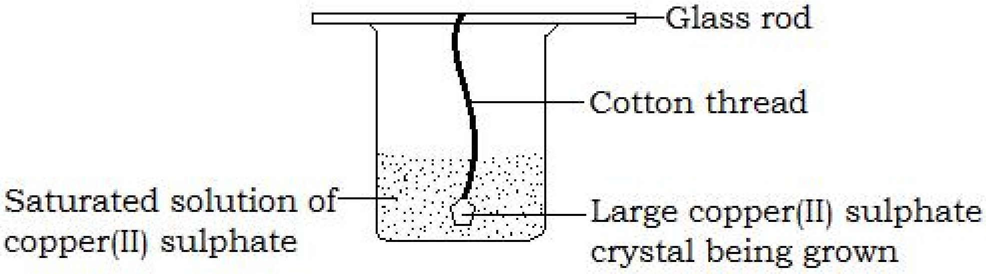
Growing of a large crystal

Large crystals can be grown from saturated solutions.

Growing of a large crystal of copper(II)sulphate.

Procedure

Drawing



- Pour a saturated solution of copper(II)sulphate in a beaker
- Dip a cotton thread into the solution and remove it. Allow the thread to dry, as small crystals form on it.
- Put back the thread containing small crystals into the saturated solution of copper(II)sulphate and allow the beaker to stand for some days in a warm place.

Observation

Large blue crystals of copper (II) sulphate will form on the cotton thread.

Water of crystallization

This is the definite amount of water with which some substances chemically combine when they form crystals (from their solutions in water). A compound that contains water of crystallization is referred to as a **hydrated compound (hydrate).** Examples of hydrated compounds include:

Name	**Formula**
Sodium carbonate decahydrate (sodium carbonate 10-water)	$Na_2CO_3.10H_2O$
Calcium chloride hexahydrate (Calcium chloride 6-water)	$CaCl_2.6H_2O$
Copper (II) sulphate pentahydrate (Copper (II) sulphate 5-water)	$CuSO_4.5H_2O$
Magnesium sulphate heptahydrate (Magnesium sulphate 7-water)	$MgSO_4.7H_2O$

Some crystals do not have water of crystallization e.g. sodium chloride (NaCl) and lead(II)nitrate ($Pb(NO_3)_2$).

Dry compounds that do not contain water of crystallization are called **anhydrous compounds.** E.g. anhydrous copper(II)sulphate (white powder); anhydrous calcium chloride (white powder). They normally exist in powdery forms and not as crystals.

Exchange of water takes place between some crystals and the atmosphere. Some substances absorb water from the atmosphere while some release water to the atmosphere. These substances are grouped in to three as:

1. Hygroscopic substance

This is a substance that absorbs water from the atmosphere and remains physically unchanged. I.e. does not dissolve to form solution. Examples include: anhydrous copper(II)sulphate ($CuSO_4$); calcium oxide (CaO); Copper(II)oxide (CuO);anhydrous calcium chloride ($CaCl_2$) and concentrated sulphuric acid(H_2SO_4). Hygroscopic substances are used for drying gases.

Hygroscopy is the process of absorbing moisture/water from the atmosphere without the substance changing physically.

2. Deliquescent substance

This is a substance that absorbs water from the atmosphere and dissolves in it to form a solution. E.g. sodium hydroxide pellets(NaOH); Potassium hydroxide pellets (KOH); and Copper(II)chloride ($CuCl_2$).

Deliquescence is the process whereby a substance absorbs water from the atmosphere and forms a solution.

3. Efflorescent substance

Is a substance that loses its water of crystallization to the atmosphere. The process in whereby a substance loses its water of crystallization to the atmosphere is called **efflorescence.** Examples of wfflorescent substances include; sodium carbonate decahydrate ($Na_2CO_3.10H_2O$); Sodium suphate decahydrate ($Na_2SO_4.10H_2O$).

$$Na_2CO_3.10H_2O(s) \longrightarrow Na_2CO_3.H_2O(s) + 9H_2O(l)$$
$$Na_2SO_4.10H_2O(s) \longrightarrow Na_2SO_4(s) + 10H_2O(l)$$

When hydrated compounds are heated, they lose their crystalline shapes as the water of crystallization escapes, become powdery and as well their colours change.

For example when hydrated copper(II)crystals are heated, the blue crystals turn into white powder (anhydrous copper(II)sulphate) and a colorless liquid (water) condenses on the cooler parts of the test tube.

$$\underset{\text{Blue crystals}}{CuSO_4.5H_2O(s)} \longrightarrow \underset{\text{white powder}}{CuSO_4(s)} + \underset{\text{colorless liquid}}{5H_2O(l)}$$

On strong heating, the white powder turns to black producing white fumes of a gas with a choking smell.

$$\underset{\text{White powder}}{CuSO_4(s)} \longrightarrow \underset{\text{Black powder}}{CuO(s)} + \underset{\text{White fumes with a choking smell}}{SO_3(s)}$$

Sample questions on Laboratory apparatus; Matter; Elements, Mixtures and compounds

Bunsen burner and flames

1. *A Bunsen burner is one of the apparatus used for heating in the laboratory.*

a) *Define the term laboratory*

b) *Make a labeled drawing of a Bunsen burner and give the functions of all labeled parts.*

c) *How are you able to light a Bunsen burner?*

2. *Define a flame? Mention the two common types of flames and conditions under which each is produced. Make labeled drawings of the two types of flames and explain what happens in each of the labeled zones. Give four differences and two similarities between the two types of flames. What is strike back?*

Matter

1. *What is matter? Mention the different states of matter and give the characteristics of each state.*
2. *State kinetic theory of matter? Use kinetic theory of matter to explain why a solid object changes into a liquid and finally a gas when heated. Uses a schematic diagram to illustrate the different processes involved in the transition of matter from one state to another. Define the processes mentioned in the schematic diagram.*
3. *Define Brownian motion and diffusion. Describe experiments to demonstrate Brownian motion and diffusion in liquids and gases. Outline four factors that affect the rate of diffusion of gases. Describe an experiment to show that ammonia and hydrogen chloride gases diffuse at different rates.*
4. *Define the terms melting, freezing and boiling points. Describe experiments to show how you can determine the melting point of naphthalene and boiling point of ethanol. Sketch a temperature-time graph for cooling liquid naphthalene. Identify the melting point and label it T on the graph axis.*
5. *Define the terms physical and chemical changes and state at least two examples of each. Mention the characteristics of physical and chemical changes and clearly give the differences between them.*
6. *Describe the observations you would make when each of the following substances are heated: candle wax; magnesium, iodine; potassium permanganate; lead(II) nitrate; hydrated copper(II) sulphate; sulphur and zinc oxide.*

Elements, mixtures and compounds

1. *Define the terms element, mixture and compounds giving three examples of each (where applicable write the symbol or formula). Give the names and formulae of five compounds stating clearly the elements that they are composed of. What are the differences between a mixture and a compound?*
2. *On what property is each of the following methods of separation of mixtures based: fractional distillation; fractional crystallization; sublimation; chromatography and magnetic separation?*
3. *Describe how the following mixtures can be separated: oil and water; salt from water; sugar from sand; sand and water; ethanol and water; diesel and kerosene; a mixture of different colours; iodine and common salt; potassium chloride and potassium nitrate; iron and sulphur.*
4. *What an alloy, mention three examples is of alloys any their uses. Brass, bronze, duralumin, solder and stainless steel are alloys. Mention the components in each alloy.*

Solutions, suspension and crystals

1. *Define the following terms as applied to chemistry: solution, suspension, solute, solvent, saturated solution, unsaturated solution, super saturated solution and crystallization.*
2. *What is a crystal? Explain with a the aid of a drawing how a large crystal of copper(II) sulphate can be grown.*
3. *Some compounds are said to have water of crystallization. What is water of crystallization? Mention three examples of hydrated compounds and write their formulae.*
4. *Define the following terms: hygroscopy, deliquescence and efflorescence. What are hygroscopic and efflorescent substances (illustrate with at least two examples of each)*

THE ATMOSPHERE AND COMBUSTION

THE ATMOSPHERE

The atmosphere is a layer of air surrounding the earth. Air is a mixture because; it has variable composition; its constituents or components can be separated by physical means and its formation is not accompanied by heat or light (energy changes).

The composition of the atmospheric air

The atmospheric air consists of

1. A mixture of gases mainly oxygen, nitrogen, carbon dioxide and noble gases or inert of rare gases.(the noble gases include; helium, neon, argon, krypton and xenon). It also consists of minute quantities of other gases such as sulphurdioxide, hydrogen dioxide, carbon monoxide, traces of methane and ozone.
2. Water vapor
3. Solid particles such as pollen grains, soot, smoke, dust particles etc.
4. Micro organisms such as bacteria.

The percentage composition of air by volume

Nitrogen	78%
Oxygen	21%
Carbondioxide	0.03%
Noble gas	0.9%
Water vapor	0.4% (variable)
Impurities (e.g.hydrogensulphide, sulphur dioxide)	Variable

How to detect the presence of atmospheric gases and water vapor

1. **Carbondioxide**

When carbondioxide is passed through lime water, the lime water turns milky.

2. **Water vapor**

The presence of water vapor in the atmosphere can be detected using anhydrous copper (II) sulphate or cobalt chloride paper. When water vapor gets in to contact with anhydrous copper(II) sulhate, it turns from white to blue or when water gets into contact with blue cobalt chloride paper, it turn to pink.

3. **Oxygen**

Since oxygen supports burning, its presence in the atmosphere can be detected using a glowing splint. When a glowing splint is in contact with oxygen, it relights of rekindles. As well, burning e.g. of a candle takes place in the open due to the presence of oxygen in the atmosphere.

4. **Nitrogen**

Nitrogen is inert and very difficult to detect. However, when oxygen and carbondioxide have been removed from the atmosphere, the remaining gases constitute mainly nitrogen.

Uses of the atmosphere

The atmosphere supports life in many ways:

1. The atmosphere filters off the harmful ultra violet radiations from the sun, thus protecting life.
2. It provides oxygen needed for respiration and combustion (burning).
3. Nitrogen is inert and therefore slows down various chemical reactions such as burning and rusting. Nitrogen is also needed by plants.
4. Carbondioxide is needed for photosynthesis by green plants.
5. Water vapor is necessary for rain formation. It also provides moisture to sooth our bodies and make plants cool.

6. Noble gases provide inert conditions for various chemical reacrtions. Other uses of noble gases include:
   - i) Helium is used to fill balloons because its light.
   - ii) Neon is used in electric sign boards for advertisement.
   - iii) Argon is used in electric bulbs.

Air pollution

Air pollution is the introduction of substances in the atmosphere that directly or indirectly endanger the lives of living organisms. Air pollutants (substances introduced in the atmosphere that endager life of living organisms) include:

i) Carbondioxide which is responsible for green house effect.
ii) Carbonmonoxide which is very poisonous reacting with blood and forming a stable compound (carboxyheamoglobin) and stopping it from carrying oxygen.
iii) Oxides of sulphur e.g. Suphur dioxide and sulphur trioxide which are toxic.
iv) Oxides of nitrogen e.g. nitrogen monoxide, nitrogen dioxide and dinitrogen oxide.
v) Compounds of heavy metals like leadf, mercury and cadmium are toxic and poisonous.
vi) Radioactive particles which cause abnormalities to plant and animals as a result of mutation.

Sources of air pollutants

i) Burning of fossils (coal, coke, oil fuel) produce oxides of carbon and sulphur, hydrogen sulphide and unburnt oil vapor.
ii) Industrial and domestic depositions. Thgese produce oxides of carbon, sulphur. They may produce methane and oher heavy metals like mercury.
iii) The exhaust fumes of locomotives e.g cars contain carbon monoxide, carbondioxide, nitron dioxide, unburnt fuel and some heavy metals like lead.
iv) Nuclear explosions produce radioactive particles.
v) Tobacco smoking
vi) Natural processed like volcanic eruption.
vii) Gases released from manure in farm land.

Effects of atmospheric pollution

1. Causes respiratory diseases e.g. cough, lung cancer e.t.c.
2. Formation of fog which distracts vision and may also be poisonous.
3. Formation of acid rain due to the presence of sulphur dioxide, sulphur trioxide, carbon dioxide e.t.c.
4. Darkens buildings especially those painted using lead paints.
5. Causes global warming, particularly carbondioxide and other green house gases like methane.

COMBUSTION

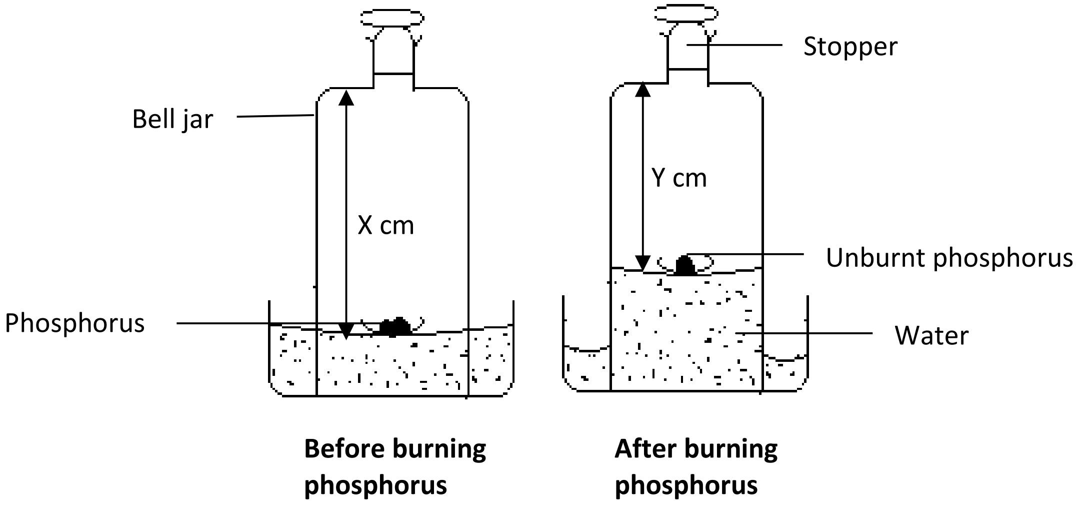
This is a chemical reaction in which a substance combines with air to form energy. Combustion can be slow (with no flame) or rapid (rapid combustion produces flame). Examples of rapid combustion are: burning bush and burning charcoal while examples of slow combustion include: digestion of food, rusting of iron, and fermentation/decay.

Experiments on combustioon

Experiment to determine the percentage of air used in combustion

i) Using phosphorus

Drawing of setup



Procedure

Place a small piece of phosphorus in a crucible, float the crucible in a water trough and place a bell jar over it.

Measure and record the height of air (X cm) in the bell jar.

Ignite the phosphorus by touching it with the end of a hot wire and quickly replace the stopper of the jar.

Observation

The phosphorus burns with a bright yellow flame producing dense white fumes. The water level drops for a while due to the increased pressure as a result of expansion of air, then rises as the white fumes dissolve in the water.

When phosphorus stops burning and the water level comes to rest, measure the height of air left in the bell jar (Ycm).

Results

Height of air in the bell jar before burning phosphorus =X cm
Height of air in the bell jar after burning phosphorus =Y cm
Air used in the burning of phosphorus = $(X-Y)$ cm
Percentage of air used in the burning of phosphorus =$\left(\frac{X-Y}{X}\right)$x 100

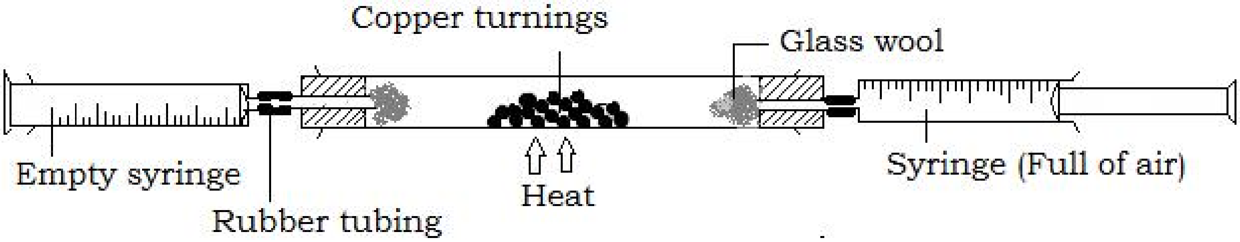
Conclusion

The percentage volume of air used in the burning of phosphorus was 20.8% indicating that the gas used was oxygen. The remaining gas in the gas jar was mainly nitrogen.

(Use this data and calculate the percentage of air used up in the combustion of phosphorus: Y= 55.8cm, X= 70.5cm)

ii) Using copper turnings

Drawing of set up



Procedure

Place some copper turnings (small pieces of copper) in a combustion tube and connect the combustion tube to two syringes (each 100cm^3 in volume) as shown above. One of the syringes must be full of air and the other empty.

Record the volume of air, Y cm^3 in the syringe.

Heat the copper turnings strongly while passing air over it from one syringe to another until there is no further change in the volume o air in the syringes.

N.B. The glass wool allows air to easily be blown from one syringe to another without the copper turnings entering the syringes.

Allow the apparatus to cool and record the remaining volume of air in the syringe, Y cm^3.

Observation

The copper turnings turn from brown to black as the copper reacts with ooxygen to form copper(II) oxide which is black. The volume of air in the combustion tube and the syringes reduced as part of the air is used up during the process of combustion.

Results

Volume of air in the syringe before heating = Y cm^3

Volume of air in the syringe after heating = X cm^3

Volume of air used during combustion = (Y-X) cm^3

Percentage of air used up during combustion = $\left(\frac{Y-X}{Y}\right) \times 100$

Conclusion

The percentage volume of air used up during the combustion was 20.8%, indicating that the gas was oxygen since it occupies approximately 21% of the air. Oxygen is there fore the active component of air used during combustion.

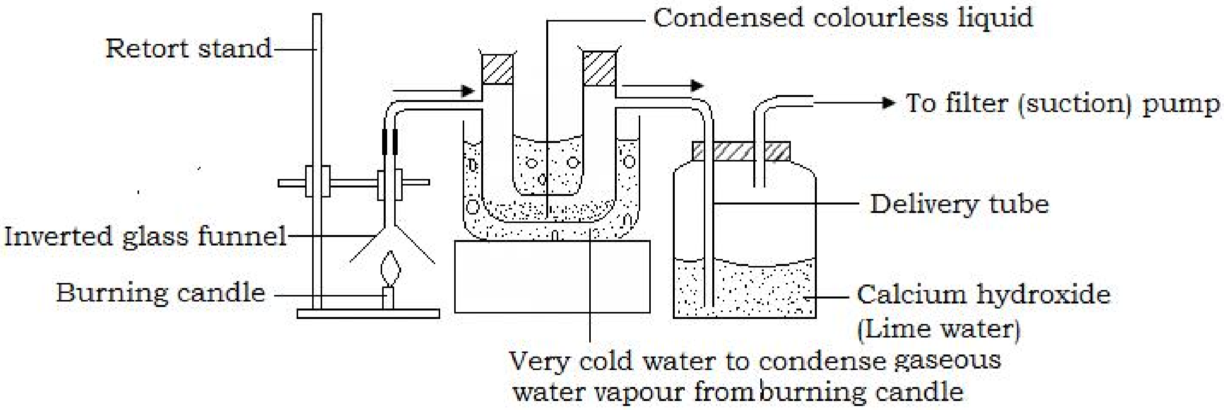
(Use this data and calculate the percentage of air used up in the combustion of copper: Y= 100cm^3, X= 80cm^3)

Experiment to determine products formed when hydrocarbons burn in air

A hydrocarbon is a compound containing hydrogen and carbon only. Examples of hydrocarbons include: candle, natural gas like methane (CH_4),petrol and paraffin.

Products of combustion of a candle

Set up of apparatus



Procedure

The arrangement of the apparatus is made as above and the candle lit.

As the candle burns, the pump is switched on to suck the gaseous products of the burning candle through the glasas funnel into the setup.

Observation

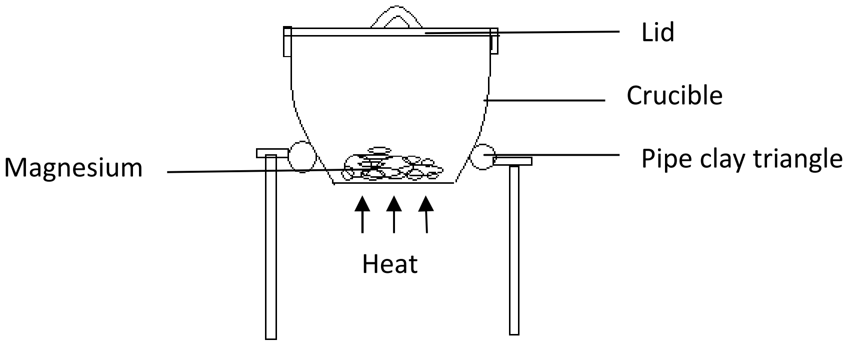
A colorless liquid is collected in the U tube that turns white anhydrous copper (II) sulphate to blue, indicating that it is water. Lime water in the test tube turned milky indicating the presence of carbondioxide.

Conclusion

When a candle burns water and carbondioxide are produced as the only products showing that the gas used up in the combustion is oxygen.

Experiment to find out if the mass of substances change when they burn in air

Setup of apparatus for the combustion of magnesium



Procedure

A known mass of polished magnesium is placed in a crucible and weighed with the lid of the crucible on. The apparatus is then set as above.

The crucible is first heated gently and then strongly. Occasionally the lid of the crucibl;e is opened to allow more air in for the complete burning of magnesium.

When all the magnesium has burnt, the crucible is allowed to cool and then weighed with its lid on.

Results

Initial weight = a grams
Final weight = b grams
It is observed that b is greater than a. The mass of magnesium increased by (b-a) grams.

Conclusion

When a substance burns in air, its mass increases due to the addition of oxygen to it.

RUSTING

This is the process by which iron combines with oxygen in the presence of water to form brown solids of hydrated iron (III) oxide called rust.

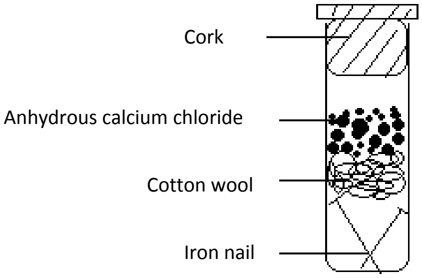
Rust is a brown solid of hydrated iron (III) oxide formed when iron combines with oxygen in the presence of water. The chemical name of rust is hydrated iron (III) oxide and its chemical formula is $Fe_2O_3.xH_2O$.

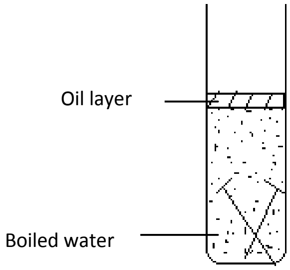
Conditions necessary for rusting

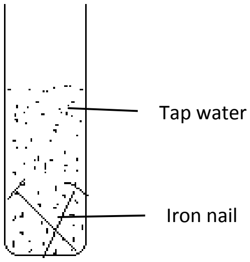
The conditions necessary for rusting are air/oxygen and moisture/water.

Experiment to investigate the conditions necessary for rusting

Set up of apparatus







Procedure

Place some iron nails in three test tubes and arrange them as shown above.

In test tube A, plug cotton wool and add anhydrous calcium chloride on top of the wool then close the tube with a cork. The calcium chloride absorbs all the moisture and keeps the air completely dry.

In test tube B, add freshly boiled distilled water to cover the nails completely. then pour a layer of oil on top of the water. The water is boiled to dive off any dissolved oxygen/ air in the water. The layer of oil is added to prevent air/ oxygen from entering the water.

In test tube C, tap water is added to the nails.

Leave the set up to stay for 5 days.

Observations

The nails in test tube A did not rust as there was no moisture/ water.

The nails in test tube B did not rust as there was no air/ oxygen.

The nails in tes tube C rusted as there was both air and water.

Conclusion

Both air and water are necessary for rusting.

Effects of dissolved substances on rusting

Presence of dissolved salts or acids in water tend to accelerate rusting while alkaline solutions tend to slow down rusting, therefore iron will rust faster in a salty or acidic solution than in tap water.

Advantages of rusting

- It helps in the decomposition of disposed iron and increases the iron content in the soil necessary for plants.
- It creates and increases market for the iron manufacturing industry.
- Rust is used for making paints
- Study purposes in the laboratory.

Disadvantages

- It weakens and wears the metal.
- It causes the metal to lose shape.
- It makes metals to appear dark.
- It increases friction in machines.
- It increases the maintenance costs of materials made of iron.
- It contaminates tinned food.

Prevention of rusting

Rusting can be controlled by preventing the iron from getting in to contact with air/ oxygen and water/ moisture. The following methods can be used.

1. Oil and grease. When a layer of oil or grease is applied to the surface of iron, air and moisture are kept away and no rusting occurs. Oil and grease is applied to moving parts of machines.
2. Paint or tar. Paint and tar keeps moisture and air away from contact with iron. Paint is used on window frames, doors, bicycles e.t.c. while tar is used on iron used in the construction of bridges and ships.
3. Enamelling. Iron coated with enamel does not rust as moisture and air are kept away from the iron.
4. Coating with other metals. The coating of iron with other metals is referred to as **metal plating**. The iron can be coated with zinc, tin, chromium, silver and nickel to prevent rusting from occuring.

a) Coating with zinc. The process of coating iron with zinc is called **galvanization** and iron coated with zinc is called **galvanized iron.** The galvanized iron is used for making iron sheets, bath buckets and pails. Galvanized iron does not easily rust even when the zinc coat is scratched because zinc being more reactive than iron will react with air forming solids that seal off the scratch.

b) Coating with tin. This is called **tin plating.** Iron coated with tin is used to make tins that carry packed food, paints oil, e.t.c. Tin plated iron can rust when the tin coating is broken because iron is more reactive than tin and there fore will react with the air and moisture forming rust.

c) Coating with chromium. Chromium plated iron shines just like silver and can be used to make bicycle handle, car bumpers e.t.c.

d) Iron coated with nickel prevents rusting from occurring.

5. Alloying. In stainless steel, iron is mixed with carbon, chromium, nickel to prevent it from rusting.
6. Sacrificial protection. The iron being protected from rusting is connected via a wire to a block of a more reactive metal (magnesium and zinc are commonly used). The magnesium/ zinc being more reactive than the iron will react in preference to iron thus sacrificing to protect the iron. This method sometimes is referred to as cathodic protection. i.e. the iron being protected acts as the cathode and the more reactive metal acts as the anode.

Corrosion

Corrosion is the reaction between metals and the atmosphere which causes metals to lose etrheir structural properties like shape, color, strength e.t.c. Examples of corrosion include: rusting of irons, reaction of copper metal with the atmosphere (oxygen and carbondioxide) to form green solid of copper carbonate; reaction of aluminium with atmospheric oxygen to form aluminium oxide.

Self protection from corrosion

Some metals are self protecting from corrosion e.g. magnesium and aluminium react with the air to form oxides which form a thin layer on the surface of the metal. These oxide coatings prevent air from reaching the metal beneath it hence the metal is protected from further corrosion. Unfortunately, iron cannot protect its self in this way because the rust formed is porous and can allow air and water to reach the metal below it causing further corrosion.

OXYGEN

Occurrence

Oxygen occurs free in air and makes up about 21% by volume of the atmosphere. It also occurs in combined states as oxides, sulphates, carbonates, nitrates etc, and widely distributed in rocks, clay and mineral compounds. Oxygen is also common in carbohydrates, fats, proteins and water. The amount of oxygen in the atmosphere can be increased by processes such as photosynthesis and can be reduced by processes like combustion, respiration, rusting, germination etc.

Laboratory preparation of oxygen

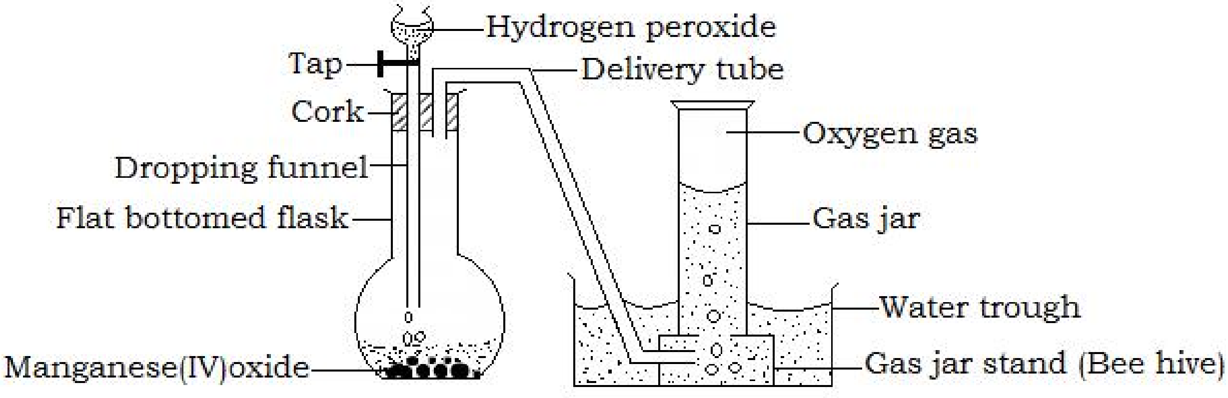
Oxygen can be prepared in the laboratory from oxygen rich compounds such as hydrogen peroxide (H_2O_2), potassium chlorate ($KClO_3$), potassium manganate (VII) ($KMnO_4$), and sodium peroxide (Na_2O_2). On a large scale oxygen is prepared by electrolysis of sodium hydroxide (NaOH) and fractional distillation of liquid air.

Preparation of oxygen from hydrogen peroxide (H_2O_2)

Apparatus and materials

Flat bottomed flask, dropping funnel, delivery tube, gas jar, beehive (gas jar stand), water trough, water, hydrogen peroxide, manganese (IV) oxide, cork.

Setup of apparatus



Procedure

The gas jar is filled with water and inverted over a gas jar stand.

Manganese (IV) oxide is added in the flask. The apparatus is then arranged as shown above.

Hydrogen peroxide is added to the manganese (IV) oxide through the dropping funnel.

Observation

Effervescence occurs as a colorless gas (oxygen gas) is given off. The colorless gas is collected over water because it is slightly soluble in water or when it is not required dry.

Equation of reaction

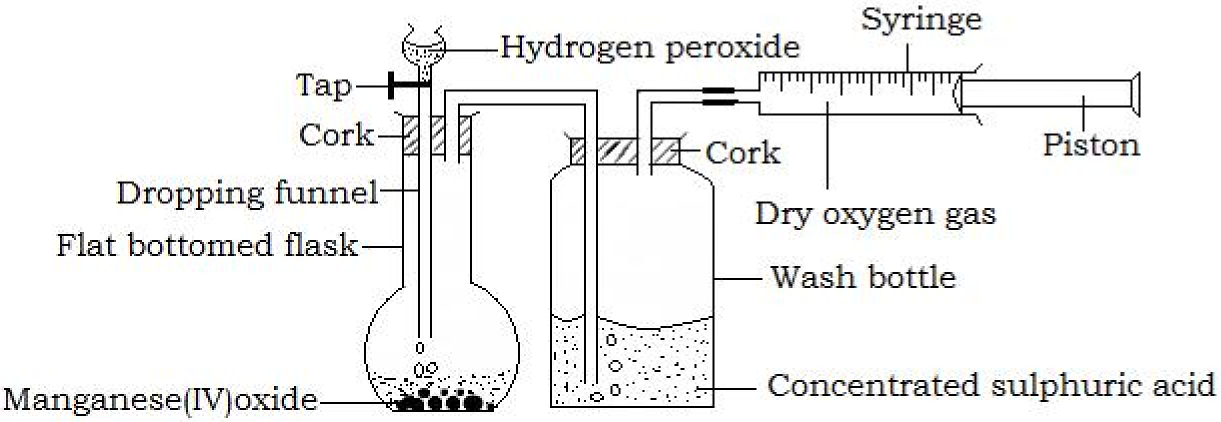
Word equation

$$\text{Hydrogen peroxide} \xrightarrow{\text{Manganese (IV) oxide}} \text{oxygen + water}$$

Symbolic equation

$$2H_2O_{2(aq)} \xrightarrow{MnO_2} O_{2(g)} + 2H_2O_{(l)}$$

NB. If the oxygen gas is required dry, it is passed through a wash bottle containing concentrated sulphuric acid as shown below or through a U tube containing calcium chloride and then collected using a syringe.



Purpose of Manganese (IV) oxide

Manganese (IV) oxide acts as a catalyst to speed up the rate of decomposition of hydrogen peroxide. Without it, the decomposition would take place but at a very slow rate. The rate of decomposition can also be increased by exposing the hydrogen peroxide to sunlight.

Using the same setup as above, hydrogen peroxide is added drop by drop from the funnel into potassium permanganate in the presence of dilute sulphuric acid. Oxygen is liberated until when all the potassium permanganate is decomposed and turns colorless.

Equation

$$5H_2O_2(aq)+2KMnO_4(aq)+3H_2SO_4(aq) \longrightarrow K_2SO_4(aq)+2MnSO_4(aq)+8H_2O(l)+5O_2(g)$$

A catalyst is a substance that speeds up the rate of a chemical reaction but remains unchanged chemically at the end of the reaction.

N.B. When sodium peroxide is used for the preparation of oxygen, the sodium peroxide is placed in the flat bottomed flask and water in the dropping funnel, then the water is run into the sodium peroxide and effervescence occurs as oxygen is given off.

Equation

Sodium peroxide + water ⟶ sodium hydroxide + oxygen gas

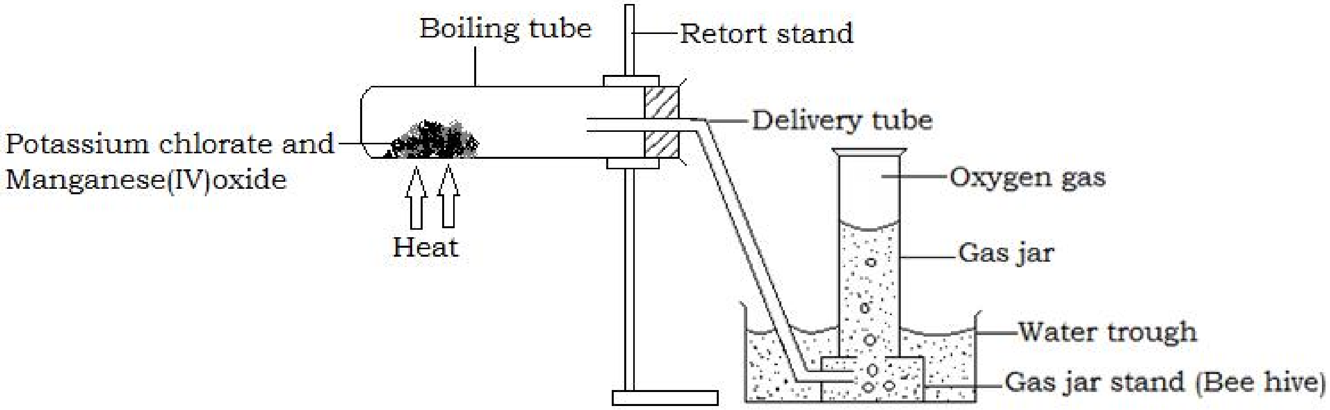
$$Na_2O_2(aq) + H_2O(l) \longrightarrow 2NaOH(aq) + \frac{1}{2}O_2(g)$$

Chemical test for oxygen

Oxygen relights a glowing splint. When a glowing splint is lowered into a gas jar containing oxygen, the glowing splint relights.

Preparation of oxygen from potassium chlorate

A mixture of potassium chlorate and manganese (IV) oxide is heated in a boiling tube as shown below.



Results

The potassium chlorate decomposes giving off colorless oxygen gas and white residue of potassium chloride.

Equation

Word equation

Potassium chlorate ⟶ potassium chloride + oxygen gas

Symbolic equation

$$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$$

Industrial preparation of oxygen

Oxygen is manufactured on a large scale by fractional distillation of liquid air. Air is passed through sodium hydroxide to remove carbon dioxide and through silica gel to dry the air i.e remove water vapor.(carbon dioxide and water vapor are removed because at low temperatures, they solidify and block the pipes).

Equation for removal of carbon dioxide.

$$2NaOH(aq) + CO_2(g) \longrightarrow Na_2CO_3(aq) + H_2O(l)$$

The dry air containing nitrogen, oxygen and noble gases is repeatedly compressed at very high pressure (about 200 atmospheres) and cooled until liquid air is obtained. The liquid air appears blue because of the presence of oxygen. Fractional distillation of the liquid air is carried out to obtain oxygen which boils at -183˚C.

N.B. During fractional distillation of liquid air, nitrogen with the lowest boiling (-196˚C) point evaporates first followed by argon (boiing point -186˚C) and then oxygen (boiling point -183˚C)

Liquefaction of air

The dry air containing nitrogen, oxygen and noble gases is compressed at high pressures of up to 200 atmospheres. The compression of air takes place with evolution of heat. As the air comes out of the compressor, it passes through tinny nozzles and a cooling system, in the process the compressed mixture is cooled. The mixture is compressed and cooled repeatedly as the temperatures get lower and lower until when liquefaction of air occurs.

Properties of oxygen

a) Physical properties

- It's odorless, tasteless and colorless.

- It's slightly soluble (sparingly soluble) in water. This is why it is collected over water.
- It is approximately the same density as that of air.
- It is a neutral gas i.e. it has no effect on litmus paper.

b) Chemical properties

Oxygen supports burning (combustion). Most metals and non metals burn in oxygen forming basic and acidic oxides respectively.

Metal + oxygen ⟶ metallic oxide (most of which are basic)

Non metals + oxygen ⟶ non metallic oxides (most of which are acidic)

Reaction of oxygen with metals

Metals burn in oxygen forming basic oxides which when dissolved in water forms basic or alkaline solution i.e. solutions the turn red litmus paper blue and have effect on blue litmus paper.

1. Magnesium

Magnesium burns in air(oxygen) with a brilliant white flame forming white ash (powder) of magnesium oxide.

Magnesium + oxygen ⟶ magnesium oxide

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

The magnesium oxide dissolves slightly in water forming an alkaline solution of magnesium hydroxide.

$MgO(s) + H_2O(l) \longrightarrow Mg(OH)_2(aq)$

2. Sodium

Sodium burns in excess oxygen with a bright yellow flame to form a yellow powder of sodium peroxide.

$2Na(s) + O_2(g) \longrightarrow Na_2O_2(s)$

The sodium peroxide dissolves in water forming an alkaline solution of sodium hydroxide with evolution of a colorless gas that relights a glowing splint (oxygen).

However in limited supply of oxygen, sodium burns with a bright yellow flame to form white solids of sodium oxide.

$4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$

N.B. Sodium and potassium are kept under oil/paraffin since they are very reactive and react violently with both air and water.

The sodium oxide dissolves in water forming an alkaline solution of sodium hydroxide only.

$Na_2O(s) + H_2O(g) \longrightarrow 2NaOH(aq)$

3. Calcium

Calcium burns in air with a bright red flame forming white solids of calcium oxide.

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

The calcium oxide slightly dissolves in cold water forming an alkaline solution of calcium hydroxide. The calcium hydroxide appears cloudy due to the presence of undissolved calcium oxide.

4. Potassium

Potassium burns in oxygen with a lilac flame (purple flame) forming white ash of potassium oxide.

$2K(s) + \frac{1}{2}O_2(g) \longrightarrow K_2O(s)$

The potassium oxide dissolves in water to form an alkaline solution of potassium hydroxide.

$K_2O(s) + H_2O(aq) \longrightarrow 2KOH(aq)$

5. Iron

Oxygen burn is oxygen ith a shower of bright sparks leaving behind blue black solids of Tri iron tetra oxide.

$3Fe(s) + 2O_2(g) \longrightarrow Fe_3O_4(s)$

The oxide of iron formed is insoluble and therefore has no effect on litmus paper.

6. Copper

Copper metal burns with a blue flame turning red hot, on cooling forms black power of copper (II) oxide.

$2Cu(s) + O_2(g) \longrightarrow 2CuO(s)$

The copper (II) oxide is insoluble in water and therefore has no effect on litmus paper.

7. Lead

Lead melts into shinny beads and finally forming yellow powder of lead (II) oxide.

$2Pb(s) + O_2(s) \longrightarrow 2PbO(s)$

Reaction of non metals with oxygen

Non-metals burn in air to form acidic oxides (acid andydrides). These oxides dissolve in water to form acidic solutions that turn blue litmus paper red and have no effect on red litmus paper.

An **acid anhydride** is an oxide of a non metal that dissolves in water to form an acid. Examples include:

Acid andydride	**Acid formed**
Carbondioxide	Carbonic acid
Sulphur trioxide	Sulphuric acid

Sulphur dioxide	Sulphurous acid
Phosphorus (V) oxide	Phosphoric acid
Phosphorus(III) oxide	Phosphorus acid

1. Carbon

Carbon burns in oxygen with an orange flame and bright sparks forming a colorless gas that turns lime water milky (carbondioxide gas).

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

The carbondioxide gas dissolves in water forming a weakly acidic solution of carbonic acid. This solution turns blue litmus paper pink and not red.

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

2. Phosphorus

Phosphorus burns in air with a bright yellow flame forming white clouds (fumes) of a oxides of phosphorus. The white fumes is a mixture of phosphorus(V)oxide and phosphorus (III)oxide.

$4P(s) + 3O_2(g) \longrightarrow 2P_2O_3(s)$(phosphorus (III)oxide)

$4P(s) + 5O_2(g) \longrightarrow 2P_2O_5(s)$(phosphorus(V)oxide)

Phosphorus(V)oxide and phosphorus (III)oxide dissolve in water forming phosphoric and phosphorus acids respectively.

$P_2O_5(s) + 3H_2O(l) \longrightarrow 2H_3PO_4(aq)$ (Phosphoric acid)

$P_2O_3(s) + 3H_2O(l) \longrightarrow 2H_3PO_3(aq)$ (Phosphorus acid)

N.B. Phoshorus is kept in water because if it is in contact with air, it smolders (burns without a flame slowly giving off white smoke)

3. Sulphur

Sulphur burns in air with a bright blue flame forming cloudy fumes which have a choking smell. The white fume is a mixture of sulphur dioxide and sulphur trioxide.

$2S(s) + 2O_2(g) \longrightarrow SO_2(g)$(sulphur dioxide)

$2S(s) + 3O_2(g) \longrightarrow 2SO_3(g)$(sulphur trioxide)

The sulphur dioxide and sulphur trioxide dissolve in water forming sulphurous and sulphuric acids respectivey.

$SO_2(g) + H_2O(l) \longrightarrow H_2SO_3(aq)$ (Sulphurous acid)

$SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq)$ (sulphuric acid)

Uses of oxygen

- Oxygen is used by living things in the process of respiration.
- Oxygen is essential for combustion or burning therefore providing heat source for various activities like cooking, burning wastes e.t.c.
- It is used to aid breathing where the natural supply of oxygen is insufficient e.g. in high altitude flying or climbing, and also in hospitals for patients.
- Oxygen when mixed with ethyne produces a very hot flame (oxy acetylene flame) which is used for welding and cutting heavy metals.
- Oxygen is used in te manufacture of steel i.e. in the conversion of pig ifron to steel.
- Liquid oxygen is used as fuel in space rockets.
- Liquid oxygen can be used as explosives in mines when mixed with charcoal and petrol.

Classes of oxides

An oxide is a compound of oxygen with another element. The following are major classes of oxides (types of oxides).

1. Basic oxides

These are oxides of metals which react with acids to form salt and water only. E.g. magnesium oxide, copper(II) oxide and calcium oxide.

Magnesium oxide reacts with hydrochloric acid to produce magnesium chloride (salt) and water.

$MgO(s) + 2HCl(l) \longrightarrow MgCl_2(aq) + H_2O(l)$

Copper (II) oxide reacts with sulphuric acid to form copper (II) sulphate (salt) and water.

$CuO(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + H_2O(l)$

2. Acidic oxides

These are oxides of non metals that dissolve in water to form acids. These oxides are called acid anhydrides. E.g. Carbondioxide (CO_2), sulphurdioxide (SO_2), sulphur trioxide (SO_3), nitrogen dioxide(NO_2) e.t.c.

Carbondioxide dissolves in water to form carbonic acid

$CO_2(g) + H_2O(l) \longrightarrow H_2CO_3(l)$

Sulphur trioxide reacts with water to form sulphuric acid

$SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq)$ (sulphuric acid)

These oxides react with alkalis to form salt and water. E.g. carbondioxide reacts with sodium hydroxide to form sodium carbonate(salt) and water.

$CO_2(g) + 2NaOH(aq) \longrightarrow Na_2CO_3(aq) + H_2O(l)$

3. Neutral oxides

These are oxides of non metals which do not show acidic or basic properties (they are neither acidic nor basic.). I.e. they don’t react with water to form acids and neither do they react with acids to form salts.E.g carbon monoxide (CO), nitrogen monoxide (NO),dinitrogen oxide (N_2O),water (hydrogen oxide).

4. Amphoteric oxides

These are metallic oxides which show both acidic and basic properties. i.e they react with both acids and alkalis to form salt and water. E.g aluminium oxides(Al_2O_3), lead(II)oxide(PbO), zinc oxide(ZnO).

Reaction of the amphoteric oxides with acids

Zinc oxide reacts with sulphuric acid to form zinc sulphate (salt) and water.

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

Aluminium oxide reacts wth hydrochloric acid forming aluminium chloride (salt) and water.

$$Al_2O_3(s) + 6HCl(aq) \longrightarrow 2AlCl_3(aq) + 3H_2O(l)$$

Reaction of amphoteric oxides with alkalis

Zinc oxide reacts with sodium hydroxide to form sodium zincate (salt).

$$ZnO(s) + 2NaOH(aq) + H_2O(l) \longrightarrow Na_2Zn(OH)_4(aq)$$

Aluminium oxide reacts with sodium hydroxide forming a salt called sodium aluminate.

$$Al_2O_3(s) + 2NaOH(aq)+ 3H_2O(l) \longrightarrow 2NaAl(OH)_4(aq)$$

5. Peroxides

These are oxides that contain twice as much oxygen as they would contain in their normal oxides. E.g. hydrogen peroxide (H_2O_2), Sodium peroxide (Na_2O_2).

6. Mixed oxides

These are oxides that react like a mixture of two simpler oxides. E.g. lead tetra oxide (Pb_3O_4) and Tri iron tetra oxide (Fe_3O_4).

Electrochemical series (Reactivity series)

This is the arrangement of elements in the order of their readiness to react with other substances. The most reactive elements are placed at the top of the series and the least reactive elements are placed at the bottom.

Metal competition for oxygen

Metals which are higher in the reactivity series (more reactive metals) tend to remove oxygen from the oxides of metals that are lower in the reactivity series (less reactive metal). Consider the reaction below;

$$Mg(s) + CuO(s) \longrightarrow Cu(s) + MgO(g)$$

Magnesium is more reactive than copper and therefore removes oxygen from it.

Comparing reactivity of metals

1. **Potassium and sodium**

When potassium and sodium are freshly cut, they are silvery shiny but their surfaces rapidly tarnish in air because they react quickly with oxygen in the atmosphere to form oxides. Potassium tarnishes more rapidly than sodium indicating that it is more reactive than sodium.

2. **Magnesium**

When a piece of magnesium is polished, tinny silvery surface is exposed. The surface remains shinny for some times before turning grey due to the slow reaction of the magnesium with air forming magnesium oxide (the grey coating). Magnesium is less reactive than sodium and potassium.

3. **Aluminium and Zinc**

Aluminium and zinc are shinny metals although their surfaces get coated with thin layers of their respective oxides on exposure to oxygen. Aluminium and zinc react more slowly with oxygen than magnesium forming their oxides. Hence aluminium and zinc are less reactive than magnesium.

4. **Copper**

Copper is a brown solid and does not react with oxygen from the atmosphere but reacts with carbondioxide and water from the atmosphere to form a green solid of copper (II) carbonate. Copper is therefore less reactive than aluminium and zinc. From the aboive reactions, we can establish the reactivity series of metals.

Complete reactivity series of metals starting with the most reactive is as shown below.

Element	**Symbol**
Potassium	K
Sodium	Na
Calcium	Ca
Magnesium	Mg
Aluminium	Al
Carbon (metalloid)	C
Zinc	Zn
Iron	Fe
Lead	Pb
Hydrogen (non metal)	H
Copper	Cu
Mercury	Hg
Silver	Ag
Gold	Au

Reactions between metals and metal oxides

1. Magnesium and copper

When a mixture of magnesium and copper (II) oxide is heated, a red glow spreads through the mixture leaving a white ash and a brown solid.

Magnesium removes oxygen from copper to form magnesium oxide. The white ash is magnesium oxide and the brown solid is copper metal.

This shows that magnesium is more reactive than copper

$$Mg_{(s)} + CuO_{(S)} \rightarrow Cu_{(s)} + Mg_{(s)}$$

Brown solid White ash

2. Carbon and copper (II) Oxide

When a mixture of carbon and copper (II) Oxide is heated in a test tube, a brown solid is formed and a colorless gas that turns lime water milky (CO_2) is liberated. Carbon removes Oxygen from copper (II) Oxide to form carbon dioxide is more reactive than copper.

$$C_{(s)} + CuO_{(s)} \rightarrow Cu_{(s)} + CO_{2(g)}$$

Brown solid

3. Magnesium and Carbon dioxide

Magnesium burns in carbon dioxide to form a mixture of white ash and black solid particles

Magnesium removes oxygen from carbon dioxide to form magnesium oxide.

The white ash is magnesium oxide and the black particles are carbon

$$Mg_{(s)} + CO_{2(g)} \rightarrow Mg_{(s)} + C_{(s)}$$

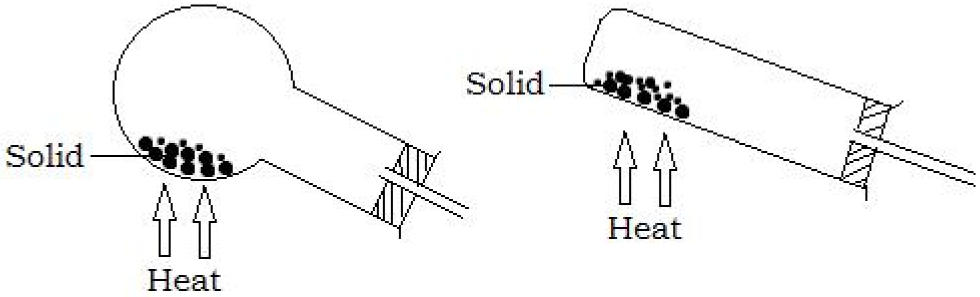
White ash Black particles

GAS GENERATION AND COLLECTION

Gas generation

Gases are produced by a number of chemical reactions e.g. action of heat on solids and reaction of acids with solids. Different setups for gas preparation are made depending on the conditions for the reactions and substances being used to prepare the gas.

a) If heat is to be applied to the reactants, a round bottom flask or a boiling tube is used. If water is one of the products produced by heating solids, the flask or boiling tube has to be placed in a slanting position (WHY?).



b) If the there is no heating involved, a conical flask with side arm or a flat bottom flask is used. The flask is placed on a flat surface.

c) If one of the reactants is a solid, and the other is an aqueous solution, the aqueous solution is added through a funnel (thistle or dropping) into the solids placed in the flask.

d) For continuous supply or a gas when solids react with aqueous solution, kipps apparatus is used.

e) Gases that are poisonous and or with unpleasant smell, are prepared in the fume cupboard. E.g. carbon monoxide and hydrogen sulphide.

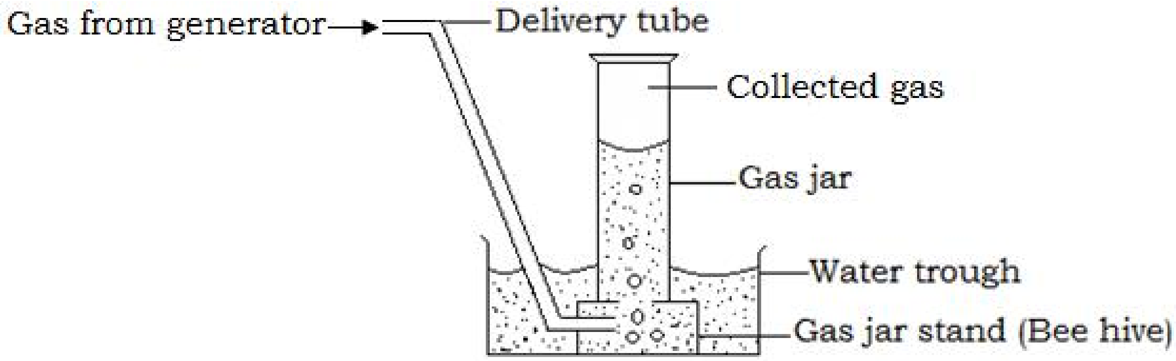
Gas collection

Gases produced are collected by different methods depending on the properties of the gases such as solubility and density. Methods of gas collection include:

1. Collection over water

In this method, the gas is bubbled through water into a gas jar stand filled with water and inverted on a gas jar stand (bee hive). The water displaces water downwardly and the method is some times referred to as **down ward displacement of water.**

Drawing



Gases collected by this method are:

- Gases that are not required dry
- Gases that are only slightly(sparingly) soluble in water

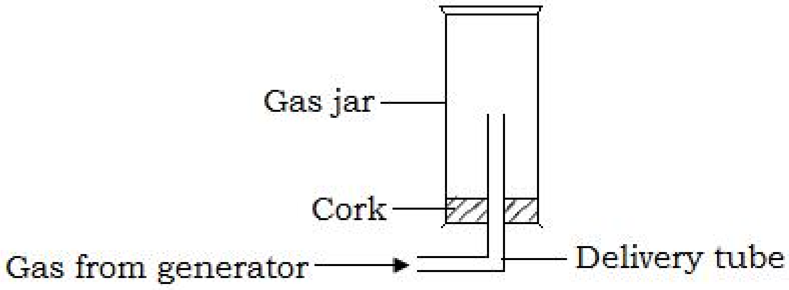
Examples of gases collected by this method include: oxygen, hydrogen, carbon monoxide and carbon dioxide.

Sulphurdioxide and chlorine are not collected over water because they react with it. Ammonia and hydrogen chloride are very soluble in water, therefore not collected over it.

2. Upward delivery method

The gas is passed through a delivery tube connected to an inverted gas jar as shown below. The gas collected displaces air downwardly and the method is also called **downward displacement of air.**

Drawing



This method is used to collect gases that are:

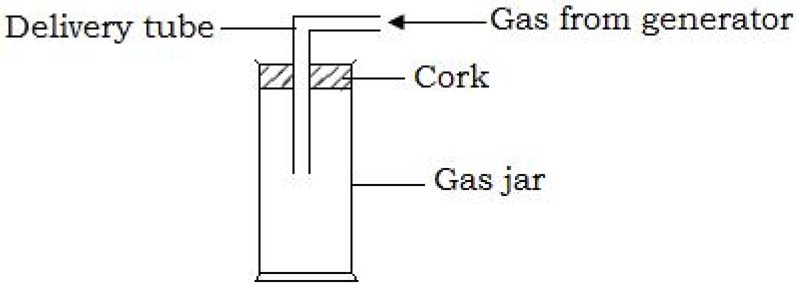
- Less dense than air
- Required dry
- Soluble in water

Examples of gases collected by this method include: ammonia and hydrogen.

3. Downward delivery method

The gas collected is passed through a delivery tube connected to a gas jar in an upright position. The gas displaces the air in the gas jar upwardly and the method is also referred to as **upward displacement of air**.

Drawing



This method is used to collect gases that are:

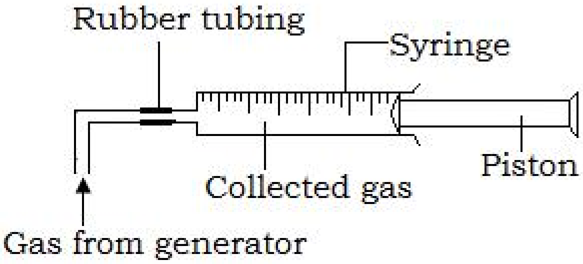
- Denser than air
- Required dry
- Soluble in water

Examples of gases collected by this method include: carbon dioxide, sulphur dioxide, chlorine and hydrogen chloride.

4. Use of a gas syringe

This method collects any gas irrespective of its density or solubility (i.e. whether the gas is soluble in water or not; whether the gas is denser than air or not). It is however suitable for collecting gases produced in small quantities.

Drawing



Advantages of using syringe method

- It is easy to know if the syringe is full of the gas or not.
- It is easy to transport the gas.
- Air does not easily mix with the gas collected, therefore it is almost pure.
- Collects any gas irrespective of its properties like solubility and density.
- Volume of gas collected is known and correct volumes of gases are used.

DRYING SUBSTANCES

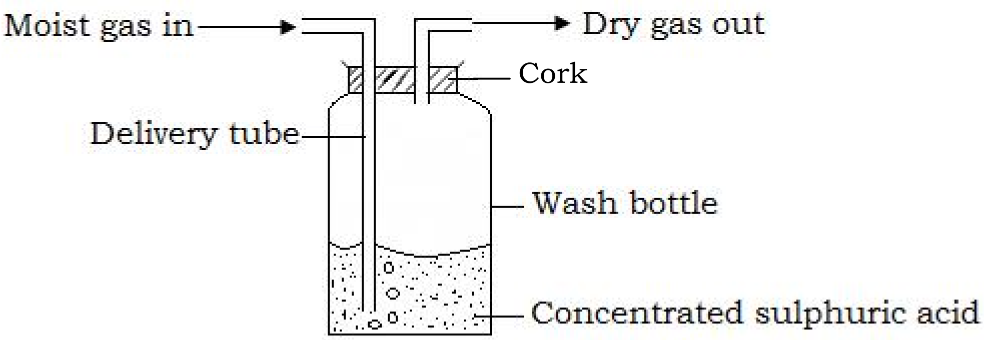
Substances that contain water can be dried by drying agents or desiccants.

Desiccants are substances that have very high affinity for water (water loving compounds) and therefore remove water from other substances. Desiccants can either be hygroscopic or deliquescent substances.

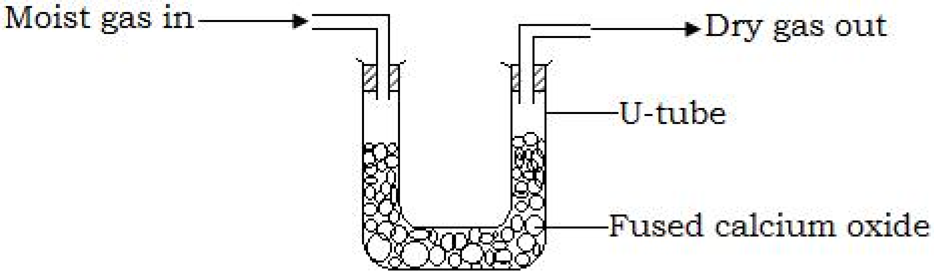
1. Drying gases

When gases are required dry, they are passed through liquid and solid desiccants. Examples of these desiccants include: concentrated sulphuric acid, anhydrous calcium chloride, calcium oxide and silica gel.

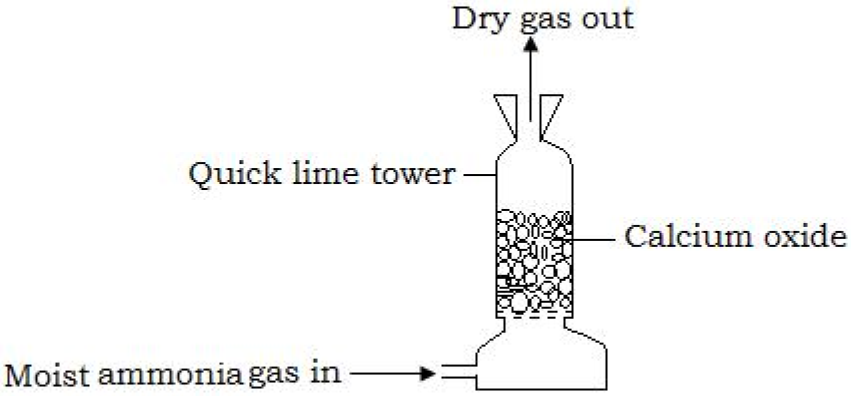
Concentrated sulphuric acid placed in a vessel is used for drying gases like sulpur dioxide, oxygen, carbondioxide, chlorine and oxygen. The gas is bubbled through concentrated sulphuric acid using a delivery tube as shown below.



Gases like ammonia and hydrogen sulphide react with concentrated sulphuric acid and are not dried by it. Instead solid desiccants like fused calcium chloride and calcium oxide (quick lime) placed in a U-tube or quick lime tower is used.

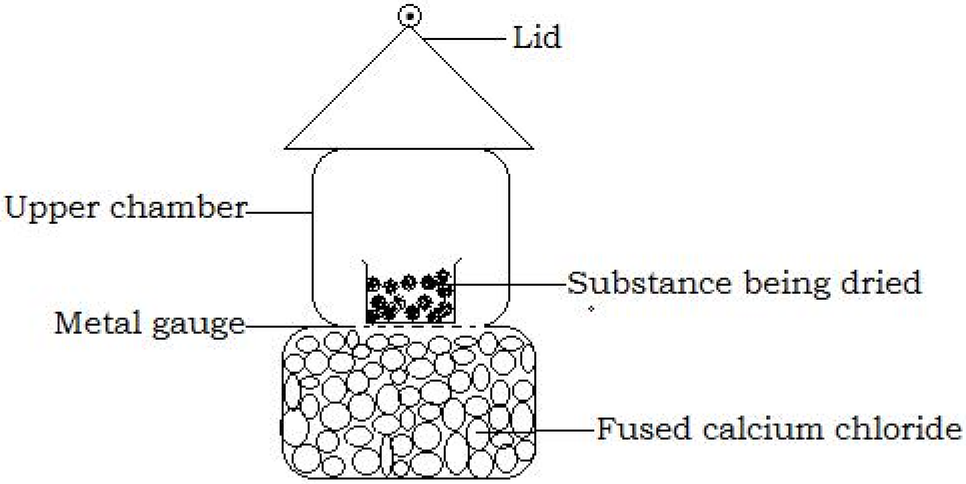


Gases dried through the U- tube include oxygen, carbon dioxide, hydrogen and hydrogen sulphide. Ammonia is dried is a quick lime tower as shown below.



2. Drying liquid

Organic liquids are the only liquids that can be dried as some contains little amount of water. The liquid can be dried by the process of fractional distillation or by use of desiccants such as fused calcium chloride, calcium oxide, magnesium sulphate e.t.c. placed in a desiccator as shown below.



A desiccator consists of a lid, upper chamber and lower chamber. The substance to be dried is placed in the upper chamber and the drying agent is placed in the lower chamber. The circumference of the lid is smeared with oil or grease to make it ai air proof and the substance remains dry as long as it is in the desiccators.

3. Drying solids

Solids with moisture can be dried by slight warming or by evaporating water from the solid at a temperature slightly below the melting point of the solids. Solids can also be dried by placing them in a special drying apparatus (desiccators) as for liquids.

Sample questions on the atmosphere and Oxygen

The atmosphere

1. *What is the approximate percentage composition of the atmosphere by volume?. Descried a laboratory experiment by which you could verify the figure you give for*

the percentage of one element in the atmosphere. Name four substances found in the atmosphere in small quantities.

2. *Air is considered a mixture. Briefly explain why it is regarded as a mixture.*
3. *What is rust?. State the conditions necessary for rusting to occur. What steps can be taken to prevent iron from rusting?. Descried experiments to prove that iron does not rust unless when both air and water are present. State briefly three ways in which rust differs from a mixture of iron and oxygen.*
4. *What is air pollution? Mention at least five environmental pollutants and briefly explain how each pollutes the environment. Give at least four sources of environmental pollutants. Outline some measures that can be taken t reduce on environmental pollution.*
5. *Describe an experiment you can carry out to investigate the products formed when a burning candle is burnt in air.*
6. *How would you prove experimentally that air contains (a) water vapour (b) carbon dioxide (c) oxygen in the laboratory.*

Oxygen

1. *Describe with the aid of labeled drawings how you can prepare oxygen in the laboratory from: hydrogen peroxide, potassium chlorate and sodium peroxide. During the preparation of oxygen a catalyst is added, name the catalyst and mention its role. What is the proof that the named substance is actually acting as a catalyst.*
2. *Describe a simple experiment to show what fraction of air is oxygen. State briefly how oxygen is prepared from air on a large scale. Mention three uses of oxygen.*
3. *With the aid of equations, describe the reaction of oxygen/air with sodium, potassium, calcium, magnesium, copper, carbon, sulphur, and phosphorus.*
4. *Non metals react with oxygen to form acid anhydrides, what is an acid anhydride? What is an oxide? Describe with the aid of equations and examples the different types of oxides formed by elements.*
5. *Explain the observation that*
    a) *Magnesium burns in carbon dioxide to form a mixture of white ash and black solid particles. Use an equation to illustrate.*

   b) *When a mixture of magnesium and copper (II) oxide is heated, a red glow spreads through the mixture leaving a white ash and a brown solid.*
   c) *When a mixture of carbon and copper (II) Oxide is heated in a test tube, a brown solid is formed and a colorless gas that turns lime water milky* (CO_2) *is liberated.*
6. *Sodium is higher in the electrochemical series than copper. Explain what this statement means.*
7. *Describe the different ways of collecting gases, mentioning clearly the examples of gasses collected by each method and why. Drawings must be included as part of the description. Outline the advantages of using a syringe over other methods to collect gases.*
8. *What is a desiccant? Explain briefly how gases, liquids and solids are dried.*

ATOMIC STRUCTURE

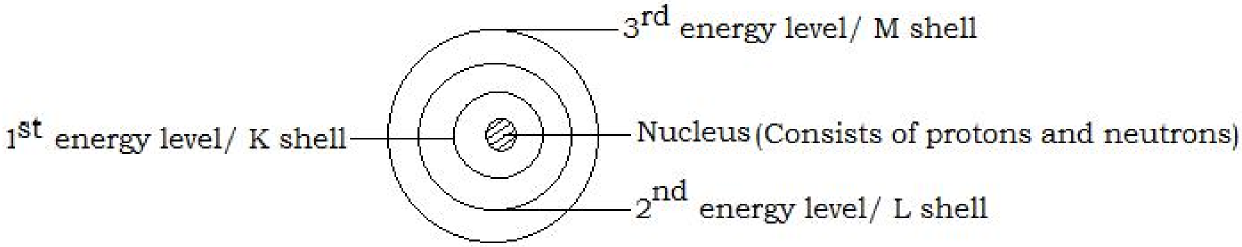
An atom is the smallest particle of an element that takes part in a chemical reaction. All atoms of the same element are identical and different from those of other elements.

Shape of an atom

An atom is spherical in shape and has a small region in the centre called the **nucleus**. The nucleus is surrounded by circular paths known as **electron shells** or **energy levels** that carry electrons. These shells are represented as circles.

The energy level (shell) closest to the nucleus is called the 1st energy level or the K shell. The 2nd closest shell to the nucleus is the 2nd energy level or the L shell. This is followed by the 3rd shell or the M shell. So, the energy levels are named 1st, 2nd, 3rd, 4th, e.t.c. or K, L, M, N e.t.c. from the shell closest to the nucleus outwardly.

General structure of an atom



An atom is composed of particles namely; protons, neutrons and electrons.

Protons (p)

Protons occur in the nucleus; Protons have a unit mass of one (1); each proton carries a charge of positive one (+1).

The total number of protons in the nucleus of an atom is the **atomic number**. The atomic number of each element is unique and it is used to identify the element. All atoms of the same element have the same atomic number.

Neutrons (n)

Neutrons occur in the nucleus of an atom; it has a unit mass of one i.e. the mass is approximately the same as that of a proton; a neutron carries no charge(it is nuetral).

The sum of all protons and all neutrons in the nucleus of an atom is called the **atomic mass** or **mass number**.

Electrons (e)

Electrons occur in the energy levels or electron shells and are constantly rotating around the nucleus; an electron has a mass of 1/1840 of a proton, there fore its mass is said to be negligible; it carries a charge of negative one(-1).

N.B. The number of electrons in the energy levels (negative charges) is equal to the number of all protons in the nucleus (positive charge) hence an atom is neutral because the number of positive charges(protons) is equal to the number of negative charges(electrons).

Mass number and Atomic number

Atomic number (Z)

Every atom has its specific atomic number. This atomic number is the same for all atoms of the same element.

Atomic number is the total number of protons in the nucleus of an atom (of an element). An atom being neutral has equal number of protons and electrons.

Mass number/Atomic mass (A)

Mass number is the sum of protons and neutrons in the nucleus of an atom of an element. The protons and neutrons occupy the nucleus and are referred to as **nucleons**.

Mass number = number of protons + number of electrons

Notation of an atom

An atom can be represented as

$^{A}_{Z}W$ Where W is the symbol of an element; A is the atomic mass or mass number and Z is the atomic number.

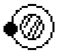
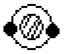
Electron arrangement in atoms

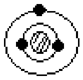
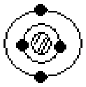
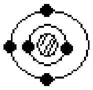
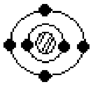
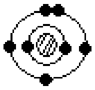
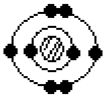
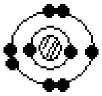
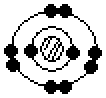
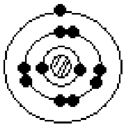
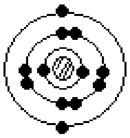
The arrangement of electrons in an atom is the **electronic configuration** and structures drawn showing the distribution of electrons in the energy levels is the electronic structure. In writing electronic configurations or structure, the following rules are considered.

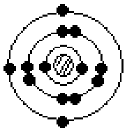
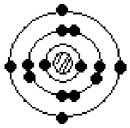
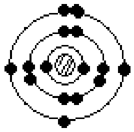
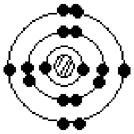
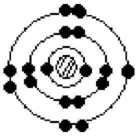
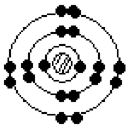
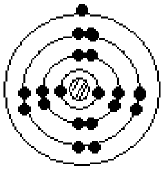
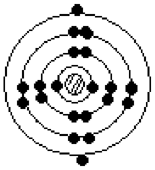
1. The 1st electron shell (one that is nearest to the nucleus) has the lowest energy level and electrons fill it first.
2. The 1st electron shell or energy level holds a maximum of two electrons. It cannot hold more than two electrons.
3. The 2nd electron shell holds a maximum of eight electrons. It can hold less the maximum number but never beyond the maximum number.
4. In the outer most shell of any atom, the maximum number of electrons possible is 8.

Electron shells are numbered 1, 2, 3, etc, or L, M, N e.t.c.outwards from the nucleus. All electrons in a given shell have approximately equal energy. This energy increases in successive shell outwards from the nucleus.

Table showing the electron arrangement of the first 20 elements

Element	**Symbol**	**No. of electrons**	**No. of electrons in each shell**				**Electronic configuration**	**Electronic structure**
			1st	2nd	3rd	4th		
Hydrogen	H	1	1				1	
Helium	He	2	2				2	
Lithium	Li	3	2	1			2;1	

								
Beryllium	Be	4	2	2			2;2	
Boron	B	5	2	3			2;3	
Carbon	C	6	2	4			2;4	
Nitrogen	N	7	2	5			2;5	
Oxygen	O	8	2	6			2;6	
Fluorine	F	9	2	7			2;7	
Neon	Ne	10	2	8			2;8	
Sodium	Na	11	2	8	1		2;8;1	
Magnesium	Mg	12	2	8	2		2;8;2	

Aluminium	Al	13	2	8	3		2;8;3	
Silicon	Si	14	2	8	4		2;8;4	
Phosphorus	P	15	2	8	5		2;8;5	
Sulphur	S	16	2	8	6		2;8;6	
Chlorine	Cl	17	2	8	7		2;8;7	
Argon	Ar	18	2	8	8		2;8;8	
Potassium	K	19	2	8	8	1	2;8;8;1	
Calcium	Ca	20	2	8	8	2	2;8;8;8;2	

Isotopes and Isotopy

Isotopes are atoms of the same element with the same number of protons but different number of neutrons. Isotopes therefore have different mass numbers. For elements that show isotopy, the most abundant (common) isotope is taken to be the representative of all the element. The abundance is usually given as a percentage.

Isotopy is the existence of atoms of the same element with the same number of protons but different number of neutrons.

Examples of isotopes

Element	Symbol	Atomic number	Isotopes	Abundance
Hydrogen	H	1	${}^{1}_{1}H$ ${}^{2}_{1}H$ ${}^{3}_{1}H$	99.99% 0.01% Rare
Carbon	C	6	${}_{6}C$ ${}^{13}_{6}C$ ${}^{14}_{6}C$	98.9% 1.1% Trace
Chlorine	Cl	17	${}_{17}Cl$ ${}^{37}_{17}Cl$	75% 25%

Example

Copy and complete the tabe below.

Relative Atomic Mass (RAM)

This is the mass of one atom of an element compared to 1/12 (a twelfth) of the mass of one atom of Carbon 12 isotope.

R.A.M $=\frac{mass\ of\ an\ atom\ of\ \ an\ element}{\frac{1}{12}of\ the\ mass\ of\ one\ atom\ of\ C-12\ isotope}$

Relative Atomic Mass of an element that shows isotopy is dependent on the relative abundance of each isotope and the atomic masses of the isotopes.
R.A.M is a ratio of the same quantity and therefore has no units.

Example
Neon has three isotopes, $^{20}_{10}Ne$(percentage abundance 90.5%); $^{21}_{10}Ne$ (percentage abundance of 0.3%) and $^{22}_{10}Ne$ (percentage abundance of 9.2%)

Solution

$\left(\frac{90.5\times20}{100}\right)+\left(\frac{0.3\times21}{100}\right)+\left(\frac{9.2\times22}{100}\right)$

$\left(\frac{(90.5\times20)+(0.3\times21)+(9.2\times22)}{100}\right)$

$\left(\frac{2018.7}{100}\right)$ =20.187.

THE PERIODIC TABLE

The periodic table is the arrangement of elements in order of increasing atomic masses. The periodic table is made up of rows and columns. The horizontal rows are called **periods** and are assigned numbers according to the number of electron shells/energy levels that are filled by electrons. All elements in the same period have the same number of electron shells /energy levels. The periods are written in roman numbers as 1, 2, 3, etc.
The vertical columns are referred to as **groups** and the group number is written in Roman numeral on top of the group (column). The group number represents the number of electrons in the outer most shell/energy level. Elements in a particular

group resemble each other chemically because they have the same number of electrons on the outer most shell/energy level.

Periodic table of the first 20 elements

Group

		I	II		III	IV	V	VI	VII	VIII
Period	1	H								He
	2	Li	Be	▨	B	C	N	O	F	Ne
	3	Na	Mg	▨	Al	Si	P	S	Cl	Ar
	4	K	Ca							

The elements in the periodic table are classified as metals, metalloid, non metals and noble/inert/rare gases.

i) **Metals**: these are on the left of the periodic table in group I, II and III. Metals react by loss of electrons.

ii) **Semi metals (metalloids)**: These are found in the middle of the periodic table in group IV.

iii) **Non metals**: they are on the right of the periodic table in groups V, VI and VII. Non metals react by gain of electrons.

iv) **Noble/rare of inert gases**: They are at the extreme right of the periodic table in group VIII (O). Non metals are not reactive because the are very stable.

Transition elements: these are found between group II and III and towards the bottom of the periodic table. Transition elements form colored compounds and have variable valencies. Examples of transition elements include copper, iron, chromium, cobalt etc.

Trend of reactivity of elements in the periodic table

1. Elements in the same group react in similar ways because they have the same number of electrons in their outer most energy level/ shell.
2. Reactivity increases as you go down the group of metals because of increase in atomic size (atomic radius). As the size of atoms increase, the outer most electrons become loosely held because they are no longer strongly attracted by the nucleus hence such electrons are easily lost. Because metals react by loss of electrons, the more easily a metal loses electrons, the more reactive that metal is.
3. As you go down the group of non metals, reactivity decreases due to increase in atomic size (atomic radius). As the atomic sizes increase, the forces of attraction of the nucleus for the electrons decrease hence electrons cannot be gained readily. So, the more easily a non metal gains electrons, the more reactive that non metal is.
4. Across the period of elements from left to right, reactivity decreases from group I to group IV and then increases up to group VII and falls sharply in group VIII. This is because, as you go across the period, the effective force of the nucleus increases hence the outer most electrons become more strongly attracted.

IONS

An ion is a charged particle formed from an atom or a group of chemically combined atoms by gaining or losing one or more electrons. Ions are either negatively charged (anions) or positively charged (cations).

Ion formation

a) Positive ions(Cations)

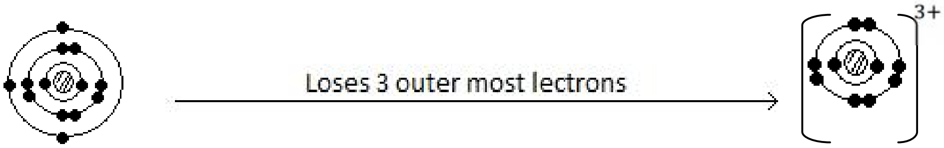
These are formed when atoms or group of chemically combined atoms lose electrons. Atoms lose electrons so as to acquire a stable electronic configuration as the noble gases. The positive ions (cations) are formed by metal atoms, hydrogen atoms and ammonium.

i) Formation of Aluminium ion, Al^{3+}

Aluminium atom with 13 electrons (electronic configuration 2.8.3), has 3 electrons more the stable configuration of noble gases,(2.8). During reaction, Aluminium loses the 3 electrons in the outer most energy level (electron shell) to become stable and form aluminium ion,Al^{3+}, (electronic configuration 2.8)

Aluminium atom (Al), unstable
Electronic configuration 2.8.3
Charge=0

Aluminium ion((Al^{3+}), stable
Electronic configuration 2.8
Charge=+3



Aluminum atom $\xrightarrow{\text{Loses 3 outer most electrons}}$ Aluminiumion

$$Al–3e \longrightarrow Al^{3+}$$

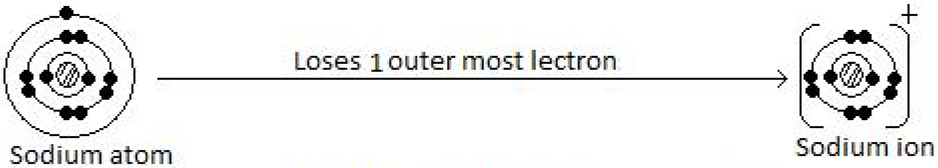
$$(Al \longrightarrow Al^{3+} + 3e)$$

ii) Formation of sodium ion, Na^{+}

Sodium atom with 11 electrons (electronic configuration 2.8.1), has 1 electrons more the stable configuration of a noble gases, (2.8). During reaction, sodium loses the 1 electron in the outer most energy level (electron shell) to become stable and form sodium ion,Na^{+}, (electronic configuration 2.8)

Sodium atom (Na), unstable
Electronic configuration 2.8.1
Charge=0

Sodium ion ((Na^{+}), stable
Electronic configuration 2.8
Charge=+1



$$Na - e \xrightarrow{\text{Loses 1 outer most lectron}} Na^{+}$$

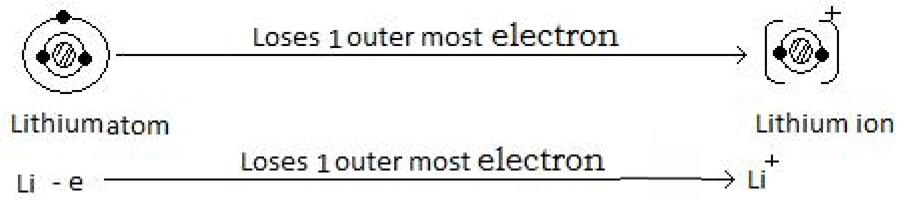
$$(Na \longrightarrow Na^{+} + e)$$

iii) Formation of lithium ion, Li^{+}

Lithium atom with 3 electrons (electronic configuration 2.1), has 1 electrons more the stable configuration of a noble gases, (2). During reaction, lithium loses the 1 electron in the outer most energy level (electron shell) to become stable and form lithium ion, Li^+, (electronic configuration 2)

Lithium atom (Li), unstable
Electronic configuration 2.1
Charge=0

Lithium ion((Li^+), stable
Electronic configuration 2
Charge=+1



$(Li \longrightarrow Li^+ + e)$

b) Formation of negative ions (Anions)

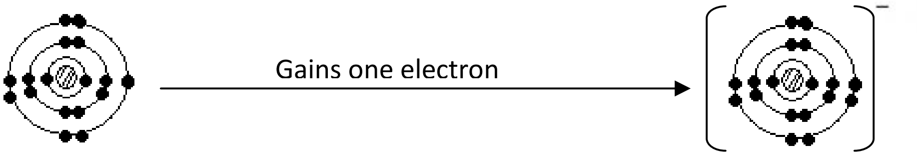
These are formed when atoms gain electrons. Atoms gain electrons so as to acquire a stable electronic configuration as the noble gases. The negative ions (anions) are formed by non metal atoms). Examples of anions are chloride ion and oxide ion.

i) Formation of a chloride ion, Cl^-

The ion is formed from chlorine atom. Chlorine atom (electronic configuration 2,8.7) lacks only one electron to attain the stable electronic configuration of a noble gas(2.8.8). The chlorine atom there fore gains the one electron during reaction and forms a chloride ion.

Chlorine atom (Cl), unstable
Electronic configuration 2.8.7
Charge=0

Chloride ion ((Cl^-), stable
Electronic configuration 2.8.8
Charge=-1



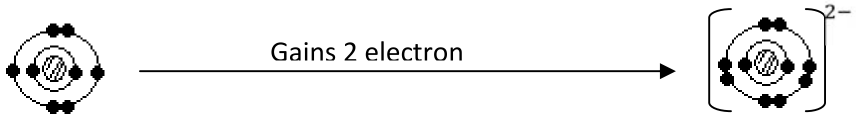
Chlorine atom Chloride ion

$$Cl+e^- \xrightarrow{\text{Gains one outer electron}} Cl^-$$

ii) Formation of an oxide ion, O^{2-}

The oxide ion is formed from an oxygen atom. Oxygen atom (electronic configuration (2,6) lacks two electrons to attain the stable electronic configuration of a noble gas(2.8). The oxygen atom therefore gains the two electrons during reaction and forms an oxide ion, O^{2-}.

Oxygen atom (O), unstable	Oxide ion ((O^{2-}), stable
Electronic configuration 2.6	Electronic configuration 2.8
Charge=0	Charge=-2



$$\text{Oxygen atom} \xrightarrow{\text{Gains two electrons}} \text{Oxide ion}$$

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

VALENCY

Valency is the combining power of an element or radical. In other words it is the measure of the power of an element to combine with others. The power of an element to combine with others is determined by the structure of its atoms. This therefore implies that the valency of an element can be determined from the electron structure of its atoms.

Elements with atoms having 1-4 electrons in the outer most shell, their valencies are equal to the number of electrons in the outer most shell.i.e. Group I elements have a valency of 1; group II elements have a valency of 2; group III elements have a valency of 3 and group IV elements have a valency of 4.

Elements in group V-VII, their valencies are equal to the number of electrons gained to acquire a stable electronic configuration i.e. Elements in group V gain three electrons each to make them stable, they therefore have a valency of 3; Elements in group VI gain two electrons each to make them stable, they therefore have a valency of 2 and elements in group VII gain one electron each to make them stable therefore have a valency of 1.

NB. Therefore, valency can also be defined as the number of electrons lost or gained by an element in order to attain a stable electronic structure.

Examples of elements with their valencies

Valency 1	**Valency 2**	**Valency 3**	**Valency 4**
Hydrogen (H)	Beryllium (Be)	Nitrogen (N)	Carbon (C)
Lithium (Li)	Magnesium (Mg)	Phosphorus (p)	Manganese (Mn)
Sodium (Na)	Calcium (Ca)	Boron (B)	Lead (Pb)
Potassium (K)	Oxygen (O)	Aluminium (Al)	Silicon (Si)
Silver (Ag)	Sulphur (S)	Iron (Fe)	Tin (Sn)
Copper (Cu)	Barium (Ba)		
Chlorine (Cl)	Zinc (Zn)		
Bromine (Br)	Iron (Fe)		
Iodine (I)	Lead (Pb)		
	Copper (Cu)		
	Mercury (Hg)		

NB. For metals with variable valencies, Roman numerals are included in their name to indicate the valencies of the metal. E.g. Iron (II) has a valency of 2,Lead (IV) has a valency of 4 etc.

RADICALS

A radical is an atom or group of atoms that exist in several compounds but does not exist on its own.

Valencies of common radicals

Radical	**Formula**	**Ion**	**Valency**
Hydroxide	OH	OH^-	1
Chloride	Cl	Cl^-	1
Nitrate	NO_3	NO_3^-	1
Nitrite	NO_2	NO_2^-	1
Ammonium	NH_4	NH_4^+	1
Hydrogen carbonate	HCO_3	HCO_3^-	1
Chlorate	ClO_3	ClO_3^-	1
Permanganate	MnO_4	MnO_4^-	1
Hydrogen sulphate	HSO_4	HSO_4^-	1
Hydrogen sulphite	HSO_3	HSO_3^-	1
Oxide	O	O^{2-}	2
Carbonate	CO_3	CO_3^{2-}	2
Sulphate	SO_4	SO_4^{2-}	2
Sulphite	SO_3	SO_3^{2-}	2
Sulphide	S	S^{2-}	2
Phosphate	PO_4	PO_4^{3-}	3
Nitride	N	N^{3-}	3

CHEMICAL FORMULAE

A chemical formula is a representation showing the proportions of elements present in a chemical compound using symbols.

Important concepts about chemical formulae

1. In a chemical formula, the number written as a subscript after the symbol of an element indicates the number of atoms of that element chemically combined.

Consider examples of the chemical formulae below.
In water, H_2O-the two (2) represents the number of hydrogen atoms chemically combined. There fore, water consist of two (2) hydrogen atoms and one (1) oxygen atom chemically combined together.

In sulphuric acid (H_2SO_4)- there are 2 atoms of hydrogen, 1 atom of sulphur and 4 atoms of oxygen chemically combined together.

$CaCO_3$ is the chemical formula of calcium carbonate (chalk), which consists of 1 atom of calcium, 1 atom of carbon and 3 atoms of oxygen chemically combined.

2. For groups of atoms (radicals), a bracket is used showing that they are being considered under the same valency.
   Examples
   In calcium nitrate, $Ca(NO_3)_2$,the 2 indicates that there are 2 nitrate radicals (NO_3). Both nitrogen and oxygen in nitrate are being considered under the same valency.

   In aluminium sulphate, $Al_2(SO_4)_{3,}$-the 3 shows that there are 2 sulphate radical (SO_4), both the sulphur and oxygen in the sulphate are being considered under the same valency. There fore, $Al_2(SO_4)_3$ consists of 2 atoms of aluminium and 3 sulphate radicals.

3. A number put infront of the formula of a compound indicates the number of molecules of the compound. For example
   $2H_2SO_4$ this means two molecules of sulphuric acid
   $8HNO_3$ means eight molecules of nitric acid
   $3CO_2$ means three molecules of carbondioxide
   H_2O means one molecule of water

4. To be able to write a chemical formula, one has to know the symbol and valency of the atoms or radicals.

Steps taken in writing chemical formulae

1. Identify from the name of the compound, the elements and radicals present e.g. in sodium chloride, there is sodium and chloride radical.
2. Write the symbol of the element or the formula of the radical separately, beginning with the one of a metal or ammonium radial followed by a non metal or radical.

$Na \qquad Cl$

3. Write the valencies as superscripts to the right of the symbols written separately.

$Na^1 \qquad Cl^1$

Compare the valencies and if the valencies are;

i) If the valencies are the same, counsel them out and write the symbols/formula close to each other.

$Na^{\cancel{1}} \qquad Cl^{\cancel{1}}$

$NaCl$

ii) If the valencies are not the same, but have a common factor, reduce them to the simplest ratio then interchange the valencies and write them as subscripts to the right of the symbol or formula of the radical. e..g. in Lead (IV) oxide

$Pb^{42} \qquad O^{\cancel{2}1}$

Pb_1O_2 the 1 is usually ignored and the formula is written as PbO_2.

iii) If the valencies are not the same and have no common factor, interchange them and write them as subscripts to the right of the symbol or formulae of radical. e.g. in aluminium oxide

$Al^3 \qquad O^2$

$$\begin{tikzcd} Al^3 \arrow[r, bend right] & O^2 \arrow[l, bend left] \end{tikzcd}$$

Al_2O_3

N.B. If the formula of the radical consists of more than one symbol, the formula must be put in brackets before the subscript is written e.g. in ammonium sulphate

$$\begin{tikzcd} NH_4{}^1 \arrow[r, bend right] & SO_4{}^2 \arrow[l, bend left] \end{tikzcd}$$

$(NH_4)_2SO_4$

Example

Write the formulae of the following compounds

a) Sodium chloride b) sodium sulpate c) Iron(III) oxide

Solution

a) Sodium Chloride

Na Cl

$Na^1 \qquad Cl^1$

NaCl

b) Sodium Sulphate

Na SO_4

$$\begin{tikzcd} Na^1 \arrow[r, bend right] & (SO_4)^2 \arrow[l, bend left] \end{tikzcd}$$

Na_2SO_4

c) Iron(III) Oxide

Fe O

$Fe^3 \qquad O^2$

$$\begin{tikzcd} Fe^3 \arrow[r] & O^2 \arrow[l] \end{tikzcd}$$

Fe_2O_3

Exercise

Write the formula of the following compounds

a) Potassium hydroxide b) Zinc carbonate c) magnesium sulphate d) ammonium phosphate e) copper(II) oxide f)potassium permanganate g) calcium hydroxide.

CHEMICAL EQUATIONS

This is the representation of a chemical change (chemical reaction) by means of symbols and formulae. So, when chemical equations are written, symbols and formulae are normally used. Word equations are not significant and not commonly used. A chemical reaction consists of reactants and products.

Consider the example below for illustration

Hydrogen reacts with oxygen to produce water vapor

Word equation: $\text{Hydrogen} + \text{Oxygen} \longrightarrow \text{water vapor}$

Symbol equation: $\underbrace{2H_2(g) + O_2(g)}_{\text{Reactants}} \longrightarrow \underbrace{2H_2O(g)}_{\text{Product}}$

Components of a chemical equation

a) **Formulae of reactants and products**. These represent compounds reacting and compounds being formed during a chemical reaction respectively.

b) **State symbols**. A state symbol is one small letter or two small letters that represent the physical state of a compound. They are written after the formulae and enclosed in brackets. There are four state symbols used. These are aqueous solution (aq); liquid (l); solid state(s) and gaseous state (g).

c) **The plus (+) sign**. The plus sign on the left hand side of the equation means 'react with' and the one at the right hand side of the equation means 'and' .

d) **The arrow**. This means to produce and the arrow head points to the products.

Interpretation of equations

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

This means 2 molecules of solid magnesium reacts with 1 molecule of oxygen gas to produce 2 molecules of solid magnesium oxide.

$$2H_2O_2(aq) \longrightarrow 2H_20(l) + O_2(g)$$

2 molecules of aqueous hydrogen peroxide produce 2 molecules of liquid water and gaseous oxygen.

For an equation to be considered correct, it must have correct formulae of reactants and products; correct state symbols and must be balanced.

Balancing is the process of making each kind of atoms on both the reactant and product sides equal. A balanced equation there fore has equal numbers of each kind of atoms on both the reactant and product sides.

N.B. When balancing a chemical equation, the formula of the compound is not altered but rather it is the number of molecules that is adjusted.

Steps in writing chemical equations

1. Write the correct formulae of the reactants on the left hand side and the correct formulae of products on the right hand side of the equation. The reactants and products are separated by an arrow pointing to the produncts.

   e.g. consider the reaction calcium and hydrochloric acid producing calcium chloride and hydrogen gas.

   $$Ca + HCl \longrightarrow CaCl_2 + H_2$$
2. Write the state symbol after each formula or symbol and enclose them in brackets.

   $$Ca(s) + HCl(aq) \longrightarrow CaCl_2(aq) + H_2(g)$$
3. Balance the equation by making the number of each kind of atoms on both reactant and product side equal. Adjust the number of molecules of the reactants and products to balance the number of atoms.

   $$Ca(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2(g)$$

Exercise

1. Write the equation of reaction between the following

a) Sodium + chlorine → Sodium chloride
b) Zinc + Sulphur → Zinc sulphide
c) Magnesium + Oxygen → Magnesium oxide
d) Hydrogen + Chlorine → Hydrogen chloride
e) Copper + chlorine → Copper chloride

2. Complete and balance the following equations.
   a) $Pb(s) + O_2(g) \longrightarrow$
   b) $H_2(g) + O_2(g) \longrightarrow$
   c) $Na(s) + H_2O(l) \longrightarrow$
   d) $Mg(s) + CO_2(g) \longrightarrow$
   e) $Fe(s) + HCl(aq) \longrightarrow$

ATOMICITY

Atomicity is the number of atoms in a molecule or compound.
A molecule is the smallest particle of an element or compound that can exist on its own(i.e. in a free and separate state).

A molecule with one atom is referred to as a **monoatomic** molecule e.g. Helium(He); argon(Ar);sodium(Na) etc.
A molecule containing two atoms is referred to as a **diatomic** molecule e.g. Oxygen(O_2); hydrogen(H_2); Chlorine(Cl_2) etc.
A molecule containing three atoms is said to be **triatomic** e.g. Ozone (O_3)
A molecule containing four atoms is said to be **tetraatomic** e.g. Phosphorus (P_4)
A molecule containing many atoms is said to be **polyatomic** e.g. Sulphr (S_8).

Example

Find the atomicity of the following elements in the compound.
a) $4CO_2$ b) $5H_2O$ c) $Zn(NO_3)_2$

Solutions

a) Carbon atoms =4, oxygen atoms ,4x2=8
   Total number of atoms =4+8=12.

b) Hydrogen atoms, 5x2=10atoms, oxygen atoms =5

Total number of atoms =10+5=15 atoms.

c) Zinc atoms=1, nitrogen atoms =2, and oxygen atoms,3x2=6

Total number of atoms= 1+2+6=9 atoms

Exercise

Find the number of atoms of each element present in the following compounds.

a) $MgCO_3$ b) $Pb(OH)_4$ c) $Ca_3(PO_4)_2$ d) $(NH_4)_2CO_3$ e) $2H_2SO_4$

Sample questions on the atomic structure and the periodic table

Atomic structure and periodic table

1. *With an aid of a labeled drawing, describe the structure of an atom showing the distribution of the different particles that make it up.*
2. *What are protons, neutrons and electrons? Give a brief detail of each.*
3. *Differentiate between mass number and atomic number, isotope and isotopy. Give three elements that show isotopy, write down and name the isotopes.*
4. *The figure below shows part of the periodic table. The letters used are not the correct symbols of the elements.*

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

Which of the elements are metals? Suppose element P reacts with element T, write the formula of the compound formed between P and T. Which element in the table is least reactive and explain why. Suggest a compound formed between any two elements shown, which would conduct electricity, give a reason for your answer.

5. *Silver is an element which exists naturally as a mixture of two isotopic forms. A and B represent atoms of these isotopes. They occur in equal numbers. A is* $^{107}_{47}Ag$ *and B is* $^{109}_{47}Ag$*. State the number of (a) protons, (b) electrons and (c) neutrons in atoms A*

*and B. What is the relative atomic mass of the naturally occurring silver (**Ans. 108**(Atomic mass).*

6. *Elements X, Y, and Z have the following atomic number, 11, 15 and 19 respectively. Write the electronic configuration of the elements. State each element is a metal or non metal.*
7. *Metals lose electrons to become stable like the noble gases while non metals gain electrons, in the process they form ions. Write down the electronic configuration of the ions formed by magnesium, chlorine and calcium.*

ACIDS, BASES AND INDICATORS

ACIDS

An acid is a compound which when dissolved in water produces hydrogen ions (H^+) as the only positively charged ion. There are basically two types of acids; mineral and organic acids.

a) Mineral acids

These are called mineral acids because they are derived from minerals, examples include, the common laboratory acids like sulphuric acid (H_2SO_4), hydrochloric acid (HCl), nitric acid (HNO_3) and phosphoric acid(H_3PO_4).

b) Organic acids

These are acids derived from organic compounds and examples include: ethanoic acid (CH_3COOOH) found in vines, citric acid found in fruits and lactic acid found in milk.

Properties of acids

1. Acids have a sharp sour taste.
2. Acids change colors of indicators e.g. turn blue litmus paper red.
3. Acids ionize in water to produce hydrogen ions (H^+).

Sulphuric acid: $H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{2-}(aq)$

Hydrochloric acid: $HCl(aq) \longrightarrow H^+(aq) + Cl^-aq)$

4. Most acids are corrosive (i.e. have burning effects) and poisonous
5. Dilute acids react with reactive metals to produce hydrogen gas and a salt.

Acid + Metal $\longrightarrow$ Salt + Hydrogen gas

For example

$H_2SO_4(aq) + 2Na(s) \longrightarrow Na_2SO_4(aq) + H_2(g)$

6. Acids react with carbonates and hydrogen carbonates to liberate carbondioxide gas. When an acid reacts with a carbonate,a salt, water and carbondioxide gas are produced.

Acid + Carbonate $\longrightarrow$ Salt + Water + Carbondioxide

Example

$HCl(l) + Na_2CO_3(s) \longrightarrow NaCl(aq) + H_2O(l) + CO_2(g)$

$H_2SO_4(aq) + 2NaHCO_3(s) \longrightarrow Na_2SO_4(aq) + 2H_2O(l) + 2CO_2(l)$

7. Acids react with bases to to form salt and water

Acid + Base $\longrightarrow$ Salt + Water

Examples

$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$

$H_2SO_4(aq) + CuO(s) \longrightarrow CuSO_4(aq) + H_2O(l)$

Preparation of acids

1. By the reaction between an acid anhydride (an oxide of a non metal) and water.

Example

Sulphur dioxide + water $\longrightarrow$ Sulphurous acid

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

Phosphorus(V)oxide + Water $\longrightarrow$ Phosphoric acid

$P_4O_{10}(s) + 6H_2O(l) \longrightarrow 4H_3PO_4(aq)$

2. By displacing a weaker acid (more volatile acid) from its salt by a stronger acid (less volatile acid).

Example

$2NaCl(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2HCl(aq)$

3. By precipitating an insoluble sulphide from a metallic salt by hydrogen sulphide.

For example

$Pb(CH_3COO)_2(aq) + H_2S(g) \longrightarrow PbS(s) + 2CH_3COOH(aq)$

Lead ethanoate Ethanoic acid

Basicity of an acid

This is the number of hydrogen ions that can be produced when one molecule of an acid ionizes in water. Acids can be categorized interms of basicity as:

a) Monobasic acid

Is an acid that ionizes to produce one hydrogen ion (H^+). Their basicity is one.

Examples

Nitric acid: $HNO_3(aq) \longrightarrow H^+(aq) + NO_3^-(aq)$

Hydrochloric acd: $HCl(aq) \longrightarrow H^+(aq) + Cl^-aq)$

Ethanoic acid: $CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$

b) Dibasic acids

These acids ionize to produce two hydrogen ions ($2H^+$) from one molecule. They have a basicity of two. Examples include;

Sulphuric acid: $H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{2-}(aq)$

Carbonic acid: $H_2CO_3(aq) \longrightarrow 2H^+(aq) + CO_3^{2-}(aq)$

c) Tribasic acids

These are acids that ionize to produce three hydrogen ions $(3H^+)$. They have basicity of 3. For example:

Phosphoric acid: $H_3PO_4(aq) \longrightarrow 3H^+(aq) + PO_4^{3-}(aq)$

Strength of acids

Strength of an acid refers to the ease with which an acid ionize to produce hydrogen ions. According to strength, acids can be categorized as strong or weak.

1. **Strong acid:** Is an acid that dissociates completely into ions when dissolved in water (it ionizes completely in water). Examples include:

Sulphuric acid: $H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{2-}(aq)$

Hydrochloric acid: $HCl(aq) \longrightarrow H^+(aq) + Cl^-aq)$

2. **Weak acid:** Is an acid that does not completely ionize in water. Some of the acid molecules do not ionize, examples include:

Carbonic acid: $H_2CO_3(aq) \rightleftharpoons 2H^+(aq) + CO_3^{2-}(aq)$

Ethanoic acid: $CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$

Uses of acids

i) Acids are used in car batteries
ii) Sulphuric acid is used in the manufacture of soap, detergents and paints. Plastics e.t.c.
iii) Ethanoic acid is used in the preservation of fruits and vegetables.
iv) Hydrochloric acid in the stomach of human being helps in the digestion of food.
v) Acids are used in the manufacture of fertilizers e.g. sulphuric acid.
vi) Ascorbic acid in fruits are useful to our bodies.
vii) Nitric acids are used in the manufacture of explosives.

Bases and Alkalis

A base is a substance which reacts with an acid to form salt and water only.

Bases are mainly oxides of metals like: copper(II)oxide, CuO; zinc oxide, ZnO;calcium oxide, CaO; magnesium oxide, MgO;or hydroxides of metals and ammonium groups like sodium hydroxide, $NaOH$; potassium hydroxide, KOH; calcium hydroxide,$Ca(OH)_2$ and ammonium hydroxide, NH_4OH.

Bases that are soluble in water are known as alkalis.

Alkalis

An alkali is a base that dissolves in water to produce hydroxide ions (OH^-) as the only negatively charged ion. A solution of a base in water is called an alkaline solution.

Examples of alkalis include: sodium hydroxide or caustic soda($NaOH$), potassium hydroxide or caustic potash (KOH), calcium hydroxide or lime water ($Ca(OH)_2$) and ammonium hydroxide or aqueous ammonia (NH_4OH).

Properties of alkalis

1. Alkalis have a bitter taste
2. Alkalis are soapy and feel slippery
3. Alkalis turn red litmus paper blue
4. Alkalis react with acids to produce salt and water.

   $$\text{Acid + Alkali} \longrightarrow \text{Salt + Water}$$

   $$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

5. Alkalis react with ammonium salts to produce ammonia gas. E.g.

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

6. Alkalis precipiatate insoluble metal hydroxides from solutions of their salts. E.g.

   $$2NaOH(aq) + Pb(NO_3)_2(aq) \longrightarrow Pb(OH)_2(s) + 2NaNO_3(aq)$$

Strength of bases/ alkalis

This is the ease with which a base/ alkali dissociate or ionize. The bases/ alkalis can be categorized as strong or weak.

1. **Strong alkali:** Is one that ionizes completely in aqueous solution. Examples include:

Sodium hydroxide: $NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$

Potassium hydroxide: $KOH(aq) \longrightarrow K^+(aq) + OH^-(aq)$

2. **Weak alkali:** This is one that does not ionize completely in aqueous solution. Some molecules remain unionized in the solution. For example:

Ammonium hydroxide: $NH_4OH(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

Neutralization reaction

This is the reaction between an acid and a base to produce salt and water only. The resulting solution is neutral (i.e neither acidic nor alkaline). Examples of neutralization reactions are:

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

$$H_2SO_4(aq) + KOH(aq) \longrightarrow K_2SO_4(aq) + H_2O(l)$$

Essentially neutralization takes place when a hydroxide ion (OH^-) reacts with a hydrogen ion (H^+) to form a water molecule (H_2O).

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

INDICATORS (ACID-BASE INDICATORS)

An acid-base indicator is a substance that shows different colors in acidic and alkaline solutions. Indicators are used to establish substances that are acidic, alkaline or neutral. The color change of the indicator depends on the strength of the acid or the base/ alkaline.

Some examples of indicators used in chemistry experiments include: Litmus paper; Methyl orange; Bromothymol blue; Bromothymol red; Phenolphthalein; Universal indicators e.t.c.

Color changes of some indicators in acidic, alkaline and neutral solutions are shown below

Plant extract as a simple acid-base indicators

Preparation of indicators from flowers

Procedure

Collect a handful of flowers with brightly colored petals such as hibiscus and morning glory.

NB Use only one type of flowers and do not mix them

- Remove the petals from the flowers, put them in a motor and crush them carefully
- Add a little ethanol to the crushed petals and stir with a pestle
- Carefully decant the mixture and keep the colored solution in a test tube. This acts as an indicator. Take note of the color of the indicator

Universal indicator

A universal indicator is a mixture of other simple indicators and under goes a range of color changes in solution of different P^H. A universal indicator shows the extent of alkalinity or acidity (i.e in addition to showing whether a solution is acidic or alkaline, it shows whether the acid is weak or strong.) The indicator may be available in solution or paper form.

Indicator	**Color in acidic solution**	**Color in neutral solution**	**Color in alkaline solution**
Litmus	Red	Purple	Blue
Methyl orange	Pink	Orange	Yellow
Phenolphthalein	Colorless	Colorless	Pink

P^H and P^H Scale

The PH of a solution is a number which shows the acidic or basic strength of the solution.

It can also be defined as the negative logarithm (to base ten) of hydrogen ion concentration. It is obtained by finding the negative logarithm of hydrogen ion concentration.

A PH scale measures the strength of an acid or a base/alkali. The PH scale runs from Zero (O) to fourteen (14).

The following are values on the PH scale for solutions of different PH.

- An acidic solution has a PH less than 7
- A neutral solutions has a PH of 7
- An alkaline solution has a PH greater than 7

Representation on a PH scale

PH 0 1 2 3	4 5 6	7	8 9 10	11 12 13 14
Strong acid	Weak acid	Neutral	Weak alkali	Strong alkali

← Increasing acidity

Increasing alkalinity →

Color changes on a universal indicator PH scale

PH 0 1 2 3	4 5	6	7	8 9	10 11	12 13 14
Red	Orange	Yellow	Green	Blue	Indigo	Violet

To find the PH of a solution using a solution of a universal indicator.

1. Add some drops of the universal indicator to the solution being tested.
2. Match the color change of the solution to the corresponding color on the universal indicator PH chart.
3. Record the color change and the PH value.

CHEMICAL FAMILIES

Elements in the same group are referred to as chemical families because of the similarities in their chemical and physical properties. These elements in the same group have similar chemical properties because they have the same number of electrons in their outer most energy levels. The common chemical families are; alkaline metals; alkaline earth metals; halogens and noble gases.

Alkaline metals

These are elements of group I in the periodic table. These metals are very reactive and are kept under oil in the laboratory where they have no contact with water and air. They include lithium (Li), sodium (Na) and potassium (K). Atoms of all these elements each have a single electron on the outer most shell, which is easily lost during chemical reactions leaving a single positively charged ion. i.e.

$Li \longrightarrow Li^+ + e^-$

$Na \longrightarrow Na^+ + e^-$

$K \longrightarrow K^+ + e^-$

Physical properties

1. The metals are soft and can be cut with a knife. The softness decreases down the group.
2. Their melting and boiling points are low. The melting point decreases down the group.`
3. They have low densities and can float on water.
4. They have shinny surfaces when freshly cut. The shinny surfaces soon tarnish due to reaction with oxygen to form the oxides.
5. They are good conductors of electricity and heat.
6. They do not posses tensile strength to any appreciable extent.

Chemical properties

Reactivity of the alkaline metals increases down the group due to increase in atomic size.(i.e potassium is the most reactive followed by sodium and lithium is the least reactive).

1. Reaction with water

All the metals react with cold water to produce hydrogen gas and alkaline solution of metal hydroxides.

Potassium reacts very vigorously, darts on water and burns with a bright purple (lilac) flame producing fumes of hydrogen gas and an alkaline solution of potassium hydroxide.

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

Sodium melts into a silvery ball, darts on water reacting vigorously with it producing fumes of hydrogen gas with a hissing sound and an alkaline solution of sodium hydroxide. The reducing in size as the reaction proceeds until when it finally disappears.

$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$

Lithium reacts slowly with cold water liberating hydrogen gas.
$2Li(s) + 2H_2O(l) \longrightarrow 2LiOH(aq) + H_2(g)$

2. Reaction with air

When the metals are heated in oxygen, they burn readily forming oxides of the metal. Potassium burns with a bright purple flame producing white solids of potassium oxide.

$4K(s) + O_2(g) \longrightarrow 2K_2O(s)$

Sodium burns in air with a bright yellow flame forming yellow solids of sodium peroxide in plenty of air. In limited in air, sodium forms white solids of sodium oxide.

$4Na(s) + 2O_2(s) \longrightarrow 2Na_2O_2(s)$ (In excess air)
$4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$ (In limited air)

Lithium burns with a red flame forming white residue of lithium oxide.
$4Li(s) + O_2(g) \longrightarrow 2Li_2O(s)$

3. Reaction with chlorine

The alkaline metals react with chlorine to form metal chloride salts. Eg. Sodium continuous to burn in chlorine to form white fumes of sodium chloride.

$Na(s) + Cl_2(g) \longrightarrow NaCl(s)$
Potassium burns in chlorine to form fumes of potassium chloride.

$K(s) + Cl_2(g) \longrightarrow KCl(s)$

4. Reaction with acids

Alkaline metals react explosively with acids to liberate hydrogen gas and forming a corresponding salt. E.g. sodium reacts with hydrochloric acid to form sodium chloride and hydrogen gas.

$2Na(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2(g)$

Potassium reacts with sulphuric acid to form potassium sulphate and hydrogen gas.

$2K(s) + H_2SO_4(a) \longrightarrow K_2SO_4(aq) + H_2(g)$

Alkaline earth metals

These are elements in group II of the periodic table. They include beryllium, magnesium and calcim. These metals have two electrons on their outer most shell, they lose the two electrons during reactions to form a dipositive ion (an ion with a charge of $^{+}2$).

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

$Ca \longrightarrow Ca^{2+} + 2e^-$

$Be \longrightarrow Be^{2+} + 2e^-$

Physical properties

1. They re harder than group I metals.
2. They are good conductors of electricity and heat.
3. They are silvery grey in color when freshly cut. However the shinny surface slowly tarnishes on exposure to air forming the metal oxide.
4. Their melting and boiling points are higher than those of group I. This is because these metals release more electrons into the electron cloud forming a stronger metallic bond in the metal structure. This explains why they are also harder than group I metals.

Chemical reactions

Alkaline earth metals are more reactive than alkaline metals. This is because the valence electrons in alkaline earth metals are held more strongly due to increased effective nuclear charge making them not easily to be released. Reactivity of the metals increase down the group due to increase in atomic size.(calcium is the most reactive followed by magnesium then beryllium)

1. Reaction with air

The metals burn in air with their characteristic flame forming oxides of the metal. Magnesium burns in air with a brilliant white flame forming white ash of magnesium oxide.

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

Calcium burns in air with a bright orange flame forming white ash of calcium oxide.

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

Both metals also form some nitrides especially if the supply of oxygen is limited.

$3Mg(s) + N_2(g) \longrightarrow Mg_3N_2(s)$ (Magnesium nitride)

$3Ca(s) + N_2(g) \longrightarrow Ca_3N_2(s)$ (Calcium nitride)

When the nitrides are dissolved in water and the solution warmed, ammonia gas is liberated and an alkaline solution is also formed.

$Mg_3N_2(s) + 6H_2O(l) \longrightarrow 3Mg(OH)_2(aq) + 2NH_3(g)$

$Ca_3N_2(s) + 6H_2O(l) \longrightarrow 3Ca(OH)_2(aq) + 2NH_3(g)$

2. Reaction with water

Beryllium does not react with water. Magnesium piece reacts very slowly with cold water producing small bubbles of hydrogen gas and an alkaline solution of magnesium hydroxide.

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

Magnesium reacts very rapidly with steam to produce white ash of magnesium oxide and hydrogen gas.

$2Mg(s) + 2H_2O(g) \longrightarrow 2MgO(s) + 2H_2(g)$

Calcium reacts steadily with water evolving hydrogen gas and calcium hydroxide.

$Ca(s) + 2H_2O(l) \longrightarrow Ca(OH)_2(aq) + H_2(g)$

3. Reaction with chlorine

The metals burn readily in chlorine to form white anhydrous chloride salts. Magnesium forms magnesium chloride and calcium forms calcium chloride.

$Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)$

$Ca(s) + Cl_2(g) \longrightarrow CaCl_2(s)$

4. Reaction with acids

Calcium reacts vigorously with acids forming a salt and hydrogen gas. Magnesium reacts steadily with acids forming a salt and hydrogen gas. E.g. magnesium reacts with sulphuric acid forming magnesium sulphate and hydrogen gas.

$Mg(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2(g)$

Halogens

These are group (VII) elements; they are chlorine, bromine, iodine and fluorine. Halogens exist as diatomic molecules. This is because they lack one electron to completely fill their outer most energy level hence their atoms share electrons in order to completely fill the outer most energy level. These elements react by gaining one electron each forming negatively charged ions.

$F + e^- \longrightarrow F^-$

$Cl + e^- \longrightarrow Cl^-$

$Br + e^- \longrightarrow Br^-$

$I + e^- \longrightarrow I^-$

Physical properties

1. They are colored (chlorine is greenish yellow gas; fluorine is pale yellow gas; bromine is a brown volatile liquid and iodine is a shinny black solid).

2. Down the group, physical state changes from gas to liquid to solid. (fluorine and chlorine are gases at room temperature; bromine is a liquid and iodine is in solid form, the shinny black iodine solids sublime into a purple vapor on slight heating).
3. Solubility in water decreases down the group. Fluorine, chlorine and bromine are soluble and iodine is slightly soluble in water.
4. They have low melting and boiling points because the intermolecular forces of attraction are weak. The melting and boiling point increase down the group as the weak vanderwaal's forces of attraction increase with increase in atomic size.

Chemical properties

The order of reactivity decreases down the group.(i.e. fluorine is the most reactive and iodine in the least reactive). This is because these elements react by gaining electrons and down the group, the atomic radius increases there fore, an in coming electron is less readily attracted by the nucleus. There fore, the smaller the atom, the more readily and strongly does the nucleus attract an electron to the outer most shell and the more reactive that atom becomes.

1. Reaction with water

All the halogens dissolve in water forming a mixture of acids. Chlorine dissolves in water forming a mixture of hypochlorous acid(chloric (I) acid)(HOCl) and hydrochoric acid(HCl).

$$Cl_2(g) + H_2O(l) \longrightarrow HOCl(aq) + HCl(aq)$$

This solution turns blue litmus paper red then bleaches it. The bleaching action is due to the presence of hypochlorous acid(HOCl) that readily gives up its oxygen to the dye. A solution of bromine in water is weakly acidic and weakly bleaching due to formation of weak hydrobromic acid (HBr) and bromic(I)acid(HOBr) respectively.

$$Br_2(g) + H_2O(l) \longrightarrow HOBr(aq) + HBr(aq)$$

Chlorine is slightly soluble in water forming very weak acids. A solution of chlorine is too weak to bleach dyes, therefore, bleaching power of the elements decrease down the group.

2. Reaction with alkalis

Halogen react with alkalis to form a mixture of salts and water. E.g. chlorine is absorbed by a solution of sodium hydroxide forming a pale yellow solution of a mixture of sodium chloride(NaCl) and sodium hypochlorite(NaOCl)

$Cl_2(g) + 2NaOH(aq) \longrightarrow NaCl(aq) + NaOCl(aq) + H_2O(l)$

In hot concentrated solution of an alkali (e.g. NaOH), sodium hypochlorite is not formed but instead sodium hypochlorate is formed($NaOCl_3$).

$3Cl_2(g) + 6NaOH(aq) \longrightarrow 5NaCl(aq) + NaOCl_3(aq) + 3H_2O(l)$

3. Reaction with metals

Metals continue to burn in halogens forming salts. E.g. sodium continues to burn in a gas jar of chlorine to form dense white fumes which settle as white solids of sodium chloride. A similar reaction occurs with bromine and iodine.

$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$

$2Na(s) + Br_2(g) \longrightarrow 2NaBr(s)$

4. Displacement reaction of halogens

When chlorine is bubbled through a colorless solution of potassium bromide, the solution gradually change to red-brown as bromine is displaced from its solution by chlorine.

$2KBr(aq) + Cl_2(g) \longrightarrow 2KCl(aq) + Br_2(g)$

If chlorine is bubbled through a colorless solution of potassium iodide,the solution turns dark brown as iodine is displaced from its solution.

$2KI(aq) + Cl_2(g) \longrightarrow 2KCl(aq) + I_2(g)$

Chlorine being the most reactive displaces all the halogens from their solutions.

Noble Gases (Group (VIII) Elements)

These elements are referred to as inert elements because they are non reactive. The elements include; helium,neon,argon,krypton,xenon and radon. Characteristics of these elements include:

1. Their outer most energy levels are completely filled with electrons making them very stable elements and chemically non reactive. However in spite of their stable electronic configurations, a few compounds have been formed from xenon(e.g. Xenon(II)fluoride,XeF_2) and krypton(E.g. krypton(II)fluoride, KrF_2)
2. The noble gases exist as monatomic gases.
3. They have very low melting and boiling points due to negligible forces of attraction between the atoms.
4. Noble gases are colorless

Though nobles gases are non reactive, they are used for a variety of purposes including the following:

- Argon is filled in light bulbs to prevent the tungsten fillament from reacting with air.
- Argon provides an inert environment for gas-liquid chromatography, for risky welding jobs and for some chemical reactions.
- Neon is used extensively in advertising signs and in lasers.
- Helium is used to inflate air ships, weather balloons and aeroplane tyres.
- A mixture of oxygen and helium is used by divers instead of air.

Sample questions on Acids, Bases, Indicators and Chemical Families

Acids, bases and indicators

1. *Write short notes on: acids, bases, alkalis, indicators and basicity. Illustrate with examples and equations.*

2. *Outline three properties typical of acids and three properties typical of alkalis. (illustrate with equations). Differentiate between: a strong acid and a weak acid; an alkali and a base.*
3. *Describe how you can use a named material from the environment to make an indicator in the laboratory.*
4. *What is meant by neutralization reaction? Illustrate by use of equations.*

Chemical families

1. *Describe with equations, the reactions of*

a) *Alkaline metals with air, water, acids and chlorine.*
b) *Halogen with water and alkalis.*
c) *Alkaline earth metals with water.*

2. *Discuss the displacement reaction of halogens.*
3. *Mention three examples of noble gases. Outline four uses of noble gases.*

WATER AND HYDROGEN

WATER

Water is a chemical compound consisting of oxygen and hydrogen. Its chemical formula is H_2O.

Sources of water

1. **Rain water**

The water originates from rain and it is relatively pure with no dissolved minerals. However, it contains suspended materials and dissolved gases such as carbon dioxide and sulphur dioxide. Rain water is good for washing as it forms lather easily with soap and does not contain dissolved salts. It is also good for drinking since it has a flat taste.

2. **Wells and spring water**

The water comes from underground and is obtained by drilling the ground inform of a bore hole or, the water just springs out of the ground. It is actually rain water that has drained into the ground and collected into a pool of underground water. It has less suspended materials but contain dissolved mineral salts making it good for drinking. It may not be good for washing because of the dissolved mineral salts that delay the formation of lather.

3. **River and lake water**

The water is obtained from rivers and lakes. It contains both suspended materials and dissolved minerals. As water flows in to rivers and lakes, it gathers floating materials and pollutants such as sewage and industrial wastes, so it is not good for washing and drinking.

4. **Sea and ocean water**

These provide large sources of water. The water contains a lot of suspended materials and dissolved substances. The mineral salts accumulate settle at the bottom of the water and the water may become saturated with the salt. This is why the water tastes salty.

5. **Ice caps**

The water is obtained from melting ice and it is relatively pure. The ice formed at the peaks of mountains melts and the water runs down forming spings.

Uses of water

1. **Biological use**

All biological processes in living organisms require water as a medium for the reactions. For example, photosynthesis, excretion, transport of materials and transmission of impulses.

2. Domestic use

Drinking, cooking, washing, bathing, mopping, construction, recreation e.t.c.

3. Farm use

Irrigation, for dipping animals, for drinking by animals, for fish farming e.t.c.

4. Industrial use

Water is used as a solvent in many industries, for cleaning, for cooling purposes especially in power stations, used in steam engines, generating hydro electricity, transportation using boats, ships e.t.c, for recreation purposes and construction.

Water cycle

Water cycle is the process by which water circulates round the earth. The driving force for the cycle is the sun.

The water from the water bodies like seas, lakes, rivers e.t.c. evaporates into the atmosphere due to the sun heat. The vapor cools in the atmosphere and then condenses to form clouds which fall as rain. The rain then flows back into the water bodies.

N.B. Water vapor in the atmosphere may also result from transpiration processes in plants and respiration processes of all living organisms and burning of starchy food materials forming carbon dioxide and water which are expelled in to the atmosphere.

Water pollution

This is the discharge of unwanted materials/ substances into water bodies. The un wanted materials are generally called pollutants and are harmful to living organisms. Examples include; acids, bases, insecticides, fertilizers, sewage, solid particles, oil e.t.c.

Sources of water pollutants

1. Atmospheric gases like carbondioxide, sulphur dioxide, nitrogen dioxide and hydrogen sulphide. These gases dissolve in the rain water making it acidic.
2. Dissolved maerials from rocks and soil.
3. Industrial discharge (effluent)
4. Fertilizers washed from farm lands by running water.

5. Insecticides and herbicides resulting from spraying from homes and farms.
6. Soaps, detergents and other chemicals from homes, salons e.t.c.

Water treatment

This is the removal of harmful substances from water making it suitable for use (domestic use. In Uganda, the treatment of water is done by Uganda National Water and Sewage Co operation (UNWSC).

Processes involved in water treatment are:

1. Addition of chemicals

Chemicals are added to water to remove harmful substances and neutralize acids and bases. The chemicals as well remove hardness of water, precipitate metals and cause suspended solid particles to settle.

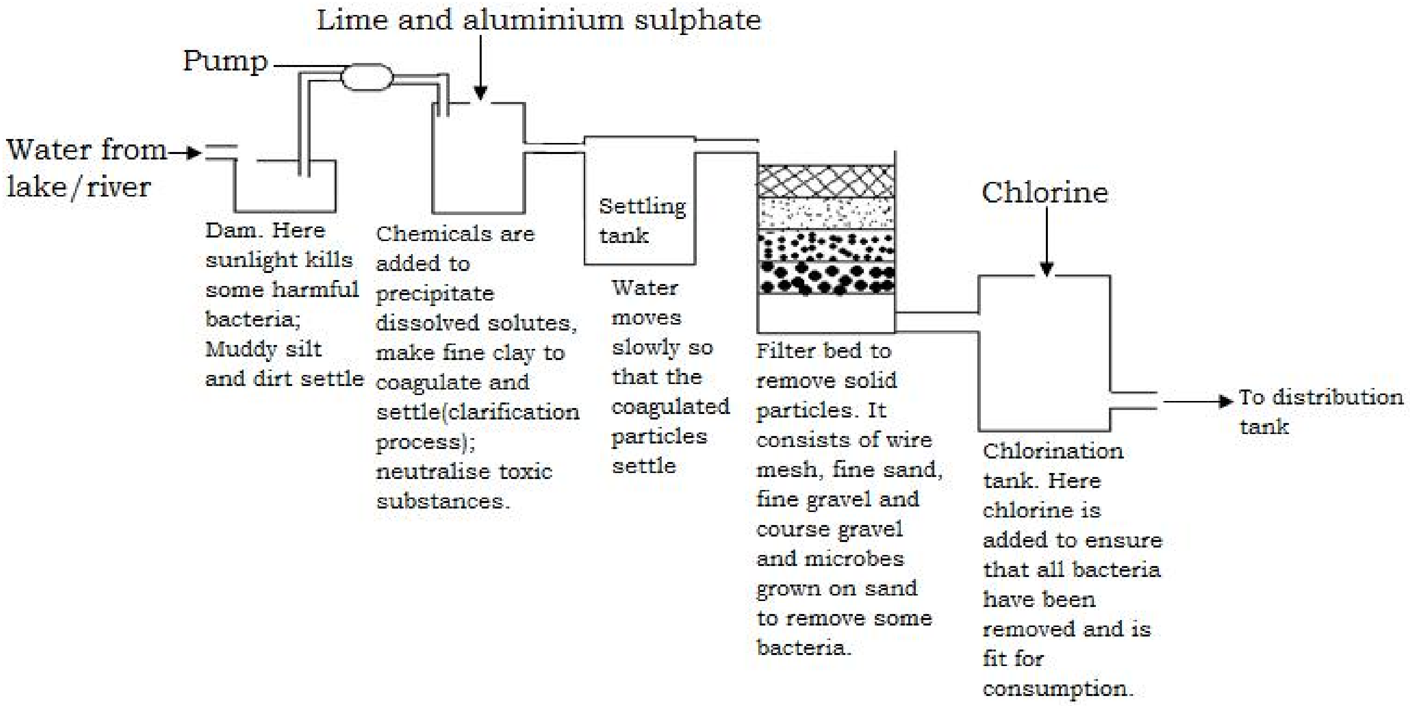
2. Filtration

Filtration is carried out to remove the solid particles. Water is made to pass through wire mesh screens to remove large and floating objects and then through sand and gravel to filter out smaller solid particles. The last part of the filter bed consists of sand with some microbes grown to remove some bacteria that might be harmful.

3. Chlorination and fluoridation

Chlorine or fluorine is added to kill the harmful organisms such as bacteria. Only a small concentration of chlorine is used because it is harmful. However, in swimming pools, high concentration of chlorine is used as the water may contain high bacteria concentration is not meant for drinking. In some countries, ozone is used instead of chlorine.

Summary of water purification process



Sewage

Sewage is running water containing wastes from toilets, bathrooms, sinks, factories and streets of towns.

Sewage treatment

The reason for treating sewage is to remove harmful pollutants before allowing the sewage to join water bodies. The following are the processes involved in sewage treatment:

1. Separation of solid and liquid wastes through sieving and sedimentation.
2. Addition of certain bacteria to the liquid wastes to convert harmful materials to harmless substances.
3. Addition of chlorine to kill the harmful bacteria
4. The liquid waste (effluent) from the treatment plant is the discharged into the water body. The solid waste from the treatment of sewage is the called the **sludge** and has a number of uses.

Uses of sludge

- It is used as fertilizer because it is rich in nitrogen and phosphorus.
- It is used as a raw material to produce bio gas.

- It is used for road surfacing since it forms hard solids.
- It is used in land reclamation to fill quarries.

Properties of water

Physical properties

1. Pure water is a clear, colorless and tasteless liquid.
2. It boils at 100°C and melts at 0°C.
3. It has a density of $1g/cm^3$ at 4°C.
4. Pure water is neutral to litmus i.e. has no effect on litmus.
5. Water expands on freezing.
6. It is a universal solvent.

Test for water

Water is chemically tested for using anhydrous copper (II) sulphate of cobalt chloride paper.

When water is added to anhydrous copper(II)sulphate, it changes color from white to blue. Or when water is added to cobalt chloride paper, it changes color from blue to pink.

Chemical properties

1. Reaction with metals

Water reacts with some metals and the vigour depends on the position of the metal in the electrochemical series.

Potassium }
Sodium } reacts with cold water
Calcium
Magnesium }
Aluminium } reacts with steam
Zinc }
Iron
Tin }
Lead }
(Hydrogen) }

Copper do not react with water in any form
Silver
Mercury
Platinum
Gold

a) Potassium

When a piece of potassium is placed into a trough containing water, it reacts very vigorously, catches fire and burns with a purple flame; it darts on the water surface producing a hissing sound and white fumes of hydrogen gas. It explodes with a pop sound as it disappears. The resultant solution turns red litmus paper to blue.

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

b) Sodium

Sodium reacts vigorously with water. When a piece of sodium is placed in cold water, it melts into a silvery ball and darts on the surface of water as it produces a hissing sound with evolution of a colorless hydrogen gas. The sodium gradually becomes smaller and smaller as it reacts with water to produce an alkaline solution of sodium hydroxide.

N.B. Sodium burns with a yellow flame if its movement on water is restricted.

$$Na(s) + H_2O(l) \longrightarrow NaOH(l) + H_2(g)$$

c) Calcium

Calcium reacts steadily with water. When a small piece of calcium is dropped is water, it sinks to the bottom producing streams of bubbles of a colorless gas (hydrogen gas) and an alkaline solution that turns red litmus paper blue. The water turns milky as the calcium hydroxide is slightly soluble in water.

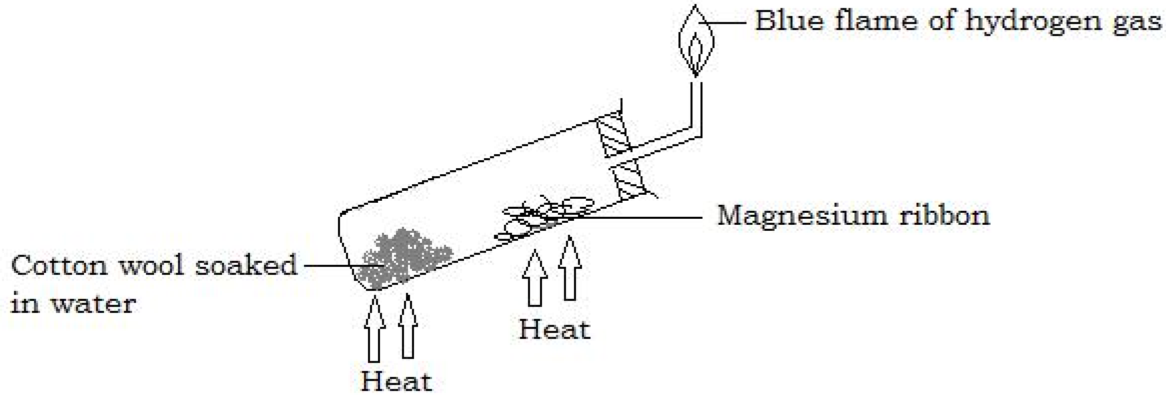
$$Ca(a) + H_2O(l) \longrightarrow 2Ca(OH)_2(aq) + H_2(g)$$

d) Magnesium

Magnesium sinks to the bottom and reacts very slowly with cold water to produce very few bubbles of hydrogen gas on the surface of the metal and an alkaline solution of magnesium hydroxide.

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

However, hot magnesium reacts rapidly with steam, burning with a bright white flame leaving white ash of magnesium oxide. The hydrogen gas produced burns with a blue flame in air.



$Mg(s) + H_2O(g) \longrightarrow MgO(s) + H_2(g)$

Note

1. The magnesium metal used is first cleaned to remove the oxide coating due to corrosion.
2. The cotton wool soaked in water is heated to release steam that reacts with the heated magnesium.
3. The boiling tube is first heated to drive out all air (oxygen).

e) Iron and zinc

Zinc and iron do not react with cold water but react with steam when heated to red hot producing their respective oxides and hydrogen gas. Zinc reacts faster than iron.

$Zn(s) + H_2O(g) \longrightarrow \underset{\text{Yellow when hot}}{ZnO(s)} + H_2(g)$

$$3Fe(s) + 4H_2O(g) \xrightarrow{\text{White when cold}} \underset{\substack{\text{Triiron tetra oxide}\\ \text{(Blue-black)}}}{Fe_3O_4(s)} + 4H_2(g)$$

Aluminium does not react with water or steam due to the formation of an insoluble oxide layer which prevents contact between water and the metal inhibiting any further reaction.

All the other metals below iron do not react with either cold water or even steam.

2. **Reaction with oxides**

Water reacts with metallic oxides (basic oxides) to form alkaline solutions. E.g.

$$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$$

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

With non metallic oxides (acidic oxides), water forms acidic solutions. E.g.

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

$$SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq)$$

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

HARDNESS OF WATER

Water can be classified as hard or soft.

Soft water is one which forms lather readily with soap. Examples of soft water include: rain water, distilled water, deionised water.

Hard water

This is water that does not form lather readily with soap. Examples of hard water include: tap water, sea water, borehole, well water, spring water.

Causes of hardness of water

Hardness of water is caused by the presence of the calcium ion (Ca^+) and magnesium ion (Mg^+). These ions reach water when the soluble salts of calcium and magnesium are dissolved by rain water and carried to water bodies. The compounds that cause hardness of water include; calcium hydrogen carbonate ($Ca(HCO_3)_2$), magnesium hydrogen carbonate($Mg(HCO_3)_2$),calcium chloride ($CaCl_2$), magnesium chloride ($MgCl_2$), Calcium sulphate ($CaSO_4$), Magnesium sulphate ($MgSO_4$).

Effects of magnesium and calcium ions on soap

Soap is a sodium salt of a long chain carboxylic acid. It is known as sodium stearate (Na-st). Soap reacts with magnesium or calcium ions to form dirty white solids called **scum**. So, when hard water is used for washing, a lot of soap is used because part of the soap initially reacts with calcium and magnesium ions to form scum.

$$CaSO_4(aq) + 2Na\text{-}st(aq) \longrightarrow Na_2SO_4(aq) + Ca\text{-}st_2(s)$$

Types of hard water

There are two types of hard water i.e. permanent and temporary hard water.

Temporary hard water

This is a type of hard water whose hardness can be removed by boiling. I.e. the water can be made soft by boiling.

Causes: it is caused by the presence of dissolved calcium hydrogen carbonate or magnesium hydrogen carbonate.

Permanent hardness

Is a type of hard water whose hardness cannot be removed by boiling. I.e. the hard water cannot become soft on boiling.

Causes: it is caused by calcium sulphate, calcium chloride, magnesium sulphate and magnesium chloride.

Distinguishing between permanent hard water, temporary hard water and soft water

Procedure

- Place equal volumes of each unboiled water samples in three different conical flasks.
- Run soap solution from the burette into each sample while shaking the conical flask until when lather forms.
- Note the volume of soap solution required to form lather with each of the unboiled water samples.

Results

The soft water required the least volume of soap solution to form lather compared to the samples of hard water.

Repeat the above procedure with boiled water sample.

Results

Soft water required the same volume of soap solution to form lather as the unboiled soft water.

Temporary hard water after boiling required less volume of soap solution to form lather than the unboiled one. This is because; the water was made soft by boiling and thus formed lather readily.

Permanent hard water even after boiling required the same volume of soap solution as the unboiled to form lather. This is because even after boiling, the water still remained hard and took time to form lather.

Methods of softening water

Removal of only temporary hardness of water

1. **Boiling** (Physical method)

Boiling decomposes calcium and magnesium hydrogen carbonates to their respective carbonates, water and carbondioxide gas. The calcium and magnesium ions are removed from the water as insoluble calcium and magnesium carbonates. The calcium and magnesium carbonates may be deposited as solids at the bottom of the boiler and are referred to as boiler scales of kettle fur.

$$Ca(HCO_3)_2(aq) \xrightarrow{\text{(boiling)}} CaCO_3(s) + CO_2(g) + H_2O(l)$$

Disadvantages of boiler scale/fur

- The fur produced is an insulator hence a lot of heat is required for boiling and this wastes a lot of fuel or electricity.
- The fur produced may also block pipes carrying boiled water and possibly causing it to burst.

By chemical means the fur can be removed from he kettle by adding mineral acids like hydrochloric acid and sulphuric acid.

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

2. **Addition of calcium hydroxide (lime water)** (chemical method)

Calcium hydroxide reacts with the soluble magnesium or calcium hydrogen carbonates to form insoluble magnesium and calcium carbonates.

$$Ca(HCO_3)_2(aq) + Ca(OH)_2(aq) \longrightarrow 2CaCO_3(s) + 2H_2O(l)$$

The disadvantage of this method is that when excess lime water is used, the water becomes hard again.

It has an advantage in that it is cheaper than boiling as calcium hydroxide is readily available.

3. **Addition of aqueous ammonia (ammonium hydroxide solution)** (chemical method)

During the process, the soluble hydrogen carbonates are converted into insoluble carbonates. This enables magnesium and calcium ions to be removed from the water.

$2NH_4OH(aq) + Mg(HCO_3)_2(aq) \longrightarrow MgCO_3(s) + (NH_4)_2CO_3(aq) + H_2O(aq)$

Methods for softening both temporary and permanent hardness

4. **Distillation** (physical method)

During distillation, the soft water is collected as the distillate and the dissolved calcium and magnesium ions remain in the distillation flask as residue.

The advantage of this method is that the water obtained is actually pure. However, it can only be used in small scale because it is expensive.

5. **Addition of sodium carbonate solution (washing soda)** (chemical method)

Sodium carbonate converts magnesium or calcium ions into insoluble carbonates.

$Na_2CO_3(aq) + Ca(HCO_3)_2(aq) \longrightarrow 2NaHCO_3(aq) + CaCO_3(s)$

The method is cheap and easy to carryout on a large scale.

6. **Ion exchange method (permutit)** (chemical method)

The calcium and magnesium ions in hard water can be exchanged with sodium ions using a suitable ion exchange material. This makes the water to become soft. The common ion exchange materials are zeolites and permutit. Zeolites are natural occurring forms of sodium aluminium silicate and permutit is the artificial form. Zeolite/permutit can be represented by the formula Na_2Y.

$CaCl_2(aq) + Na_2Y(aq) \longrightarrow CaY(s) + 2NaCl(aq)$

$MgSO_4(aq) + Na_2Y(aq) \longrightarrow MgY(s) + 2Na_2SO_4(aq)$

$Ca(HCO_3)_2(aq) + Na_2Y(aq) \longrightarrow CaY(s) + NaHCO_3(aq)$

The method works on the principle of exchange of ions. During the process, calcium and magnesium ions are removed from water and they appear as precipitates of aluminium silicate.

The advantage here is that, the permutit can be recovered and used several times.

Advantages of hard water

1. Calcium compounds in water are important in our diet for strong teeth and bones.
2. Magnesium in hard water is important for photosynthesis since it is used for the formation of chlorophyll.
3. Calcium in hard water is responsible for the formation of strong shells of many animals e.g. snails and egg shells.
4. Hard water tastes better and it is used in the brewing industry.
5. Hard water reduces lead poisoning and it is the only type of water transported by lead pipes.

Disadvantages of hard water

1. It wastes a lot f soap, as more of the soap reacts the calcium and magnesium ions to form scum. There for lather does not form readily.
2. When hard water ids used for washing, it leaves dirty marks on cloth due to formation of scum.
3. Hard water forms kettle fur and boiler scales which are insulators, there fore a lot of fuel or electricity is used during boiling of water.
4. The boiler scales may block the pipes carrying boiling water hence causing the pipe to burst.

Exercise

1. Explain why soap does not form lather readily at first with hard water and then eventually forms the lather?
2. Explain what is meant by hardness of water. What are the causes of temporary and permanent hardness of water?. How do these compounds that cause hardness of water reach the water?
3. A white precipitate is formed when

i) Carbondioxide is blown into lime water

ii) Temporary hard water is boiled

iii) Washing soda is added to water containing dissolved magnesium chloride.

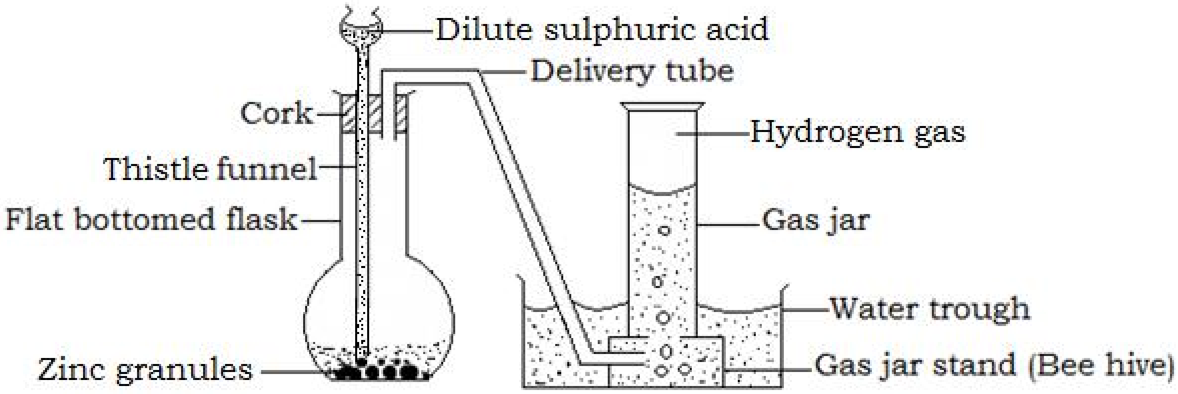
HYDROGEN

Hydrogen is the smallest element and the lightest gas. Hydrogen usually does not occur is Free State but in combined states as water, acid, hydrocarbons and other organic compounds.

Laboratory preparation of hydrogen

Hydrogen is prepared in the laboratory by the action of either dilute hydrochloric acid or dilute sulphuric acid on zinc granules or zinc metal. (zinc granule are used instead of pure zinc because some impurities in the zinc granules act as catalysts).

Set up



Place some zinc granules in the flask and add to it a little copper (II) sulphate solution. The copper (II) sulphate acts as a catalyst in the preparation of hydrogen gas.
Arrange the apparatus as shown above and add dilute sulphuric acid/ hydrochloric acid to the zinc granules through the funnel.

Effervescence occurs as hydrogen gas is produced. The gas is then collected over water. However, if the gas is required dry, it is passed through a wash bottle containing concentrated sulphuric acid and collected by upward delivery method or the gas is passed through a U-tube containing fused calcium chloride to dry the gas.

Equation

$Zn(s) + H_2SO_4(aq) \xrightarrow{\qquad} ZnSO_4(aq) + H_2(g)$

Test for hydrogen

When a burning splint is brought into a gas jar of hydrogen, the splint will be extinguished with a "pop" sound.

Properties of hydrogen

a) Physical properties

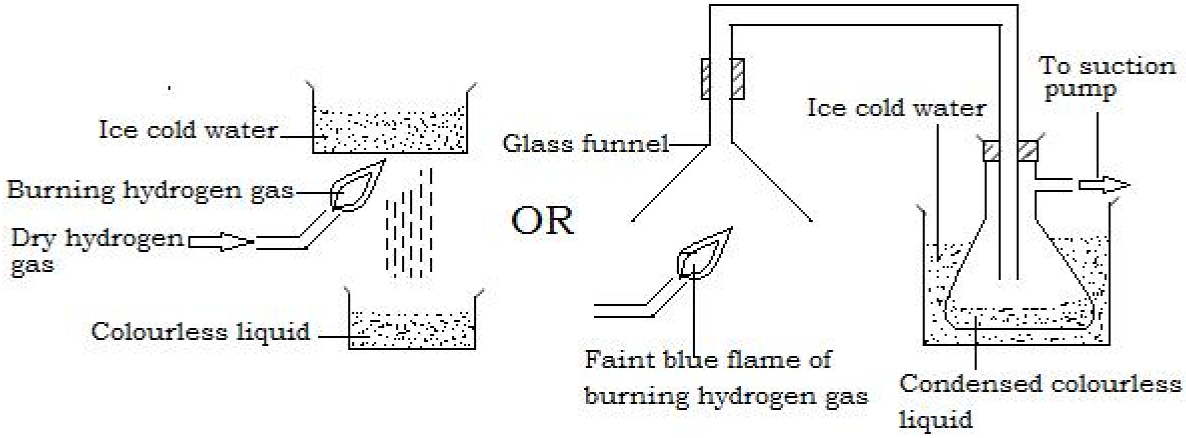
- It is the lightest gas known (lighter than air)
- It is colorless, odourless and tasteless
- It is slightly soluble in water
- It is a neutral gas (has no effects on indicators)

b) Chemical properties

1. Combustion (burning) of hydrogen

Hydrogen burns in air with a faint blue flame to produce water vapor. The gaseous product can be condensed in a cool environment to form a colorless liquid the turns white anhydrous copper (II) sulphate to blue indicating that it is water.

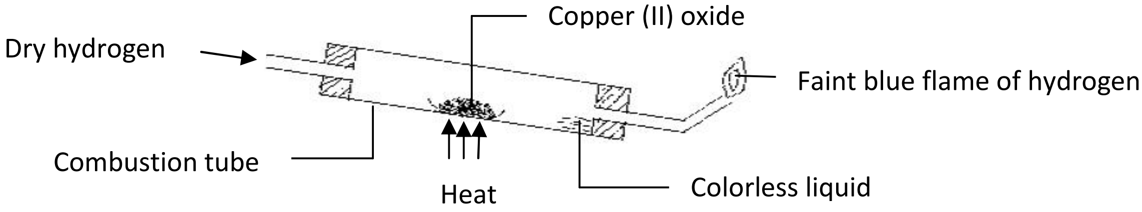
$2H_2(g) + O_2(g) \xrightarrow{\qquad} 2H_2O(g)$



2. Reduction action of hydrogen

Hydrogen is a reducing agent. It removes oxygen from the oxides of some metals (less reactive metals) like lead and copper forming the metal and the hydrogen gas itself is oxidized to water.

Reduction of copper(II)oxide



When dry hydrogen gas is passed over heated copper (II) oxide in a combustion tube, a red glow spreads through the copper (II) oxide and the oxide turns from black to brown as the oxide is reduced to copper metal (brown in color). The hydrogen itself is oxidized to water which collects as a colorless liquid.

$$CuO(s) + H_2(g) \xrightarrow{\qquad\qquad} Cu(s) + H_2O(l)$$

Hydrogen also reduces lead (II) oxide and triirontetraoxide to lead and iron respectively.

Precautions

1. Dry hydrogen gas must be used.
2. The combustion tube must be in a slanting position so as to prevent the water formed from running back to the hot part of the tube.
3. Hydrogen should be passed through the combustion tube to expel out air before heating the oxide.
4. Hydrogen should be passed through the tube for some times after heating .This prevents re oxidation of the copper.
5. The excess hydrogen should be burnt to prevent explosion with air.

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

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

3. Reaction with chlorine

A mixture of chlorine and hydrogen appears not to react at room temperature, but when exposed to sun light or heated, the mixture explodes forming misty fumes of hydrogen chloride gas.

$$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$$

4. Reaction with reactive metals

Hydrogen reacts with highly reactive metals to form hydrides. For example, sodium reacts with hydrogen to form sodium hydride.

$$Na(s) + \tfrac{1}{2}H_2(g) \longrightarrow NaH(s)$$

Uses of hydrogen

1. Hydrogen is a very light gas and therefore used to fill balloons.
2. Hydrogen with oxygen form oxy-hydrogen flame which is very hot and used for welding and cutting metals.
3. Hydrogen is used in the manufacture of ammonia by Haber process.
4. Hydrogenation of vegetable oil makes it hard and used for making margarine and cooling fats like blue band and kimbo.
5. It is used in the manufacture of hydrogen bombs.
6. Hydrogen is also used as fuel for rockets.

Sample questions on water and hydrogen

1. *Briefly describe the water cycle. Mention the major sources of water. How is waater beneficial to you and to the community from where you come?*
2. *What is water pollution? What are the main sources of water pollution? How do the pollutants mainly reach the water bodies?. Describe briefly the effects of water pollution to living organisms in water bodies.*
3. *In an attempt to make water available for domestic use, water is treated. Explain in detail the steps taken in water purification processes.*
4. *Describe the reactions of sodium, potassium, magnesium and calcium with water stating clearly the conditions under which the reactions take place.*

5. *What is meant by soft and hard water? Give examples of each. Explain the types of hard water that you know giving the causes of each; outline the physical and chemical methods of softening each type of hardness. (Where appropriate use equations to illustrate). Mention the advantages and disadvantages of hard water.*
6. *Describe the aid of a labeled drawing how hydrogen gas is prepared in the laboratory. How do you confirm that the gas produced in your description is hydrogen gas?*
7. *How do you show that water is an oxide of hydrogen? Describe the reduction reaction of hydrogen using of copper(II) oxide (What are the precautions for the reaction). Outline at least five uses of hydrogen gas.*

SALTS

A salt is a compound formed when either all of part of the ionisable hydrogen of the acid is replaced by a metallic ion or ammonium ion. Or

A salt is an ionic compound consisting of a positive metallic or ammonium ion and a negative ion derived from an acid.

Salts get their names from the acids they are derived from. Examples are in the table below.

Acid	**Salt**	**Example**
Hydrochloric acid (HCl)	Chlorides	Sodium chloride ($NaCl$)
Sulphuric acid (H_2SO_4)	Sulphates	Ammonium sulphate $(NH_4)_2S0_4$
Carbonic acid (H_2CO_3)	Carbonates	Calcium hydrogen carbonate, $Ca(HCO_3)_2$
Nitric acid (HNO_3)	Nitrates	Barium nitrate,$Ba(NO_3)_2$
Phosphoric acid (H_3PO_4)	Phosphates	Sodium phosphate,Na_3PO_4
Ethanoic acid (CH_3COOH)	Ethanoates	Sodium ethanoate,CH_3COONa
Sulphurous acid (H_2SO_3)	Sulphites	Sodium sulphite. Na_2SO_3

Types of salts

a) Normal salt

This is a salt produced when all the ionisable hydrogen of the acid is replaced by a metallic or ammonium ion. These salts do not contain ionisable hydrogen. Examples include; sodium chloride, $NaCl$; ammonium nitrate, NH_4NO_3; Magnesium sulphate, $MgSO_4$; lead (II) bromide, $MgBr_2$ and sodium phosphate, Na_3PO_4.

All normal salts have PH of 7 except salts formed from

i) Strong bases and weak acids e.g. sodium carbonate (Na_2CO_3) and potassium ehanoate (CH_3COOK). These salts in solution have PH value more than 7.

ii) Strong acids and weak bases e.g. ammonium chloride (NH_4Cl). The salts have a PH value less than 7 in solution.

b) Acid salts

An acid salt is a salt formed when only part of the ionisable hydrogen of the acid is replaced by ammonium or metallic ion. These salts contain ionisable hydrogen and examples include: sodium hydrogensulphate, $NaHSO_4$; calcium hydrogencarbonate, $Ca(HCO_3)_2$; and potassium hydrogen carbonate, $KHCO_3$.

These acid salts behave like salts because they contain metallic ion and a negative ion derived from an acid; they behave like acids because the negative ions are capable of further ionization to yield hydrogen ion(H^+) (i.e $HSO_4^- \rightleftharpoons H^+ + SO_4^{2-}$).

c) Basic salts

Basic salt are formed when insufficient acid is present to neutralize the available base. E.g. basic zinc chloride ($Zn(OH)Cl$) and basic magnesium chloride ($Mg(OH)Cl$).

All monobasic acids form normal salts while dibasic and tribasic acids form both normal and acid salts.

Naming of salts

Salts are named by adding the name of the radical or ion of the acid after the name of the metal or ammonium. Examples are:

Name of metal/ammonium radical	Name of acid radical (negative ion)	Name of salt
Sodium	Chloride	Sodium chloride, $NaCl$
Aluminium	Sulphate	Aluminium sulphate, $Al_2(SO_4)_3$

Ammonium	Nitrate	Ammonium nitrate NH_4NO_3
Potassium	Carbonate	Potassium carbonate, K_2CO_3

SOLUBILITY OF SALTS

Solubility is the amount of solute in grams required to saturate 100g of solvent (water) at a particular temperature.

A salt is described as soluble if it can dissolve in a given solvent and insoluble if it cannot dissolve in the solvent. Salts have varying degree of solubility in water as described below:

All ammonium, sodium, and potassium salts are soluble in water.
All nitrate salts are soluble in water.
All chloride salts are soluble in water except silver chloride, lead (II) chloride (sparingly soluble) and mercury (I) chloride.
All sulphate salts are soluble in water except lead (II) sulphate and barium sulphate. Calcium sulphate is sparingly soluble in water.
All carbonate salts are insoluble in water except sodium, potassium and ammonium carbonates.

Determining the solubility of a salt e.g. sodium chloride

Procedure

- Take about 50cm^3 of distilled water in a beaker.
- Add sodium chloride crystals to the water a little at a time while stirring continuously until when no more salt dissolves. The solution formed is saturated.
- Weigh a clean evaporating dish and pour into it a little of the clear salt solution.
- Weigh the evaporating dish with the salt solution and evaporate the solution to dryness carefully through a water bath.
- Allow the evaporating dish to cool and reweigh the dish with the dry salt.

Results

Mass of empty dish= a g

Mass of dish + saturated solution = b g

Mass of dish + dry salt = c g

Calculations

Mass of saturated solution = (b-a) g

Mass of dry salt = (c-a) g

Mass of water (solvent) = (b-a)-(c-a)

$= b-\cancel{a}-c+\cancel{a}$

= (b-c) g

(b-c) g of water dissolves (c-a) g of NaCl

1 g of solvent dissolves $\frac{(c-a)}{(b-c)}$ g of NaCl

100g of solvent dissolves $\frac{(c-a)}{(b-c)}$ x 100 g of NaCl

There fore, the solubility of NaCl at room temperature is $\frac{(c-a)}{(b-c)}$ x 100 g/100 g of water.

Examples

1. 30g of sodium chloride crystals were dissolved in 75 g of water at 80°C. Calculate the solubility of sodium chloride at this temperature.

Solution

75 g of water dissolves 30 g of NaCl

1 g of solvent dissolves $\frac{30}{75}$ g of NaCl

100g of solvent dissolves ($\frac{30}{75}$ x 100) g of NaCl

=40g

There fore, the solubility of NaCl at 80°C temperature is 40g/100 g of water.

2. 12.0g of potassium chlorate was carefully evaporated to dryness, 2.4 g of potassium chlorate crystals were left on the evaporating dish. Calculate the solubility of potassium chlorate in grams per 100g of water at room temperature.

Solution

Mass of solvent = mass of solution–mass of solute

=(12.0-2.4)g = 9.6g of solvent (water)

9.6 g of water dissolves 2.4g of potassium chlorate

1 g of solvent dissolves $\frac{2.4}{9.6}$ g of potassium chlorate

100g of solvent dissolves ($\frac{2.4}{9.6}$ x 100) g of potassium chlorate

=25g

There fore, the solubility of potassium chlorate at room temperature is 25g/100 g of water.

3. 75g of a saturated solution contains 30g of salt. Calculate,
   i) The solubility of the salt
   ii) The percentage of salt in the saturated solution

Solution

i) Mass of solvent = masss of solution- mass of solute

=(75-30)g = 45g of solvent

45 g of solvent is saturated by 30g of salt

1 g of solvent is saturated by $\frac{30}{45}$ g of salt

100g of solvent saturated by ($\frac{30}{45}$ x 100) g of salt

=66.67g

There fore, the solubility of the salt at room temperature is 66.67g/100 g of water.

ii) Percentage of salt = $\frac{mass\ of\ salt}{mass\ of\ solution}$ x 100

= $\frac{30}{75}$ x 100

=40%

Exercise

1. In an experiment to determine the solubility of potassium nitrate at 20˚C, the following results were obtained.
   Mass of evaporating dish + saturated solution = 100.7g
   Mass of evaporating dish = 65.3g
   Mass of dish + dry salt = 73.8 g
   Use the data above to calculate the solubility of potassium nitrate at 20˚C. Clearly show your working.

Factors that affect the rate of solubility of salts

1. **Amount of solvent**

Solubility of most salts increase with increase in the amount of solvent used.

2. **Nature of solvent/solute**

Solubility of a salt may increase or decrease depending on the nature of solvent or solute.

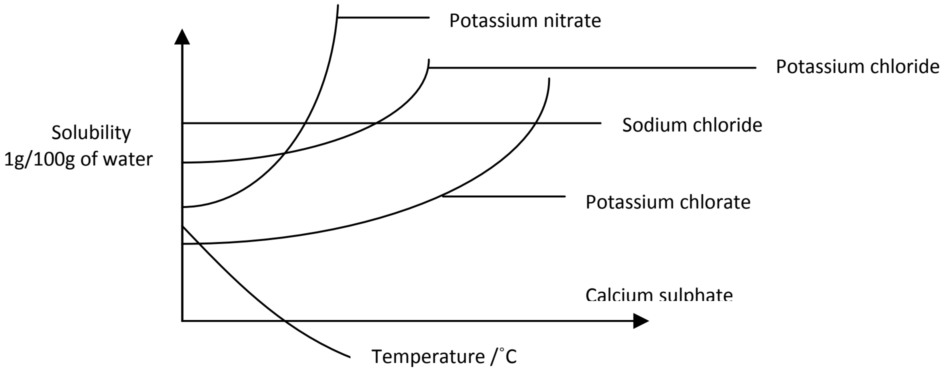
3. **Temperature**

Solubility of most salts increase with increase in temperature. For example, potassium chlorate and potassium nitrate. Solubility of a few salts like calcium chloride and calcium sulphate decrease with increase in temperature. The solubility of sodium hydroxide and gases as well also decrease with increase in temperature.

Solubility curve

A solubility curve is a graph that shows how the solubility of a salt varies with temperature. The graph is obtained by plotting solubility (on the vertical axis) against temperature (on the horizontal axis).

Solubility curve of some common salts



The solubility of potassium chloride, potassium nitrate and potassium chlorate increase with increase in temperature. The solubility of potassium nitrate increases most rapidly, followed by potassium chlorate then potassium chloride.
The solubility of sodium chloride increases very slightly with increase in temperature.
The solubility of calcium sulphate decreases with increase in temperature.

Uses of solubility curves

1. It can be used to find the solubility of a salt at a given temperature.
2. It gives the temperature at which a given amount of salt saturates 100g of solvent.
3. It can be used to explain the trend of solubility of salts.
4. A solubility curve can be used to calculate the mass of salt obtained by cooling a solution from a higher temperature to a lower temperature.

   Mass of salt= (solubility at a higher temperature – solubility at a lower temperature)

For example, if a salt P with solubility of 180g/100g of water at 90°C was cooled to a temperature of 30°C where its solubility is 25g/100g of water. Calculate the mass of salt formed formed after cooling the solution.

Solution

Mass of salt= (solubility at a higher temperature – solubility at a lower temperature)
= (180-45)g
= 155g

Application of solubility

1. Solubility is used to separate soluble salts from a mixture by fractional crystallization.
2. It is used in the extraction of salts from large water bodies like lakes and seas.

Exercise

1. a) Describe an experiment that you would carryout to determine the solubility of potassium nitrate at 15˚C.

   b) Determine the solubility in water of substance S at room temperature from the following data.

Mass of evaporating basin	25g
Mass of evaporating basin + Saturated solution of S	55g
Mass of evaporating basin + Solid S	30g

2. a) Define the term solubility?

   b) The table below shows the solubility (ies) of salt P in water at different temperatures.

Temperature/˚C	10	20	30	40	50	60
Solubility, g/100g of water	18	20	24	30	38	50

   i) plot a graph of solubility of P against temperature
   ii) use your graph to determine
      a) solubilities of P at 25˚C and 45˚C
      b) the mass of crystals deposited when a solution of P is cooled from 50˚C to 25˚C
   iii) calculate the mass of P that would dissolve in 45g of water at 25˚C

PREPARATION OF SALTS

The method of salt preparation depends on whether the salt is soluble in water or not. Soluble salts are prepared by **crystallization** and **neutralization**. Insoluble salts are prepared by **precipitation** or **double decomposition.** Other salts are prepared by direct synthesis.

Preparation of soluble salts

Soluble salts are prepared using dilute acids and metals, metals oxides, metal hydroxides and metal carbonates.

General procedure

1. Place some dilute acid in a beaker
2. Warm the acid and add the metal, metal oxide, metal hydroxide and metal carbonate bit by bit until in excess to ensure that the acid is completely used up.
3. Filter the excess metal, metal oxide, metal hydroxide or metal carbonate and collect the filtrate.
4. Saturate the filtrate by evaporating and allow the solution to cool as it cools to form the salt crystals.
5. Filter the crystals and wash them with water.
6. Dry the crystals in an oven, or under sun shine or between filter papers.

1. Preparation of salts from metals and dilute acids

Salts prepared by this method are soluble salts of iron, magnesium, aluminium and zinc. (I.e. metals higher than lead and lower than calcium in the reactivity series.)

N.B. Nitrates cannot be prepared using this method because dilute nitric acid being an oxidizing agent, does not react with metal to liberate hydrogen gas.

Example

Laboratory preparation of zinc sulphate crystals from zinc metal/powder

- Put dilute sulphuric acid in a beaker and heat it gently until when it’s hot.
- Add zinc powder to the hot acid bit by bit while stirring until when the zinc powder is in excess.
- Filter off the excess zinc powder to obtain zinc sulphate solution as the filtrate.
- Saturate the filtrate by evaporating.
- Allow it to cool and form crystals of the salt.
- Filter the crystals and wash them with distilled water.
- Dry the crystals either in qaan oven of under the sun or between filter papers.

Equation

$$H_2SO_4(aq) + Zn(s) \longrightarrow ZnSO_4(aq) + H_2(g)$$

Other salts formed in similar ways are:

Zinc chloride

Hydrochloric acid + zinc $\longrightarrow$ Zinc chloride + Hydrogen gas

$HCl(aq) + Zn(s) \longrightarrow ZnCl_2(aq)$

Iron (II) sulphate

Sulphuric acid + Iron $\longrightarrow$ Iron (II) sulphate + Hydrogen gas

$H_2SO_4(aq) + Fe(s) \longrightarrow FeSO_4(aq) + H_2(g)$

2. Preparation of salts from metal oxides and dilute acids

Example

Preparation of copper (II) sulphate from copper (II) oxide in the laboratory

- Put dilute sulphuric acid in a beaker and heat it gently until when it's hot.
- Add copper (II) oxide to the hot acid bit by bit while stirring until when the copper (II) oxide is in excess.
- Filter off the excess copper (II) oxide to obtain copper sulphate solution as the filtrate.
- Saturate the filtrate by evaporating.
- Allow it to cool and form crystals of the salt.
- Filter the crystals and wash them with distilled water.
- Dry the crystals either in an oven or under sunshine or between filter papers.

Equation

$H_2SO_4(aq) + CuO(s) \longrightarrow CuSO_4(aq) + H_2O(g)$

Other examples of salts formed from metal oxides are:

Magnesium chloride

Hydrochloric acid + Magnesium oxide $\longrightarrow$ Magnesium chloride + Water

$2HCl(aq) + MgO(s) \longrightarrow MgCl_2(aq) + H_2O(l)$

Copper (II) nitrate

Copper (II) oxide + Nitric acid $\longrightarrow$ Copper (II) nitrate + Water

$CuO(s) + HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + H_2O(g)$

Aluminium sulphate

$H_2SO_4(aq) + Al_2O_3(s) \longrightarrow Al_2(SO_4)_3(aq) + H_2O(g)$

3. Preparation of salts from insoluble metal carbonates

Example

Preparation of lead (II) nitrate from lead (II) carbonate

- Pour dilute nitric acid in a beaker and warm it gently.
- Add lead (II) carbonate a little at a time. Effervescence occurs as carbondioixde is evolved.
- Continue adding the carbonate until when it is in excess and no more effervescence occurs.
- Filter off the excess carbonate to get a colourless filtrate.
- Evaporates the filtrate by heating gently to obtain a saturated solution.
- Cool the saturated solution to obtain white crystals of lead (II) nitrate salts.
- Wash the crystals with cold distilled water and dry either on sun shine, in an oven or between filter papers.

Equation

Lead (II) nitrate

Lead (II) carbonate + Nitric acid ⟶ Lead (II) nitrate + Water + carbondioxide

$HNO_3(aq) + PbCO_3(s) \longrightarrow Pb(NO_3)_2(aq) + H_2O(g) + CO_2(g)$

Other salts prepared in similar ways are:

Barium chloride

Hydrochloric acid + Barium carbonate ⟶ Barium chloride + Water + carbondioxide

$2HCl(aq) + BaCO_3(s) \longrightarrow BaCl_2(aq) + H_2O(l) + CO_2(g)$

Copper (II) sulphate

Sulphuric acid + copper (II) carbonate ⟶ Copper (II) sulphate + water + carbondioxide

$H_2SO_4(aq) + CuCO_3(s) \longrightarrow CuSO_4(aq) + H_2O(g) + CO_2(g)$

4. Preparation of salts from metal hydroxides

Example

Preparation of lead (II) nitrate starting from lead (II) hydroxide

- Pour dilute nitric acid in a beaker and warm it gently.
- Add lead (II) hydroxide a little at a time while until when it is in excess.
- Filter off the excess hydroxide to get a colorless filtrate.
- Evaporates the filtrate by heating gently to obtain a saturated solution.
- Cool the saturated solution to obtain white crystals of lead (II) nitrate salts.
- Wash the crystals with cold distilled water and dry them either on sun shine, in an oven or between filter papers.

Equation

$Pb(OH)_2(s) + 2HNO_3(aq) \longrightarrow Pb(NO_3)_2(aq) + 2H_2O(l)$

Other salts prepared in similar ways are:

Iron (II) nitrate

Iron (II) hydroxide + Nitric acid ⟶ Iron (II) nitrate + Water

$Fe(OH)_2(s) + HNO_3(aq) \longrightarrow Fe(NO_3)_2(aq) + H_2O(g)$

Copper (II) chloride

Hydrochloric acid + Copper (II) hydroxide ⟶ Copper chloride + Hydrogen gas

$2HCl(aq) + Cu(OH)_2(s) \longrightarrow CuCl_2(aq) + 2H_2O(l)$

Zinc sulphate

Zinc hydroxide + sulphuric acid ⟶ zinc sulphate + water

$Zn(OH)_2(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + 2H_2O(l)$

Laboratory preparations of salts whose carbonates, oxides and hydroxides sare insoluble

These salts include potassium, sodium and ammonium salts. The salts can be prepared by titration method (neutralization).

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

General procedure

- Put a known volume of hydroxide of a metal in a conical flask.
- Add 2 or 3 drops of an indicator.

- Run a suitable acid from the burette until when the color of the mixture just changes. Note and record the volume of acid used.
- Measure accurately the same volume of hydroxide as before and titrate with exactly the same volume of acid as recorded above.

  NB. An indicator is not used in the second titration, since the volume of acid required to neutralize the fixed volume of base was already got.
- Stir and heat the solution to make it saturated.
- Allow the hot saturated solution to cool as it forms salt crystals.
- The crystals are filtered off, washed with cold distilled water and dried in an oven, under sun shine or between filter papers.

Preparation of sodium chloride crystals in the laboratory

Procedure

- Put a known volume of sodium hydroxide in a conical flask.
- Add 2 or 3 drops of an indicator.
- Titrate the sodium hydroxide with hydrochloric acid from the burette until when the end point is reached (when the indicator changes color). Note and record the volume of acid used.
- Measure accurately the same volume of sodium hydroxide as before and titrate with exactly the same volume of hydrochloric acid as recorded above without using an indicator.
- Stir and heat the solution to make it saturated.
- Allow the hot saturated solution to cool as it forms salt crystals.
- The crystals are filtered off, washed with cold distilled water and dried in an oven, under sun shine or between filter papers.

Equation

$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$

Other salts prepared in similar ways are:

Ammonium sulphate ($(NH_4)_2SO_4$)

Ammonium hydroxide + sulphuric acid ⟶ ammonium sulphate

$2NH_4OH(aq) + H_2SO_4(aq) \longrightarrow (NH_4)_2SO_4(aq) + H_2O(l)$

Potassium nitrate, KNO_3

Potassium hydroxide + nitric acid $\longrightarrow$ potassium nitrate + water

$KOH(aq) + HNO_3(aq) \longrightarrow KNO_3(aq) + H_2O(l)$

Sodium sulphate

Sodium hydroxide + sulphuric acid $\longrightarrow$ sodium sulphate + water

$2NaOH(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$

Preparation of insoluble salts

Insoluble salts are prepared by **double decomposition** or **precipitation** method. In this method, two soluble salts are mixed to form two new salts by exchange of radicals. One of the new salts formed is a soluble salt and one is an insoluble salt that appear as precipitates. The precipitate is filtered off and washed then dried.

Precipitation is the formation of solids when solutions are mixed.

A precipitate is the solid formed when two or more solutions are mixed.

Example

Preparation of barium sulphate (by reacting barium nitrate and sodium sulphate)

Procedure

- Put a solution of barium nitrate in a beaker and add a solution of sodium sulphate to it. A white precipitate of barium sulphate immediately appears,
- Filter off the precipitate and wash with distilled water.
- Dry the precipitate (salt formed) under sun shine, in an oven or between filter papers.

Equation

$Ba(NO_3)_2(aq) + Na_2SO_4(aq) \longrightarrow BaSO_4(s) + NaNO_3(aq)$

This method can be used to prepare salts such as lead sulphate, aluminium chloride, silver chloride, silver carbonate and barium sulphate.

Preparation of lead (II) sulphate (by reacting lead (II) nitrate and sulphuric acid)

Procedure

- Add dilute nitric acid to lead (II) nitrate solution in a beaker and stir the mixture. White precipitates of lead (II) sulphate is formed.
- Filter off the precipitates and wash with distilled water to remove traces of the acid.
- Dry the precipitates in a steam oven or leave it to dry in air.

NB. If any of the compounds to be used in the preparation of the salt is insoluble in water, it must first be made to dissolve in a mineral acid. For example, in the preparation of lead (II) sulphate using lead (II) oxide, the lead (II) is first divvolved in nitric acid to form lead (II) nitrate.

Equation

$PbO(s) + 2HNO_3(aq) \longrightarrow Pb(NO_3)_2(aq) + H_2O(l)$

The lead (II) nitrate formed reacts with sulphuric acid to form lead (II) sulphate.

$Pb(NO_3)_2(aq) + H_2SO_4(aq) \longrightarrow PbSO_4(s) + HNO_3(aq)$

Preparation of salts by direct synthesis

Salts consisting of two elements (binary salts) can be prepared by direct synthesis/ direct combination.

Example:

In the preparation of sodium chloride from sodium and chlorine, burning sodium is lowered in a gas jar of chlorine. Sodium continues to burn in chlorine forming white fumes which settle into white solids (sodium chloride).

Equation

$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$

Other salts prepared by direct synthesis include:

Magnesium chloride

$Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)$

Potassium chloride

$2K(s) + Cl_2(g) \longrightarrow 2KCl(s)$

Iron (III) chloride

$2Fe(s) + 3Cl_2(g) \longrightarrow 2FeCl_3(s)$

Zinc bromide

$Zn(s) + Br_2(g) \longrightarrow ZnBr_2(s)$

Effects of heat on salts

1. Carbonates

Potassium and sodium carbonates are very stable and are not decomposed by heat. But if the salts are hydrated, they lose their water of crystallization. In such a process, salts lose their crystalline nature and become amorphous.

$Na_2CO_3.10H_2O(s) \longrightarrow Na_2CO_3(s) + 10H_2O(g)$

(Crystalline solid) (Powdery solid)

All the other metallic carbonates decompose upon heating to give the oxide of the metal and carbondioxide gas.

Example

When white zinc carbonate is heated, it produces a colorless gas that turns lime water milky leaving a yellow residue when hot which turns white on cooling.

$ZnCO_3(s) \longrightarrow ZnO(s) + CO_2(g)$

(white solids)

(yellow solids when hot but turns white on cooling)

Colorless gas that turns lime water milky

When lead (II) carbonate is heated, a brown residue (when hot) which becomes yellow on cooling and a colorless gas that turns lime water milky are produced.

$PbCO_3(s) \longrightarrow PbO(s) + CO_2(g)$

(white solids)

(Brown solids when hot but turns yellow on cooling)

Colorless gas that turns lime water milky

When copper (II) carbonate is heated, black solids of copper (II) oxide is formed and a colorless gas that turns lime water milky evolved.

$$CuCO_3(s) \longrightarrow CuO(s) + CO_2(g)$$

(white solids) (black solids) Colorless gas that turns lime water milky

Ammonium carbonate decomposes to give ammonia gas, carbondioxide and water vapor.

$$(NH_4)_2CO_3(s) \rightleftharpoons 2NH_3(g) + H_2O(g) + CO_2(g)$$

Hydrogen carbonate of metals decompose to form carbonate of metals, carbondioxide gas and water vapor.

Example

$$2NaHCO_3(s) \longrightarrow Na_2CO_3(s) + H_2O(l) + CO_2(g)$$

2. Sulphates

Sulphates of sodium and potassium donot decompose on heating. When hydrated sulphates of potassium or sodium is heated, it loses its water of crystallization and becomes amorphous.

$$K_2SO_4.10H_20(s) \longrightarrow K_2SO_4(s) + 10H_2O(g)$$

Sulphates of heavy metals decompose to give metal oxides and white fumes of sulphur trioxide gas. When heated more strongly, the sulphur trioxide gas decomposes to give sulphur dioxide and oxygen gas.

Examples

When hydrated copper (II) sulphate crystals are heated, they lose their water of crystallization and changes from blue crystals to white powder. The water condenses as a colorless liquid on the cooler parts of the test tube.

$$CuSO_4.5H_2O(s) \longrightarrow CuSO_4(s) + 5H_2O(g)$$

(blue crystals) (white powder)

On further heating,the white powder gives off white fumes of a gas (SO_3) and a black residue (CuO) is left.

$2CuSO_4(s) \longrightarrow 2CuO(s) + SO_3(g)$

(white powder) (black solids)

When green solid of iron (II) sulphate is heated, it loses it water of crystallization and changes from green to dirty-yellow anhydrous solids.

$FeSO_4.7H_2O(s) \longrightarrow FeSO_4(s) + 7H_2O(g)$

(Green solids) (Dirty yellow solids)

When heated more strongly, it gives off sulphur dioxide (a colorless gas that turns potassium dichromate solution from yellow to green), white fumes of sulphur trioxide and brown residue of iron (III) oxide is left.

$2FeSO_4(s) \longrightarrow Fe_2O_3(s) + SO_3(g) + SO_2(g)$

(dirty yellow) (brown solids)

3. Nitrates

All nitrates decompose upon heating.

Sodium and potassium nitrates melt into colorless liquids then decompose upon heating to give their corresponding nitrites that form yellow solids on cooling and oxygen gas.

$2NaNO_3(s) \longrightarrow 2NaNO_2(s) + O_2(g)$

(white crystals) (yellow powder)

All nitrates from calcium down to copper decompose to give their corresponding oxides, brown fumes of nitrogen dioxide gas and oxygen gas.

Examples

$2Zn(NO_3)_2(s) \longrightarrow 2ZnO(s) + 4NO_2(g) + O_2(g)$

(White crystals) (Yellow-hot; white-cold) (Brown fumes)

$2Pb(NO_3)_2(s) \longrightarrow 2PbO(s) + 4NO_2(g) + O_2(g)$

(White crystals) (Brown-hot; yellow-cold) (Brown fumes)

N.B. Lead(II)nitrate decomposes with a cracking sound.

$2Cu(NO_3)_2(s) \longrightarrow 2CuO(s) + 4NO_2(g) + O_2(g)$

(Green crystals) (Black powder) (Brown fumes)

Mercury (II)nitrate and silver nitrate decompose to give their corresponding metals, brown fumes of nitrogen dioxide gas and a colorless gas that relights a glowing spling (oxygen gas).

$2AgNO_3(s) \longrightarrow Ag(s) + 2NO_2(g) + O_2(g)$

(White crystals) (Grey solids) (Brown fumes)

$Hg(NO_3)_2(s) \longrightarrow Hg(l) + 2NO_2(g) + O_2(g)$

Ammonium nitrate sublimes upon heating to give dinitrogen oxide and water vapor.

$NH_4NO_3(s) \longrightarrow N_2O(g) + 2H_2O(g)$

4. Chlorides

Metallic chlorides are not affected by heat because hey are very stable. However, if they are hydrated, the lose their water of crystallization.

For example

$MgCl_2.7H_2O(s) \longrightarrow MgCl_2(s) + 7H_2O(g)$

Ammonium chloride sublimes on slight heating and on further heating, it decomposes to give ammonia and hydrogen chloride gases.

Equation

$NH_4Cl(s) \rightleftharpoons NH_4Cl(g) \longrightarrow NH_3(g) + HCl(g)$

Effects of heat on hydroxides

Hydroxides of sodium and potassium are not decomposed by heat. If they are in solid forms, they absorb moisture and melt to form solutions.

However, the hydroxides of other metals decompose to give the corresponding oxides and water vapor.

Example

$Zn(OH)_2(s) \longrightarrow ZnO(s) + H_2O(g)$

(white crystals) (yellow-hot; white-cold)

$Pb(OH)_2(s) \longrightarrow PbO(s) + H_2O(g)$

(white crystals) (brown-hot; brown-cold)

Sample questions on salts

1. *Briefly explain what is meant by the following and in each case give an example (formula and name): salt, acid salt, normal salt and basic salt.*
2. *What is solubility of a salt? Describe with the aid of a labeled drawing how you can determine the solubility of sodium chloride at 40 °C*
3. *(a) State one method for preparing: (i) lead (II) nitrate other than from lead (II) carbonate (ii)lead (II) sulphate. Describe how a pure dry sample of lead (II) nitrate can be prepared in the laboratory, starting from lead (II) carbonates.*
4. *Describe what would be observed and write equations when each of the following salts are heated: (a) lead(II) carbonate (b) lead(II) nitrate (c) sodium carbonate crystals (d) zinc carbonate (e) ammonium carbonate (f) hydrated copper(II) sulphate (g) sodium nitrate (h) zinc nitrate crystals (i) silver nitrate and ammonium nitrate.*
5. *Describe in detail how you would prepare pure dry samples of (a) zinc sulphate from zinc carbonate (b) lead(II) sulphate from lead(II) nitrate (c) sodium chloride by action of an acid and a base.*
6. *Write down the formula of the precipitate which is produced when pairs of aqueous solutions are mixed as follows: (i)* $Ba(NO_3)_2$ *and* Na_2SO_4*; (ii)* H_2S *and* $Pb(NO_3)_2$*; (iii)* NH_4Br *and* $AgNO_3$*; (iv)* $CaCl_2$ *and* K_2CO_3*.*

BONDING AND STRUCTURE

BONDING

Bonding is the chemical combination of atoms or elements to form compounds. The force of attraction holding atoms or elements together in a molecule/crystal is referred to as a **chemical bond**. Chemical bonding /combination occurs mainly in four forms as:

1. **Ionic/electrovalent bonding**-this involves transfer of electrons from a metal atom to a non metal atom. It occurs between metals and non metals.
2. **Covalent bonding**-this involves sharing of electrons between two or more non metal atoms/elements. The atoms/elements involved contribute to the bonding electrons.

3. **Dative/co-ordinate bonding**-this involves sharing of the bonding electrons which are donated by one molecule or atom involved.
4. **Metallic bonding**-this occurs between atoms of metal elements.

Electrovalent/ ionic bonding

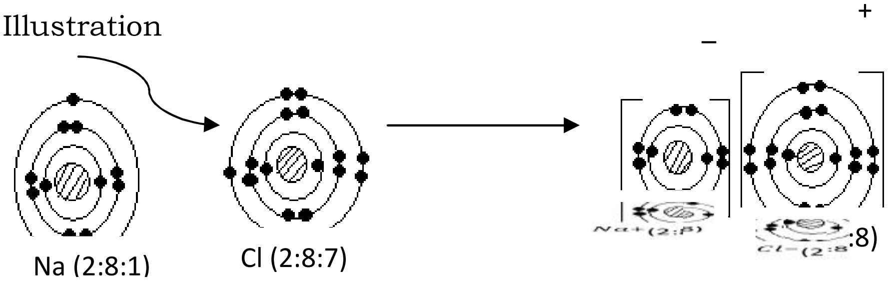
This involves transfer of electrons from a metal to a non metal. The number of electrons lost by the metal atom or gained by non metal atom is equivalent to its valency. The loss of electrons from a metal atom leads to formation of a positively charged ion (cation) and the gain of electrons by a non metal atom leads to the formation of a negatively charged ion (anion). The positively charged ion and the negatively charged ion are attracted to each other and the force of attraction holding them together is known as electrovalent/ionic bond. The compounds formed are referred to as **ionic/ electrovalent compounds**. Metal atoms lose electrons so as to gain stable electronic configuration of noble gases and non metal atoms also gain electrons to become stable.

Formation of some ionic compounds

Sodium chloride

When sodium burns in chlorine to form sodium chloride, the sodium atom loses an electron and forms a positively charged sodium ion(Na^+) while the chlorine atom gains the electron and forms a negatively charged chloride ion (Cl^-).

Illustration

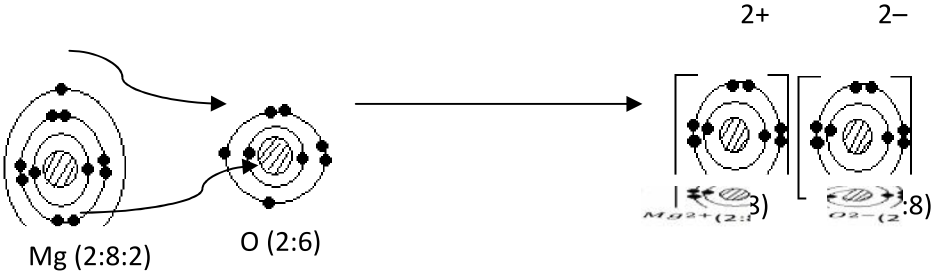


The positively charged sodium ion and negatively charged chlorine ion are attracted to each other by a strong force called ionic bond.

Magnesium oxide

Magnesium atom loses two electrons to the oxygen atom and forms a positively charged magnesium ion (Mg^{2+}) while the oxygen atom forms a negatively charged oxide ion (O^{2-}). These two oppositely charged ions are then attracted to each other.

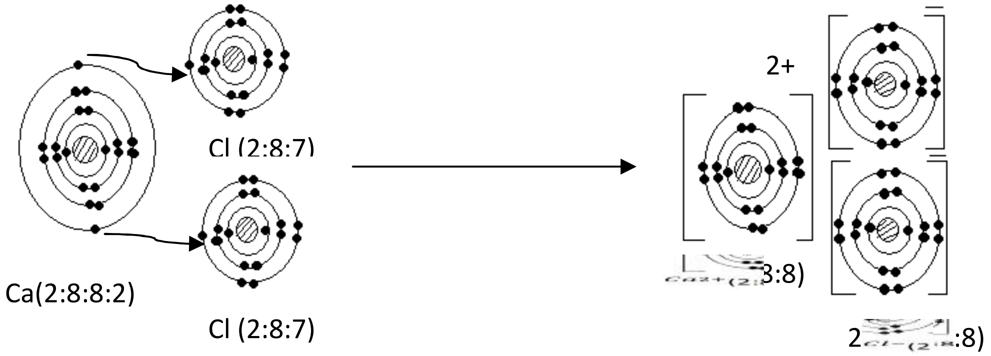
Illustration



Calcium chloride

During the formation of calcium chloride, calcium atom loses two electrons which are gained by two chlorine atoms. This leads to the formation of $Ca^{2+}$and two Cl^{-}which are then attracted to each other.

Illustration



Covalent bonding

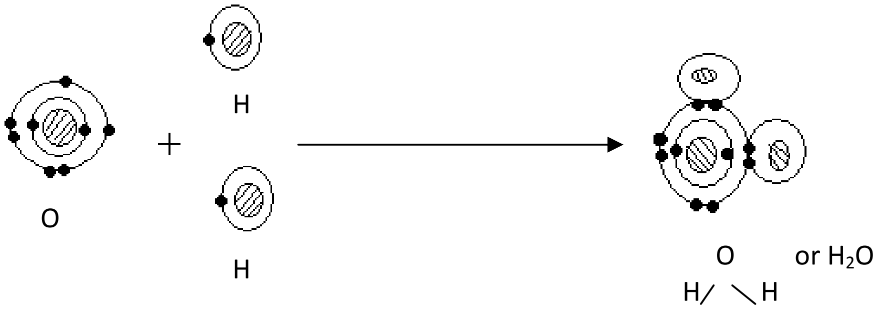
This type of bonding occurs between non metal elements/ atoms to form compounds. Covalent bonding involves mutual sharing of electrons between elements in which each of the atoms involved gives equal number of electrons to be shared. On sharing the electrons, the atoms attain stable electronic structure of noble gases. The compound formed are called covalent compounds. Covalent bonds may by classified as single, double, triple or quadruple depending on the number of electrons shared.

a) Single covalent bond: this is formed when one pair of electron is shared. Each of the atoms involved contribute an electron for the bond formation.

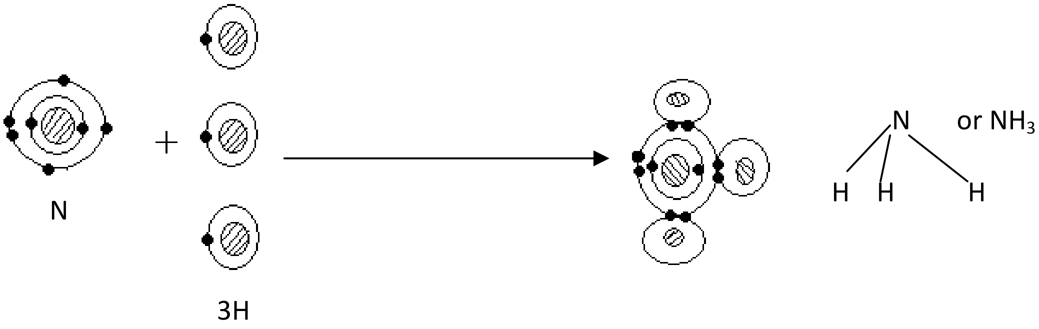
Example

i) Hydrogen molecule (H_2)

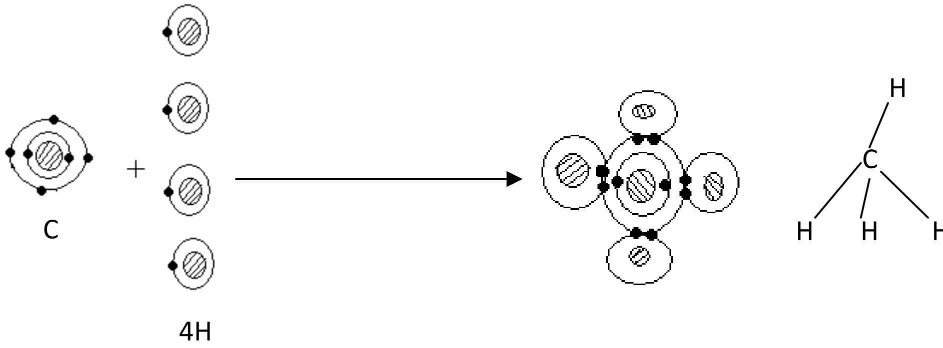
ii) Water molecule (H_2O)



iii) Ammonia (NH_3)



iv) Methane (CH_4)



or CH_4

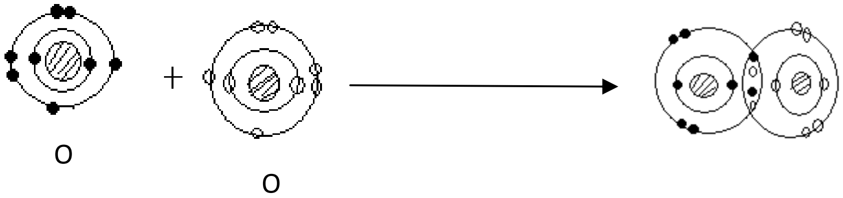
N.B. The pair of electron on the outer most energy level that does not take part in bonding is the **lone pair** or **non bonded pair.**

b) Double covalent bond

This is formed when two pairs of electrons are shared between two bonded atoms.

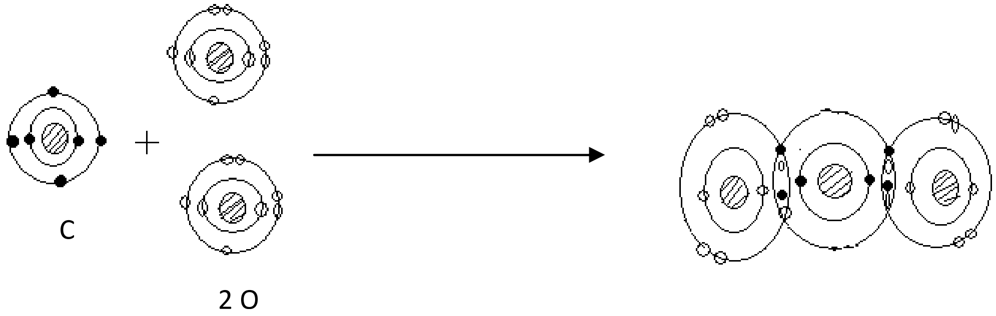
Example

i) Oxygen (O_2)



O=O or O_2

ii) Carbon dioxide (CO_2)



O=C=O or CO_2

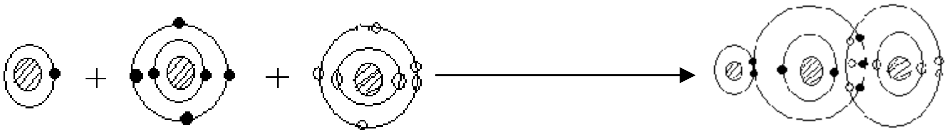
c) Triple covalent bond

This involves sharing of three pairs of electrons between atoms.

Example

i) Nitrogen molecule (N_2)

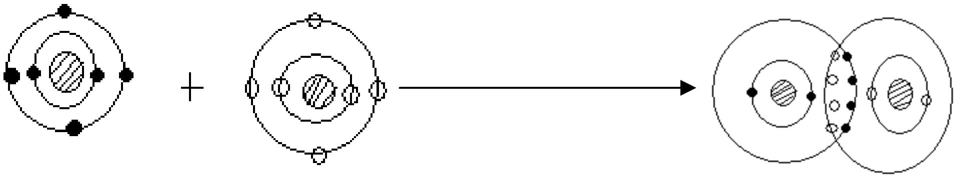
ii) Hydrogen cyanide (HCN)



d) Quadruple covalent bond

This is formed by sharing four pairs of electrons. It occurs between carbon atoms.

Illustration



Dative/ co-ordinate bond

This involves sharing of electrons but the shared pair of electrons is donated by one atom/ molecule. Here, one molecule/ atom donates the pair of electrons to be shared with an ion or another atom. Normally it is atoms/ molecules with lone pair of electrons that form this bond-by donating the lone pair of electrons to be shared with another atom/ion. The bond may be represented by an arrow originating from the donor atom (atom that donates electrons to be shared) to the atom accepting the electrons.

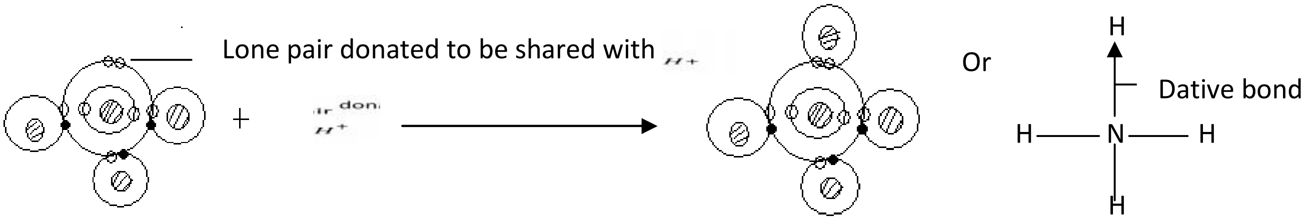
Example

i) Formation of ammonium ion (NH_4^+)

During the formation of ammonium ion, ammonia reacts with an acid. The nitrogen in ammonia has a lone pair of electron which it donates to be shared with the hydrogen ion (proton) from the acid.

$NH_3 + H^+ \longrightarrow NH_4^+$

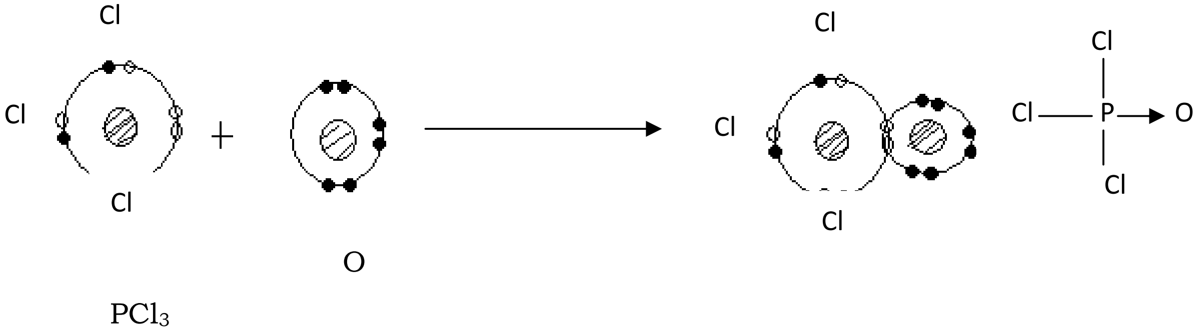
Illustration



ii) Formation of phosphorus (V) trichloride oxide.

This is formed between phosphorus (III) chloride which donates the lone pair of electron to be shared and oxygen atom.

Illustration using outer most shell

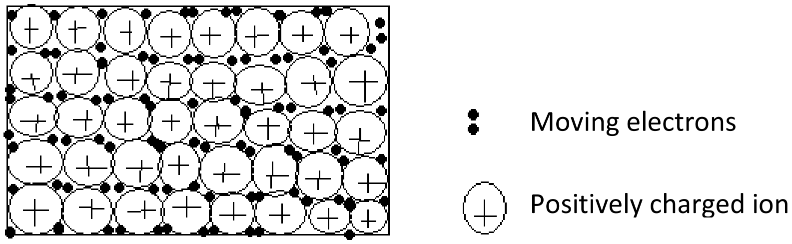


Metallic bonding

This is the type of bonding in metals due to the attraction between metal ions and the valency electrons within the structure of the metal (metal lattice). In the structure of a metal, the loosely held valency electrons (electrons in the outer mos energy level) are released in to a general pool and the atoms become positively charged. These electrons move freely around the ions formed and are termed as **mobile/delocalized electrons**. The ions formed and the electrons attract each other forming metallic bond.

The strength of metallic bond increases with increase in the number of electrons released in to the electron cloud. Thus the bond is very strong in metals like iron and aluminum that release up to three electrons each to the electron pool and weak in metals like sodium and potassium that only release one electron each to the electron pool.

Illustration



Exercise

1. Draw diagrams to show the arrangement of electrons in the following compounds
   a) KCl b) Na_2O c) $MgCl_2$ d) CS_2

2. Draw diagrams to show the arrangement of electrons in the energy levels of atoms in the following molecules. Use only the outer most electrons to illustrate.
   a) Ethane (C_2H_4) b) Ethyne (C_2H_2)

STRUCTURE OF COMPOUNDS/SUBSTANCES

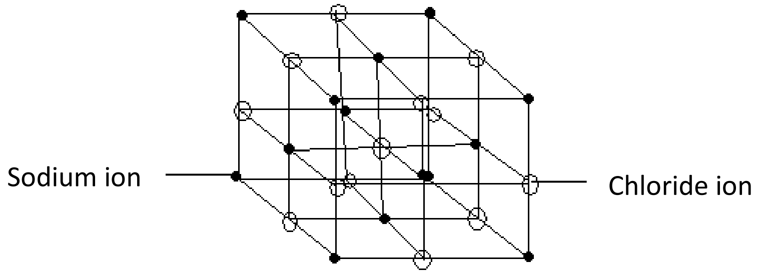
Different types of binds give rise to different types of structures. This is illustrated in the table below.

Bonding	**Structure**
Ionic/ electrovalent	Giant ionic structure
Covalent	Simple molecular structure Giant molecular structure
Metallic	Giant metallic

Giant ionic structure

This is a giant structure consisting of a regular arrangement of a large number of oppositely charged ions resulting into a three dimensional crystal. Each ion is surrounded by a number of oppositely charged ions and they are held together by strong electrostatic forces of attraction. The number of ions surrounding a given oppositely charged ion is known as the **co-ordination number**. The examples of structures include: sodium chloride and magnesium chloride.

Structure of sodium chloride



Each sodium ion is surrounded by six chloride ions and each chloride ion is also surrounded by six sodium ions. The co-ordination number is there fore 6:6.

Properties of ionic compounds

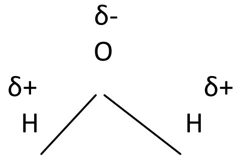
1. They are solids with a regular shape. This is because of the strong electrostatic forces of attraction keeping the ions closely packed.
2. They have high melting points. This is due to the strong electrostatic force af attraction between oppositely charged ions.
3. They do not conduct electricity in solid states but do conduct in their molten states or solution form. This is because in solid states, the ions are localized and the electrons are not mobile but in molten form, the ions and electrons are free to move (electrons become delocalized) and thus conduct electricity.
4. They have high density as the ions are closely packed.
5. Ionic compounds are soluble in water and other polar solvents but they a re insoluble in organic solvents like benzene. Ionic compounds dissolve in polar solvents like water as the ions are attracted by the polar molecule.

A polar compound is a covalent compound in which charge separation exist between atoms. This is due to the differences in electro negativity between the atoms.

Electro negativity is the tendency of an atom to attract bonding electron towards itself. Electro positivity is the tendency of an atom to push away bonding electrons from itself.

For example, in a water molecule oxygen is more electro negative and attracts the bonding electrons towards itself giving it a partial negative charge. Hydrogen attains a partial positive charge.

Structure of a water molecule



NB Most metals are malleable and ductile. This is because, since the bonding agent in a metal is a moving electron cloud, the ions of the metal; usually slide relative to one another under stress (without shattering the lattice and produce a new position of stability).

Simple molecular structures

The structure consists of simple independent molecules joined together by weak intermolecular forces of attraction. E.g. Vander Waal's force of attraction. The atoms in the molecule are held together by strong covalent bond. Compounds with this structure exist as gases, liquids or solids with low melting points. Examples include iodine, carbon dioxide , ammonia and water.

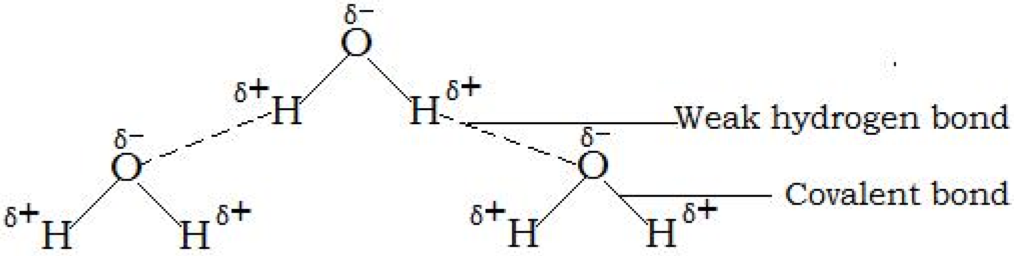
Structure of iodine

I–I-----------I–I------------I–I

Strong covalent bond holding iodine atoms together

Weak Vander Waal's force holding iodine molecules together

Illustration of weak hydrogen bond in water



Properties of simple molecular structures

1. They have low melting and boiling points because the molecules are held by weak intermolecular forces of attraction.
2. They are non conductors of electricity because they do not have mobile electrons/ ions.
3. Most of them are gases and liquids. Very few are solids.
4. They have low densities as the molecules are not closely packed.
5. They dissolve in organic solvents.

Giant atomic/molecular structure

Compounds with such structures consist of molecules /atoms linked together by strong covalent bonds resulting into a giant three dimensional or three dimensional structures like in graphite and diamond respectively (see details under Carbon and its compounds)

Properties of giant molecular structures

1. They have high melting and boiling pints because of great energy needed to break the bonds.
2. They do not conduct electricity except graphite which has delocalized electrons.
3. They are insoluble in water.

Giant metallic structure

Giant metallic structure consists of very many metal ions surrounded by a sea /cloud of electrons. These electrons are released by metal atoms.

Properties of giant metallic structures

1. They are solids with high melting and boiling points except mercury which is a liquid.
2. They conduct electricity in both liquid and solid states due to the presence of mobile electrons.

Sample questions on bonding and structure

Bonding and structure

1. *Describe briefly what is meant by the following types of bonding: electrovalent, covalent, coordinate and metallic. Use examples to illustrate how the above bonds are formed between any two named atoms or molecules.*
2. *Explain what is meant by metallic bond. Explain why a typical metal is (a) a good conductor of electricity (b) a good conductor of heat (c) in some conditions at leas malleable and ductile.*
3. *The compounds named below are all covalent compounds. With the aid of the table of electron structures, give a diagram for a molecule of each of these compounds showing the outer most electron shells only: (i) tetra chloromethane* CCl_4, *phosphorus(III) chloride* PCl_3 *(iii) silane* SiH_4 *(iv) tri chloromethane* $CHCl_3$ *(v) phosphine* PH_3 *and (vi) dichloromethane* CH_2Cl_2
4. *Explain briefly why ammonia and oxygen molecules participate readily in coordinate bonding. Give an example for the formation of an ion by ammonia by this means. Show by means of electronic diagram the formation of a coordinate linkage between phosphorus tri chloride and oxygen.*
5. *By means of electronic diagrams, show the structure of (a) calcium atom (b) chlorine atom (c) calcium chloride. State the differences between electrovalent and covalent compounds.*

6. *Describe how structures of the following compounds account for their property in term of electrical conduction (a) copper (b) graphite (c) diamond*
7. *With the aid of well labeled drawings explain the structure of the following: Graphite, diamond, sodium chloride and chlorine molecule.*

CARBON AND ITS COMPOUNDS

Carbon is an element in period II and group IV in the periodic table. It has atomic number of six.

Occurrence of carbon

Carbon occurs in both free states and combined states. In combined states, it occurs as coal, mineral oils, carbonates (e.g. lime stone, marble and sea shells) and all living things (animals and plants). As an element it occurs in both natural (as diamond and graphite) and synthetic forms (as coke, charcoal and carbon fiber).

Allotropes of carbon

Allotropy is the existence of an element in two or more forms in the same physical state.

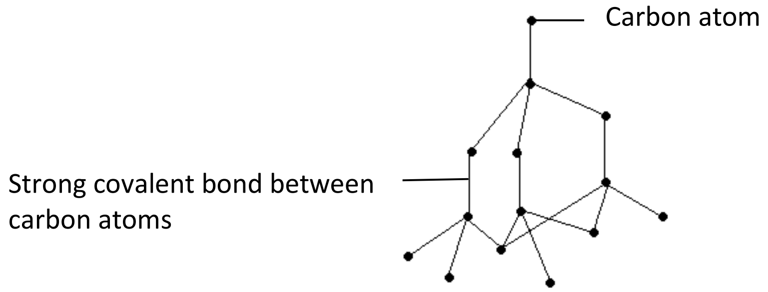
Allotropes are various forms in which elements exist without change in physical state. Different allotropes of the same element always have different crystalline structures and physical properties but the same chemical properties.

Allotropes of carbon are: diamond and graphite. Another form of carbon is amorphous carbon.

Diamond

In diamond, the structure consists of infinite number of carbon atoms. Each of the carbon atoms is joined to four other carbon atoms by covalent bonds resulting into a tetrahedral arrangement. This gives a diamond crystal a giant three dimensional structure. Diamond has no mobile electrons so cannot conduct electricity.

Structure of diamond



Physical properties of diamond

- Diamond is the hardest natural substance known. This is because the carbon atoms are closely parked and are joined by strong covalent bonds.
- Diamond has a very high melting pint because of the strong covalent bond between the carbon atoms.
- It has a very high density (3.5g/cm^3) because of the closely packed carbon atoms.
- Diamond is transparent, sparkling and glitters.

- Diamond does not conduct electricity because it has no mobile electrons.

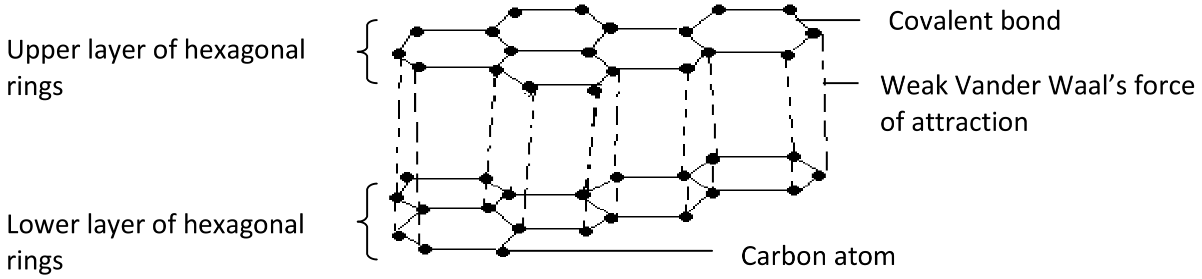
Uses of diamond

1. Diamond is very hard and used as drilling devices, rock borers and glass cutters.
2. Diamond is used jewellery because its sparkling appearance.
3. Diamond is bright and used to make laser beams.

Graphite

Graphite consists of infinite number of carbon atoms each covalently bonded to three other carbon atoms forming hexagonal rings that are arranged in layers. Each layer is a giant two dimensional structure. The different layers are held by weak Vander Waal's forces of attraction, making them to slide over each other thus they are slippery and soft. Some electrons in a layer are mobile making them to conduct electricity.

Structure of graphite



Graphite consists of infinite number of carbon atoms each covalently bonded to three other carbon atoms forming hexagonal rings that are arranged in layers. Each layer is a giant two dimensional structure. The different layers are held by weak Vander Waal's forces of attraction, making them to slide over each other thus they are slippery and soft. Some electrons in a layer are mobile making them to conduct electricity.

Physical properties of graphite

- Graphite conducts electricity because they have free mobile electrons.
- The melting point of graphite is high because of the strong covalent bond between the carbon atoms.
- Graphite is soft and slippery because its layers are held by weak Vander Waals force of attraction.

- Graphite is opaque and dark in color and shinny.

Graphite is less resistant to chemical attack than diamond because of the open spaces between the layers. The density of graphite is 2.3g/cm^3

Uses of graphite

1. Graphite is soft and can mark there fore used to make pencil “leads”. Graphite is mixed with clay to make pencil leads.
2. Graphite is a good conductor of electricity and thus used as electrodes.
3. It is soft and greasy, therefore used as lubricants especially in small bearings like those in dynamos.
4. Graphite is used to make brushes for electric motors.
5. Graphite (black lead) is used as a protective coating on iron to prevent rusting.

When 1g of both diamond and graphite burns in oxygen, they form 3.67g of carbon dioxide only. This shows that they are allotropes.

Amorphous carbon

Amorphous carbon is black and has the lowest density. It is non crystalline and fairly conducts electricity. Amorphous carbon exist in several forms like wood charcoal, animal charcoal, sugar charcoal, lamp black and coke.

i) Wood charcoal

It is formed when wood is heated in limited supply of air (destructive distillation of wood). It is a black porous solid and a very good absorbent of gases.

ii) Animal charcoal

This is made by heating animal refuse and bones in limited supply of air. Animal charcoal has a property of absorbing coloring matter and is used to remove brown color from sugar during its manufacture.

iii) Sugar charcoal

This is a very pure form of carbon made by removing elements of water from sugar. E.g. when cane sugar is dehydrated by concentrated sulphuric acid, sugar charcoal is formed.

iv) Lamp black

This is formed by burning oils (e.g. turpentine, petroleum, kerosene) in limited supply of air. It is used for making ink for printing and shoe polish.

v) Coke

This is an impure form of carbon made by heating coal in the absence of air (destructive distillation of coal). It is used as a reducing agent in the extraction of metals like iron and zinc from their ores. Coke is also used as fuel.

Chemical properties of carbon

All allotropes of carbon have similar chemical properties but different physical properties.

1. Reaction of carbon with oxygen

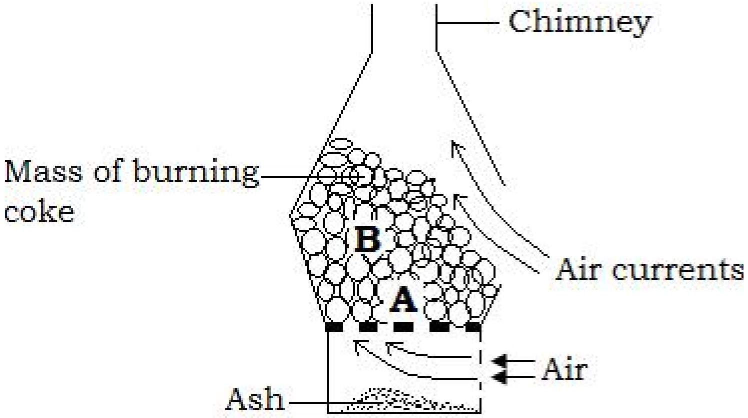
Carbon burns is excess oxygen to for carbon dioxide gas. In the process great heat is generated.

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

When carbon burns in limited amount of oxygen, carbon monoxide forms.

$$C(s) + \frac{1}{2}O_2(g) \longrightarrow CO(g)$$

Reaction of carbon in a deep, brightly glowing coke/coal fire.



In region A, plenty of air is available. Carbon burns to carbon dioxide.

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

In region B, ascending carbon dioxide is reduced by red hot coke to carbon monoxide.

$CO_2(g) + C(s) \longrightarrow 2CO(g)$

At the surface, the hot carbon monoxide burns in air to form carbon dioxide with a flickering blue flame.

$CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g)$

2. Reaction of metal oxides with carbon

Carbon readily removes oxygen from (reduces) oxides of metals lower than it in the reactivity series. Thus it acts as a reducing agent. Examples of oxides reduced by carbon are: zinc oxide, ZnO; lead (II) oxide, PbO; copper (II) oxide, CuO; iron (II) oxide, FeO and iron (III) oxide, Fe_2O_3.

$2ZnO(s) + C(s) \longrightarrow 2Zn(s) + CO_2(g)$
$2PbO(s) + C(s) \longrightarrow 2\ Pb(s) + CO_2(g)$
$2CuO(s) + C(s) \longrightarrow 2Cu(s) + CO_2(g)$

3. Reaction with acids

Carbon reduces nitric acid and sulphuric acid to nitrogen dioxide and sulphur dioxide gases respectively. The carbon itself is oxidized to carbon dioxide gas.

$C(s) + 4HNO_3(aq) \longrightarrow 4NO_2(g) + CO_2(g) + 2H_2O(l)$
$C(s) + 2\ H_2SO_4(aq) \longrightarrow 2SO_2(g) + CO_2(g) + H_2O(l)$

4. Reaction with steam

If steam is blown through red hot charcoal, a mixture of carbon monoxide and hydrogen gases are formed. This mixture is called **water gas.**

$H_2O(g) + C(s) \longrightarrow CO(g) + H_2(g)$

OXIDES OF CARBON

CARBON MONOXIDE (CO)

Occurrence

Carbon monoxide is a poisonous gas and is commonly present in coal gas and other gaseous fuels. It is also produced in car exhaust fumes due to partial combustion of carbon.

Laboratory preparation

This can be done by the action of concentrated sulphuric acid on any of the following compounds.

1. **Methanoic acid (HCOOH)**

Concentrated sulphuric acid is a very strong dehydrating agent and removes water from methanoic acid forming carbon monoxide and water.

$$HCOOH(l) \xrightarrow{\text{Conc. } H_2SO_4} CO(g) + H_2O(l)$$

No heating is required.

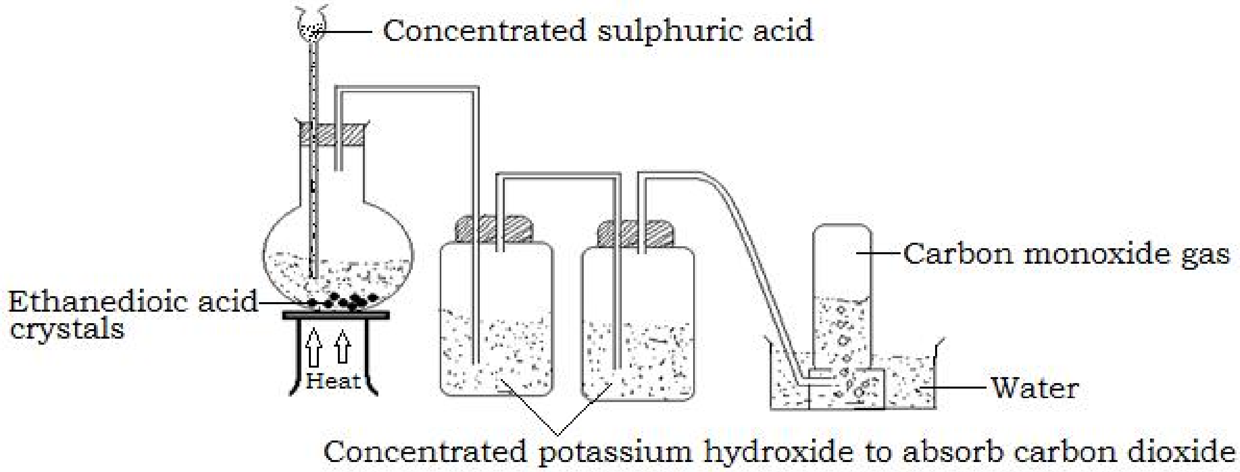
2. **Sodium methanoate (HCOONa)**

Here, concentrated sulphuric acid dehydrates sodium methanoate to form water, sodium hydrogen sulphate and carbon monoxide.

$$HCOONa(s) + H_2SO_4(aq) \longrightarrow NaHSO_4(aq) + CO(g)\ H_2O(l)$$

3. **Oxalic acid (ethanedioic acid-$H_2C_2O_4$)**

Set up of apparatus



Equation

$H_2C_2O_4(s) \xrightarrow{\text{Conc. } H_2SO_4} CO_2(s) + CO(g) + H_2O(l)$

In this preparation, the products formed are passed through concentrated potassium hydroxide solution to remove traces of carbon dioxide.

$2KOH(aq) + CO_2(g) \longrightarrow K_2CO_3(aq) + H_2O(l)$

If the carbon monoxide is required dry, it is passed through a bottle containing conc. Sulhuric acid and collected by use of a syringe or upward displacement of air.

Physical properties of carbon monoxide

- It is colorless, odourless and tasteless.
- It is insoluble in water.
- It is denser than air.
- It is neutral to litmus paper

Chemical properties

a) Reaction with air

Carbon dioxide burns in air with a pale blue flame to produce carbon dioxide.

$CO(g) + ½O_2(g) \longrightarrow CO_2(g)$

b) Reducing action of carbon monoxide

Carbon dioxide is a powerful reducing agent and it reduces the oxides of metals (iron (illustrated in extraction of iron), lead and copper) to respective metals upon heating.

$CuO(s) + CO(g) \longrightarrow Cu(g) + CO_2(g)$
(Black solids) (Brown solids)
$PbO(s) + CO(g) \longrightarrow Pb(s) + CO_2(g)$
(Yellow solids-cold) (Grey solids)

Poisonous nature of carbon monoxide

Blood contains haemoglobin which combines with oxygen to form oxyhaemoglobin. The oxyhaemoglobin transports and supplies the oxygen to all parts of the body. However, when carbon monoxide is inhaled, it combines with haemoglobin to form carboxyhaemoglobin. This stops the blood from absorbing oxygen and may lead to death.

Haemoglobin has a higher affinity for carbon monoxide than oxygen. Carbon monoxide is even more dangerous because it has no colour and smell.

Uses of carbon monoxide

- It is used extensively in the extraction of metals as a reducing agent e.g. extraction of iron in a blast furnace.
- It is used as fuel in form of producer gas and water gas.
- It is used in the manufacture of methanol used in anti freezer mixture in cold countries to prevent ice from forming in car radiators.
- It is used in the manufacture of synthetic petrol.

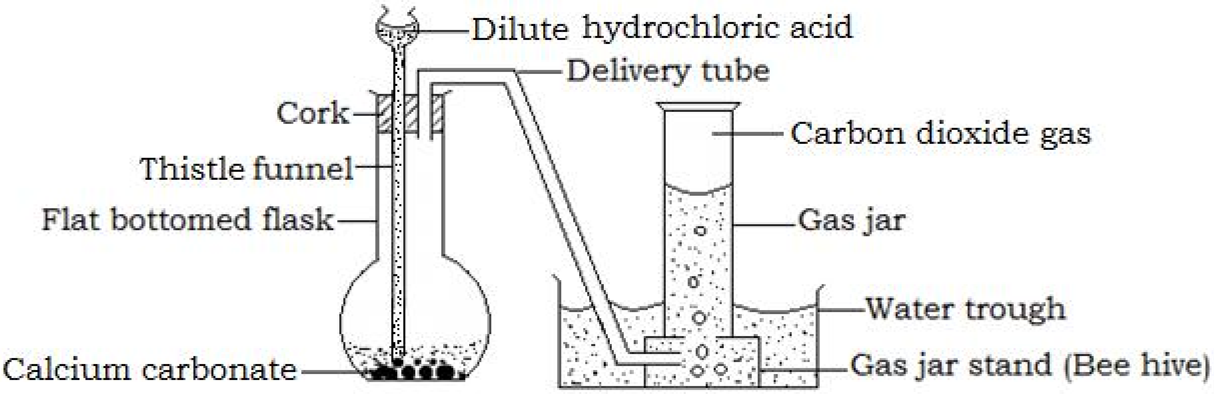
CARBON DIOXIDE (CO_2)

Carbon dioxide occurs in air and occupies 0.03% by volume. Carbon dioxide forms from rocks as a result of volcanic eruption; occurs in mines as “choke damp” and it is always present in natural drinking water because it is slightly soluble in it.

Laboratory preparation

It is prepared by the action of dilute hydrochloric acid on marble chips (a form of calcium carbonate)

Set up of apparatus



Arrange the apparatus as shown above with the calcium carbonate in the flask. Run dilute hydrochloric acid in to the flask through a funnel.

Observation

Effervescence occurs as a colorless gas is given off.

Equation

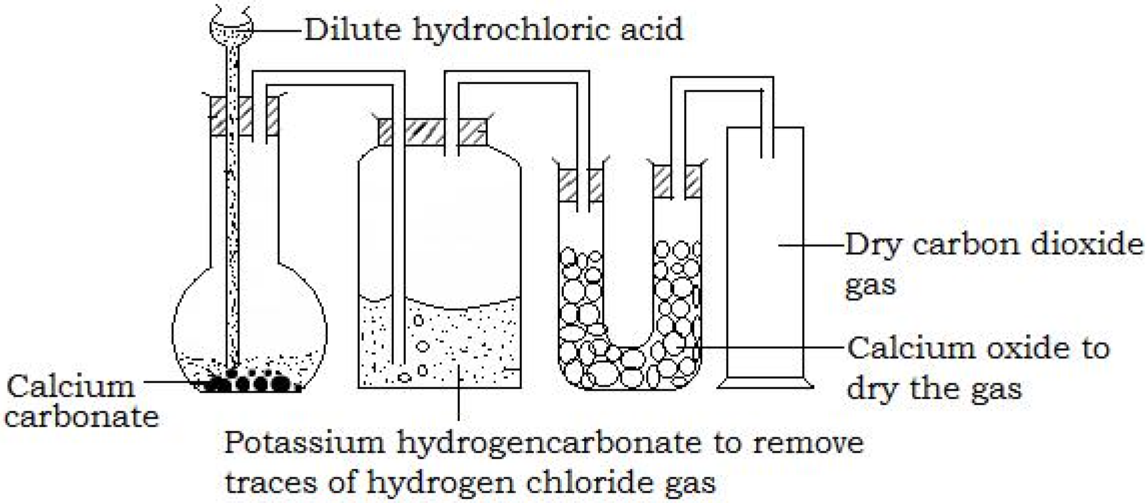
$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

Ionic equation

$CO_3^{2-}(aq) + 2H^+(aq) \longrightarrow CO_2(g) + H_2O(l)$

The gas may be collected over water as it is only slightly soluble in water. If the gas is required dry, it is collected by down ward delivery method.

If a pure dry sample of carbon dioxide is require, the gas is first passed through a wash bottle containing concentrated solution of potassium hydrogen carbonate to remove traces of hydrogen chloride gas (water placed in a wash bottle can also be used). It is then passed through a U tube containing fused calcium chloride to dry the gas. The gas is then collected by down ward delivery method since it is denser than air.



Sulphuric acid is not used in the preparation of carbon dioxide because the salt formed (calcium sulphate) forms a coating on the surface of calcium carbonate and this prevents further reaction between the acid and the calcium carbonate as the calcium sulphate coating is insoluble. The reaction there fore stops prematurely.

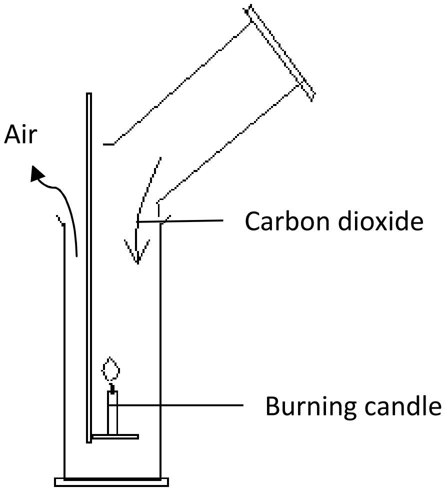
Other reactions that produce carbon dioxide include: fermentation; combustion of hydrocarbons; respiration and action of heat on carbonates and hydrogen carbonates.

Properties of carbon dioxide

Physical properties

- It is colorless, odourless and tasteless.
- It does not burn and does not support burning. Because of its density, it is also used as a fire extinguisher as it displaces oxygen that supports burning.

  Illustration



- It is denser than air that is why it is collected by down ward delivery method. Under high pressure, carbon dioxide is quite soluble in frizzy drinks (effervescent drinks)
- It is slightly soluble in water forming a weakly acidic solution of carbonic acid.

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

- It turns moist blue litmus paper pink indicating that it is weakly acidic.

Chemical properties

1. Carbon dioxide does not support burning. However when a piece of burning magnesium is lowered into a gas jar of carbon dioxide, it continues to burn forming black specks of carbon mixed with white solids of magnesium oxide.

   $2Mg(s) + CO_2(g) \longrightarrow 2MgO(s) + C(s)$

   In this case, magnesium reduces carbon dioxide to carbon ant it self is oxidized to magnesium oxide.

2. Carbon dioxide turns lime water (calcium hydroxide) milky. This is due to the formation of insoluble calcium carbonate (white precipitates)

$Ca(OH)_2(aq) + CO_2(g) \longrightarrow CaCO_3(g) + H_2O(l)$

(White precipitate)

However, when excess carbon dioxide is bubbled, through the solution (lime water), the white precipitates dissolve making the solution to appear clear. This is due to the formation of calcium hydrogen carbonate which is a soluble compound.

$CaCO_3(s) + H_2O(l) + CO_2(g) \longrightarrow Ca(HCO_3)_2(aq)$

3. When carbon dioxide is bubbled through a solution of fairly concentrated sodium hydroxide, the solution remains clear due to the formation of a soluble sodium carbonate.

   $2NaOH(aq) + CO_2(g) \longrightarrow Na_2CO_3(aq) + H_2O(l)$

   This reaction is sometimes used to remove carbon dioxide from a mixture of gases. However, when excess carbon dioxide is bubbled through the above solution, white precipitates appear due to formation of insoluble sodium hydrogen carbonate.

   $Na_2CO_3(aq) + H_2O(l) + CO_2(g) \longrightarrow 2NaHCO_3(s)$

   (White precipitate)

4. When carbon dioxide gas is bubbled through a solution of potassium hydroxide, there is no observable change as the products formed are all soluble.

   $2KOH(aq) + CO_2(g) \longrightarrow K_2CO_3(aq) + H_2O(l)$

   (Soluble)

   $K_2CO_3(aq) + H_2O(l) + CO_2(g) \longrightarrow 2KHCO_3(aq)$

   (Soluble)

Uses of carbon dioxide

- It is used as a fire extinguisher as it does not support burnng and is denser than air.

- It is used in the manufacture of effervescent drinks. This is because the solution of the gas in water has a pleasant taste (the taste of soda water).
- Solid carbon dioxide (dry ice) is used as a refrigerating agent for perishable goods.
- Carbon dioxide is used in the manufacture of baking powder.
- Solid carbon dioxide (dry ice) is fired into the cloud where it cools to form rain. This happens in places where there is unreliable rain fall.

Test for carbon dioxide

When carbon dioxide gas is passed through a solution of calcium hydroxide (lime water), the solution becomes milky. This is due to the formation of calcium carbonate which is insoluble.

$$Ca(OH)_2(aq) + CO_2(g) \longrightarrow CaCO_3(s) + H_2O(l)$$

(White precipitate)

CARBONATES AND HYDROGEN CARBONATES

These are salts derived from carbonic acid.

Properties

1. **Solubility**

a) A part from potassium, sodium and ammonium carbonates, all other carbonates are insoluble in water.

b) All solid hydrogen carbonates are soluble in water except sodium hydrogen carbonate which is only slightly soluble. Hydrogen carbonates of metals lower than magnesium in the reactivity series do not exist. Hydrogen carbonates of calcium and sometimes magnesium only exist in solution form.

2. **Action of heat on carbonates and hydrogen carbonates**

a) Carbonates of sodium, potassium and lithium are not decomposed by heat. However if they are hydrated, they only lose their water of crystallization. E.g.

$$Na_2CO_3.10H_20(s) \longrightarrow Na_2CO_3(s) + 10H_2O(l)$$

(White crystals) (White powder)

b) All the other metal carbonates decompose when heated to form the oxide of the metal and a colorless gas that turns lime water milky (carbon dioxide) e.g.

$CuCO_3(s) \longrightarrow CuO(s) + CO_2(g)$

(Green powder) (Black powder)

$ZnCO_3(s) \longrightarrow ZnO(s) + CO_2(g)$

(White powder) (Yellow-hot; white-cold)

c) Ammonium carbonate sublimes when heated and forms ammonia, carbon dioxide and water vapor.

$(NH_4)_2CO_3(s) \longrightarrow 2NH_3(g) + H_2O(g) + CO_2(g)$

d) All hydrogen carbonates decompose to give corresponding carbonates, carbon dioxide gas and water vapor. E.g.

$2NaHCO_3(s) \longrightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$

3. Action of dilute acids

All carbonates and hydrogen carbonates react with dilute acids to liberate carbon dioxide gas, water and corresponding salts. E.g.

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

$H_2SO_4(aq) + 2NaHCO_3(s) \longrightarrow Na_2CO_3(aq) + CO_2(g) + H_2O(l)$

N.B. Dilute sulphuric acid reacts with calcium carbonate and lead (II) carbonate at a very slow rate as the resulting salts formed (calcium sulphate and lead(II) sulphate) are insoluble and there fore tend to form coatings around the carbonates inhibiting further reactions between the carbonates and the acid.

The reaction between dilute hydrochloric acid and lead (II) carbonate also forms an insoluble salt (lead (II) chloride) which also forms coating stopping further reactions.

Dilute nitric acid reacts with all carbonates to form soluble nitrate salts.

Testing for carbonates and hydrogen carbonates

a) If the sample is a solid

Add a little dilute acid to the solid substance under test.

Observation

Effervescence occurs and a colorless gas that turns lime water milky (carbon dioxide) is given off.

Conclusion

This shows that a carbonate or hydrogen carbonate is present.

b) Carbonate in solution form

To 2cm^3 of test solution add 3 drops of lead (II) nitrate solution followed by excess dilute nitric acid.

Observation

White precipitate is formed and the precipitate dissolves in excess nitric acid with effervescence.

N.B. Sometimes barium nitrate solution or barium chloride solution may be used in place of lead (II) nitrate solution.

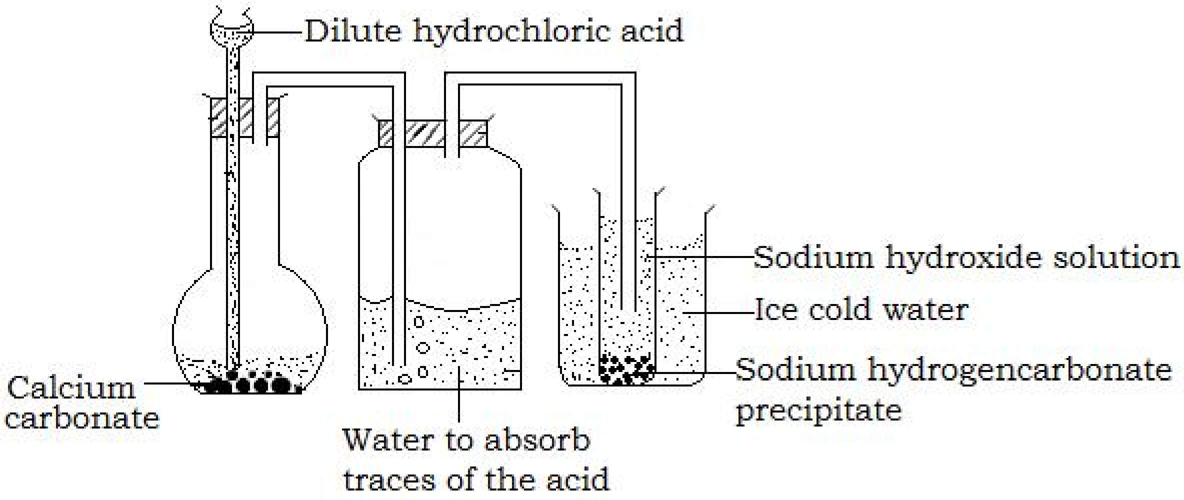
$$Na_2CO_3(s) + Pb(NO_3)_2(aq) \longrightarrow PbCO_3(s) + 2NaNO_3(aq)$$

$$PbCO_3(s) + HNO_3(aq) \longrightarrow Pb(NO_3)_2(aq) + H_2O(l) + CO_2(g)$$

Sodium carbonate

Laboratory preparation

Set up



Procedure

- Bubble carbon dioxide through a fairly concentrated sodium hydroxide solution until in excess. White precipitate of sodium hydrogen carbonate is formed according to the equation below

$$2NaOH(aq) + CO_2(g) \longrightarrow Na_2CO_3(aq) + H_2O(l)$$

$$Na_2CO_3(aq) + H_2O(l) + CO_2(g) \longrightarrow \underset{\text{(White precipitate)}}{2NaHCO_3(s)}$$

- The white precipitate of sodium hydrogen carbonate is filtered off, washed and dried.
- The sodium hydrogen carbonate is heated strongly until no further water vapor and carbon dioxide are given off. This leaves a white powder of calcium carbonate.

$$2NaHCO_3(s) \longrightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$$

Commercial preparation of sodium carbonate (Solvay process)

Very concentrated brine (28% sodium chloride) is saturated with ammonia gas in a tower to form ammonia gas in a tower to form ammoniacal brine. The ammoniacal brine is run downwards from the top of the tower while carbon dioxide (formed from decomposition of calcium carbonate) is forced to rise up the tower from the base of the tower.

The tower is fitted with perforated mushroom shaped baffles at intervals that delay the flow of ammoniacal brine and also offer surface for the reaction.

The ammoniacal brine reacts with carbon dioxide to form sodium hydrogen carbonate as precipitates since it is not very soluble in water. Precipitation is assisted by cooling the lowest third of the chamber.

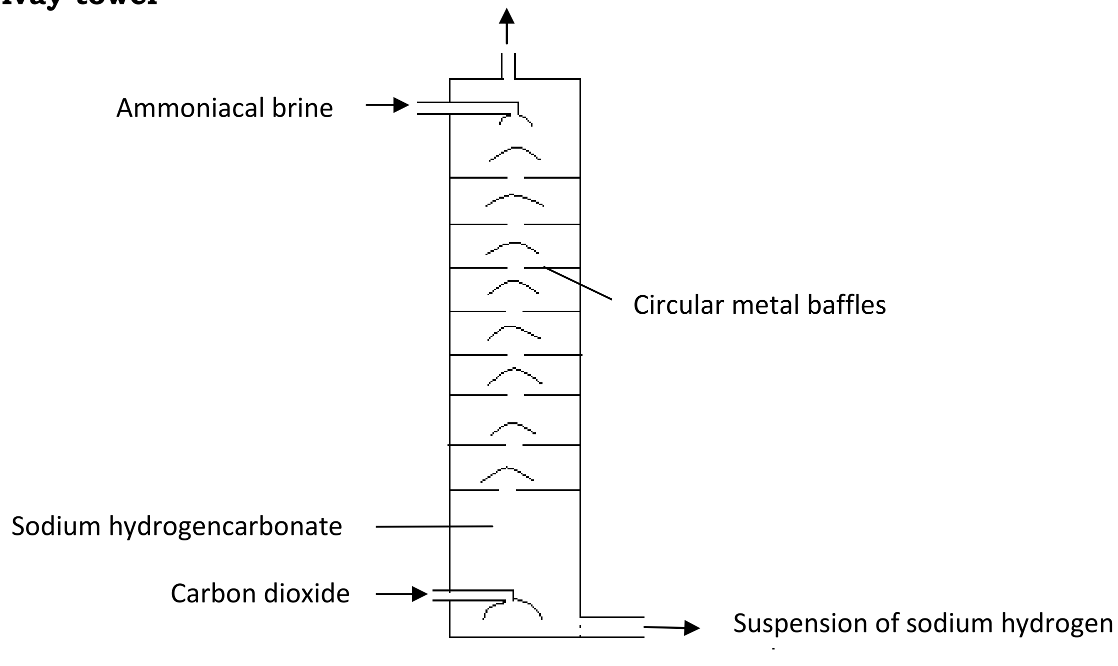
$$NaCl(aq) + NH_4OH(aq) + CO_2(g) \longrightarrow NaHCO_3(s) + NH_4Cl(aq)$$

Sodium hydrogen carbonate is filtered from the white sludge at the base of the tower and washed to remove ammonium compounds. The sodium hydrogen carbonate is the heated to form sodium carbonate.

$$2NaHCO_3(s) \longrightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$$

The anhydrous sodium carbonate formed ha a wide market.

The Solvay tower



If crystalline form (washing soda) is required, the anhydrous solid is dissolved in hot water, crystallization takes place as the solution cools. The crystals are removed and allowed to dry.

$$Na_2CO_3(aq) + 10H_2O(l) \longrightarrow Na_2CO_3.10H_2O(s)$$

Sodium carbonate decahydrate, $Na_2CO_3.10H_2O$ are large translucent crystals. When the crystals are exposed to air, they lose mass and become coated with fine powder which makes is opaque. Each molecule of washing soda gives to the atmosphere 9 molecules of water of crystallization.

$$Na_2CO_3.10H_2O(s) \longrightarrow \underset{\text{(Sodium carbonate monohydrate)}}{Na_2CO_3.H_2O(s)} + 9H_2O(g)$$

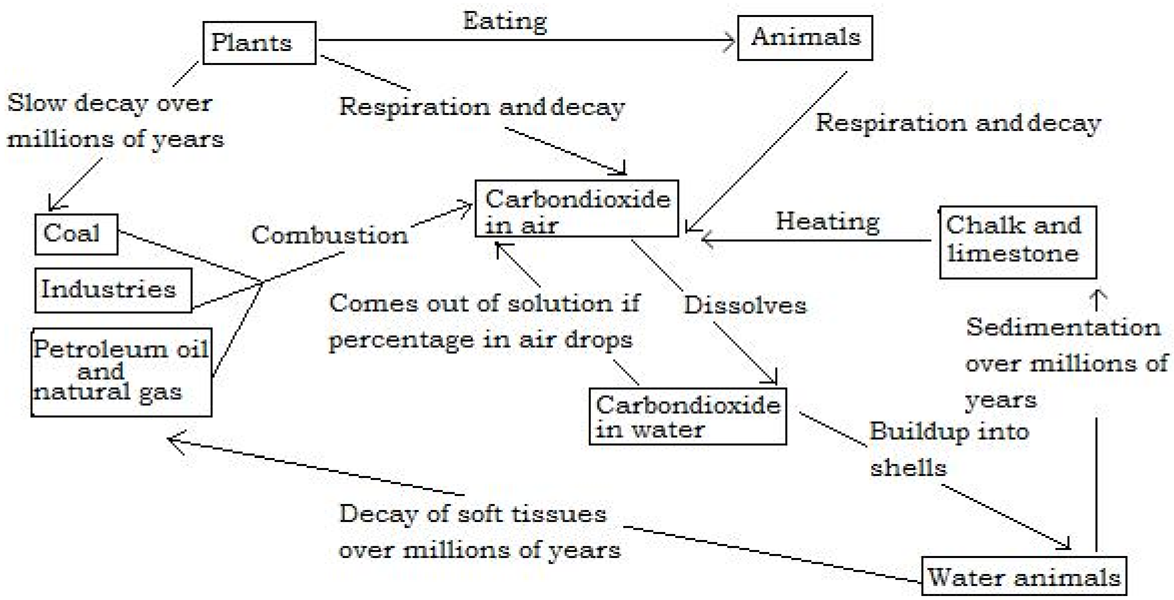
Such lose of water of crystallization to the atmosphere is termed as efflorescence. The substance that loses its water of crystallization is known as an efflorescent substance.

Uses of sodium carbonate

1. Manufacture of glass. Ordinary bottle glass is made by fusing together sodium carbonate, calcium carbonate, silicon dioxide (sand) and a little carbon (reducing agent). Broken glasses are added to assist fusion.
2. Manufacture of water glass that is used to preserve eggs, used in fire proofing and production of cement.
3. Sodium carbonate is used in the manufacture of soap powders.
4. Sodium carbonate in used in domestic water softening. Calcium ions (Ca^{2+}) which is the principal cause of hardness in water is precipitated from the water as calcium carbonate by adding sodium carbonate.

The carbon cycle

The balance of processes which give out carbon dioxide and those which use carbon dioxide is called the **carbon cycle.** Summary of the carbon cycle is given below.



Carbon dioxide in the atmosphere is added from: respiration of plants and animals; decay/ decomposition of plants and animals; heating of lime stone (to give quick lime) in lime kilns; water (if the percentage in air drops); and combustion of coal (formed as a result of slow decomposition of plants and animals over millions of years), petroleum oil and natural gas.

The processes that remove carbon dioxide from the atmosphere are: photosynthesis by green plants; and dissolution in water.

Animals obtain carbon from plants by feeding on food such as starch made by plants..

Effects of carbon dioxide on atmospheric temperature

The sun emits radiations that pass through the atmosphere of the earth with little absorption and warms up the surface of the earth (ground). The warm surface of the earth (ground) reflects back the radiations inform of infrared radiations. The infrared radiation is absorbed by gases like carbon dioxide and methane which radiate some heat back to the ground leading to warming up of the earth. The warming up of the earth is known as **green house effect**. The gases like carbon dioxide and methane that cause the warming up of the earth are referred to as **green house gases.**

When more green house gases are released into the atmosphere, much heat accumulates and this leads to general rise in temperature of the world. The general rise in temperature of the world is known as **global warming**.

Sample questions on Carbon and its compounds

1. *Define the terms allotrope and allotropy. Name two crystalline forms of the allotropes of carbon and describe their structures. Outline three differences between the two allotropes. Mention two uses of each allotrope. How do you prove that the allotropes consist solely of carbon?*
2. *Amorphous carbon exists in several forms name them and briefly explains how each form is obtained and give at least one use of each form of amorphous carbon.*
3. *With the aid of equations describe the reactions of carbon with oxygen, acids, steam and metallic oxides.*
4. *Describe carefully how you would prepare and collect carbon monoxide free from carbon dioxide. Give three ways in which this gas is different from carbon dioxide. Describe the reducing action of carbon monoxide on copper(II) oxide. Explain briefly how water gas and producer gas are obtained. How can exhaust fumes of cars act as poison?.*
5. *By means of a labeled drawing, and equation, show how a sample of carbon dioxide can be made and collected in the laboratory. Describe briefly without using diagrams how you can convert carbon dioxide to carbon monoxide. Mention two large scale uses of carbon dioxide.*
6. *Explain with equations the observations made when carbon dioxide is bubbled through calcium hydroxide (lime water) and sodium hydroxide until in excess.*
7. *Describe how sodium carbonate is prepared in the laboratory and in large scale from the Solvay process. Equations and drawings must be used to illustrate.*

EFFECTS OF ELECTRICITY ON SUBSTANCES

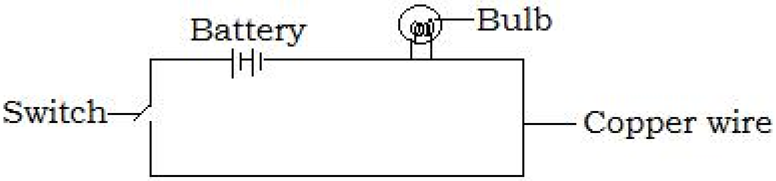
Electricity is a form of energy produced as a result of flow of electrons through materials.

Investigation of electrical conductivity through solid materials

Examples of such materials include: copper wire, zinc wire, plastic, graphite, rubber.

Procedure

Connect the copper wire to the batteries through the bulb and the switch as shown below.



Results

The bulb produced light on complete connection with copper and zinc.

Conclusion

Copper and zinc wires conduct electricity and they are called **conductors**.

A conductor is a substance in solid form which can conduct electricity. Examples include; all metals and graphite (the only non metal that can conduct electricity.

When the above experiment was repeated using rubber and plastics, the bulb did not light indicating that they do not conduct electricity and are referred to as **insulators** or **non conductors**.

A non conductor is a substance in solid form that does not conduct electricity. Examples are all non metal except graphite.

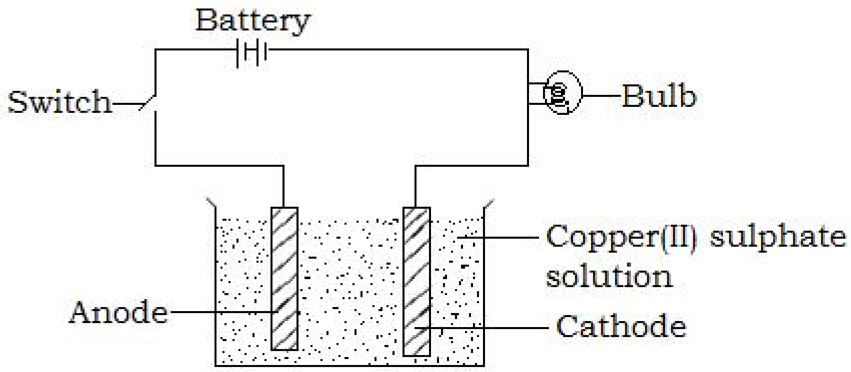
Metals conduct electricity because they have delocalized, free or mobile electrons but non metals do not have these delocalized electrons as they are all locked up in bond formation

Investigation of electrical conductivity through liquid substances in solution

Examples: ethanol, urea, hydrochloric acid, copper (II) sulphate, ethanoic acid, water, ammonium hydroxide.

Procedure

1. Put the liquid under investigation in an electrolytic cell.
2. Dip two rods in the liquid which can either be a metal or carbon(graphite) called electrodes
3. Connect the electrodes using a conductor to a bulb via a switch to the source of power (the batteries) as shown below.



4. Close the switch.
5. Repeat the experiment with hydrochloric acid, ethanol, ethanoic acid, water, urea, ammonium, ammonium hydroxide.

Results

When ethanol and urea were used there was no light produced indicating that they do not conduct electricity, they are there fore called **non electrolytes**.
When ammonium hydroxide and ethanoic acid were used, the bulb produced a dim light indicating that they weakly conduct and are there fore **weak electrolytes**.

When copper sulphate solution and hydrochloric acid were used, the bulb produced bright light indicating that they strongly conduct electricity and are **strong electrolytes**.

ELECTROLYSIS

This is the decomposition of a substance in solution form or molten form (electrolyte) as a result of passage of electric current. The decomposition of the electrolyte takes place at the electrodes.

Definitions of common terms used in electrolysis

1. Electrolytes

This is a substance in solution form or molten state that can conduct electricity. Electrolytes can be categorized as strong, weak or non electrolytes.

a) Strong electrolyte

This is a substance in solution form or molten state that ionizes completely and can easily conduct electricity. The electrolyte decomposes fully during electrolysis. Examples include, all mineral acids, alkalis, ionic compounds.

Sodium chloride; $NaCl(aq) \longrightarrow Na^+(aq) + Cl^-(aq)$

Hydrochloric acid; $HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$

b) Weak electrolytes

This is a substance in solution form which is only slightly ionized (partially ionized). The electrolyte is only partially decomposed by the electric current. Most of the ions of the electrolyte remain as un ionized ions or molecules.

Examples include:

All weak alkalis like ammonium hydroxide; $NH_4OH(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

All solutions of weak acids like ethanoic acid; $CH_3COOH_{(aq)} \rightleftharpoons CH_3COO^-(aq) + OH^-(aq)$

Impure water.

c) Non electrolyte

Is a substance in solution form or aqueous state that doesn’t conduct electricity. This substance is not decomposed at the electrodes. Example include: all covalent compounds, like pure water, benzene, methyl benzene, petrol and diesel.

2. Electrodes

These are rods or plates or poles of conductors at which electrons enter and leave the electrolyte. The electrodes are either cathodes or anodes.

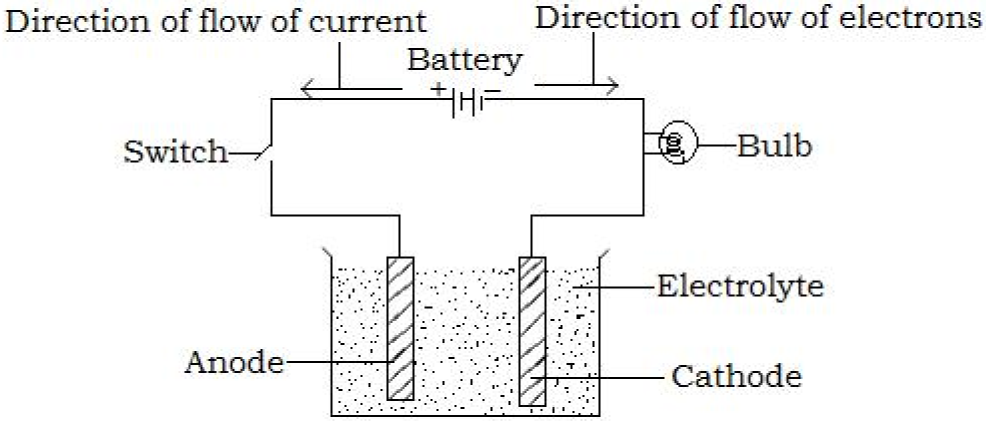
a) Anode

This is the positive electrode at which the electrons leave the electrolyte. Or is the positive electrode at which electrons enter the external circuit. It is normally connected to the positive terminal of the battery.

b) Cathode

This is the negative electrode at which the electrons enter the electrolyte or is the negative electrode at which the electrons leave the external circuit. It is connected to the negative end of the battery.

Simple electrolytic cell



In an electrolytic cell the electrolyte has to be in solution form or molten state as the ions have to be free to move so as to conduct electricity. Salts such as sodium chloride, and lead (II) bromide do not conduct electricity in solid state because the ions are held together by strong electrostatic forces of attraction and are not free and mobile. However, when the salts are melted or dissolved in water, the electrostatic forces are broken down and the ions become free and mobile and so conduct electricity.

IONIC THEORY

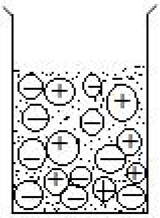
Ionic theory was put forward to explain the phenomenon of electrolysis. According to ionic theory, electrolytes are believed to contain electrically charged particles called **ions.** The ions can be positively charged (cations) and are obtained from metals, hydrogen and ammonium or negatively charged (anions) and are obtained from non metals e.g.Cl^-,NO_3^-,SO_4^{2-},O^{2-}. During electrolysis, the cations are attracted to the negative electrode (cathode) and the anions are attracted to the positive electrode (anode)

Examples of electrolytes and the ion produced.

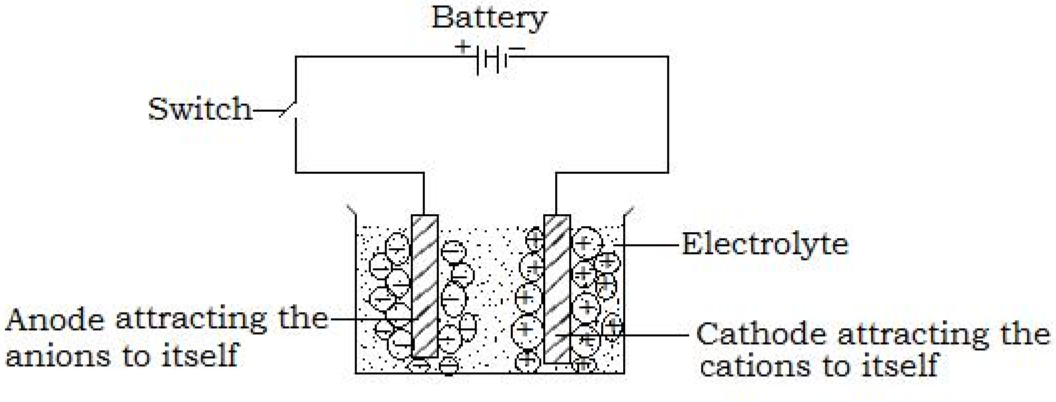
Compound	**Formula**	**Ions produced**
Sulphuric acid	H_2SO_4	H^+ and SO_4^{2-}
Sodium chloride	NaCl	Na^+ and Cl^-
Sodium hydroxide	NaOH	Na^+ and OH^-
Copper (II)sulphate	$CuSO_4$	Cu^{2+} and SO_4^{2-}
Lead(II) bromide	$PbBr_2$	Pb^{2+} and Br^-
Copper (II) chloride	$CuCl_2$	Cu^{2+} and Cl^-

Explanation of electrolysis by ionic theory

When current is not passed through an electrolyte, the ions are wandering randomly in solution.



When current is passed through the solution, the cathode attracts the to itself and the anode attracts to it self the anions.



When the cations reach the cathode (negative electrode), they stick to it, gain electrons and become ordinary atoms. e.g.

$Na^+(aq) + e^- \longrightarrow Na(s)$

When the anions reach the anode (positive electrode), they give away the electrons and become ordinary atoms.e.g.

$Cl^-(aq) - e^- \longrightarrow Cl(g)$

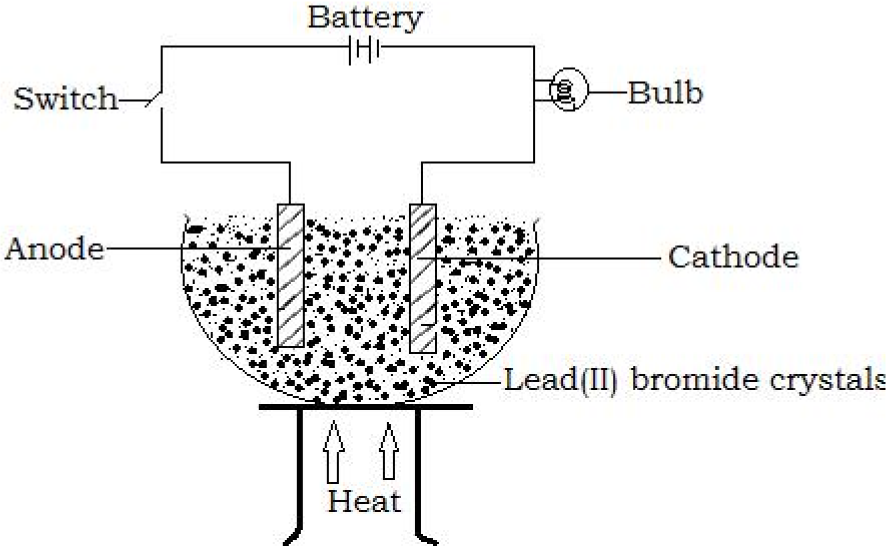
The atoms then combine to form molecules

$Cl(g) + Cl(g) \longrightarrow Cl_2(g)$

Experiment to show that Lead (II) bromide only conducts electricity in molten form

Electrolysis of leads (II) bromide

The electrolysis is done using carbon electrodes (graphite) as shown below.



When carbon electrodes are dipped into solid Lead (II) bromide crystals and the circuit completed, the bulb did not light indicating that there was no conduction because the ions responsible for conduction were locked up in the solid crystal. When heated and molten liquid formed, there was conduction as the bulb lit because the ions were free to move and thus conducted.

Ions present in Lead (II) bromide

$PbBr_2(aq) \longrightarrow Pb^{2+}(aq) + 2Br^-(aq)$

Lead ions, Pb^{2+} migrate to the cathode where it gains electrons and it is discharged to form **silvery grey solids** of lead. Thus equation at the cathode is

$Pb^{2+}(aq) + 2e^- \longrightarrow Pb(s)$

Bromine ions migrate towards the anode and are discharged by losing electrons to form **red/brown vapor or liquid.** Thus, equation at the anode is

$2Br^-(aq) - 2e^- \longrightarrow Br_2(g)$

Preferential or selective discharge of ions

When two or more ions of similar charges reach the electrode, one is preferentially selected for discharge and the selection depends on the following factors.

- The position of the ion in the electrochemical series
- The concentration of the electrolyte
- The nature of electrodes

a) The position of metal or radical in the reactivity series

The ion that is lower in the electrochemical series is selected for discharge in preference to one above it.

Cations (migrate to cathode)		**Anions (migrate to anode)**	
K^+		SO_4^{2-}	
Ca^{2+}		SO_4^{2-} NO_3^-	
Na^+	Increasing ease of discharge	NO_3^- Cl^-	Increasing ease of discharge
Mg^{2+}		Cl^- Br^-	

Al^{3+}	I^-
Zn^{2+}	I^- OH^-
Fe^{2+}	
Pb^{2+}	
H^+	
Cu^{2+}	
Ag^+	

Example: Consider the decomposition of copper(II)sulphate solution

$CuSO_4(aq) \longrightarrow Cu^{2+}(aq) + SO_4^{2-}(aq)$

$H_2O(aq) \rightleftharpoons H^+(aq) + OH^-(aq)$

At the cathode, the ions are copper ions (Cu^{2+}) and hydrogen ions (H^+)

At the anode, the ions are, sulphate ions (SO_4^{2-}) and hydroxyl ions (OH^-)

At the cathode, Cu^{2+} is preferably discharged in preference to hydrogen since copper is below hydrogen in the electrochemical series.

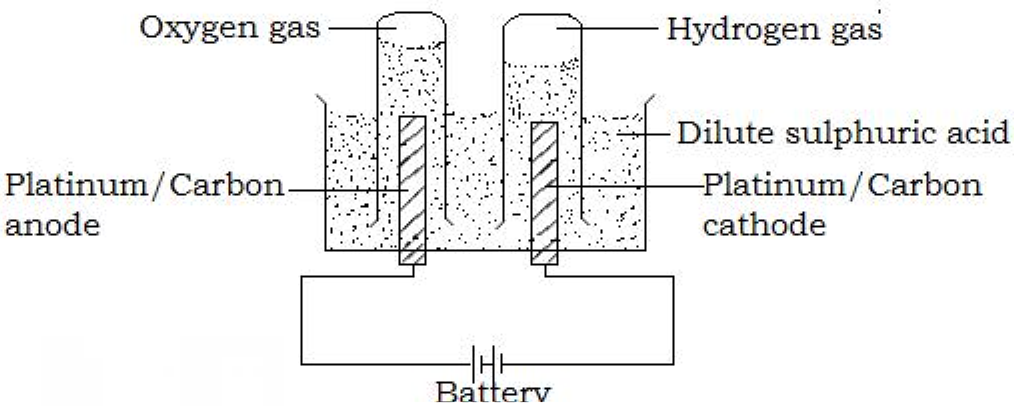
At the anode, OH^- is preferentially discharged since its below sulphate (SO_4^{2-}) in the reactivity series.

Electrolysis of some electrolytes

1. Electrolysis of dilute sulphuric acid (electrolysis of water)

The electrolysis is done by use of platinum or carbon rods for both electrodes. During the electrolysis, 2 volumes of hydrogen is produced at the cathode and one volume of oxygen is formed at the anode. Total acidity of the products remains the same as the products formed are elements of water.

Drawing of apparatus



Ions present; (H^+, SO_4^{2-}, OH^-)

$$H_2SO_4(aq) \longrightarrow H^+(aq) + SO_4^{2-}(aq)$$

$$H_2O(aq) \rightleftharpoons H^+(aq) + OH^-(aq)$$

At cathode

Hydrogen ions migrate to the cathode where they are discharged by receiving electrons from the cathode and form atoms. The hydrogen atoms pair up to form hydrogen gas (bubbles of a colorless gas that burns with a pop sound).

Equation at the cathode

$$2H^+(aq) + 2e^- \longrightarrow H_2(g)$$

At the anode

Both the sulphate ions and hydroxyl ions migrate to the anode. The hydroxyl ions are discharged in preference to the sulphate ions as it is below the sulphate ions in the electrochemical series.

Equation at the anode

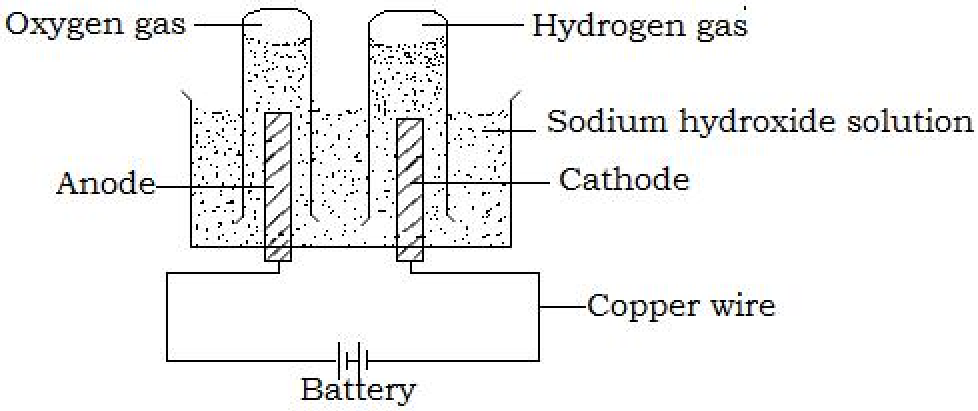
$$4OH^-(aq) - 4e^- \longrightarrow 2H_2O(l) + O_2(g)$$

Overall equation

$$4H^{+}(aq) + 4OH^{-}(aq) \longrightarrow 2H_2(g) + 2H_2O(l) + O_2(g)$$

2. Electrolysis of dilute sodium hydroxide solution using carbon electrodes

Setup of the apparatus



Ions present

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

$$H_2O(aq) \rightleftharpoons H^{+}(aq) + OH^{-}(aq)$$

At the cathode

Hydrogen ions are discharged in preference to sodium ions because it is below it in the electrochemical series. Therefore, at the cathode, bubbles of a colorless gas that burns with a pop sound is produced.

Equation

$$2H^{+}(aq) + 2e^{-} \longrightarrow H_2(g)$$

At the anode

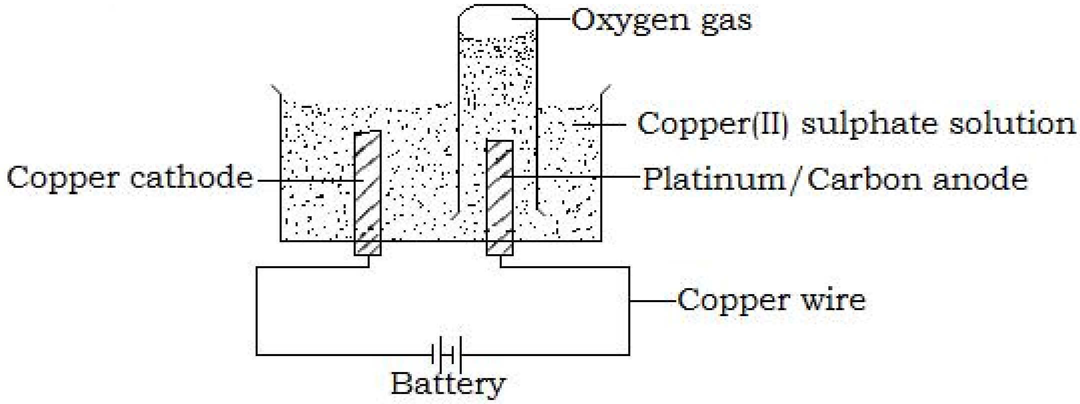
At the anode, the hydroxyl ions are discharged giving off bubbles of a colorless gas that relights a glowing splint.

Equation

$$4OH^{-}(aq) - 4e^{-} \longrightarrow 2H_2O(l) + O_2(g)$$

3. Electrolysis of copper (II) Sulphate solution using copper cathode and platinum or graphite anode

Set up of apparatus



Ions present

$$CuSO_4(aq) \longrightarrow Cu^{2+}(aq) + SO_4^{2-}(aq)$$

$$H_2O(aq) \rightleftharpoons H^+(aq) + OH^-(aq)$$

At the cathode

Both copper ions and hydrogen ions migrate to the cathode but the copper ions being lower than hydrogen in the electrochemical series, it's discharged in preference to hydrogen ions. There fore at the cathode, the copper ions gain electrons and are deposited as brown solids of copper.

Equation

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

At the anode

Both sulphate and hydroxyl ions migrate to the anode. The hydroxyl ions being lower than the sulphate ions in the electrochemical series are discharged forming water and hydrogen as final products.

Equation

$$4OH^-(aq) - 4e^- \longrightarrow 2H_2O(l) + O_2(g)$$

Note

i. The blue color of the copper (II) Sulphate fades away with time as the copper ions which are responsible for the blue color are being discharged and deposited at the cathode as brown copper metal.

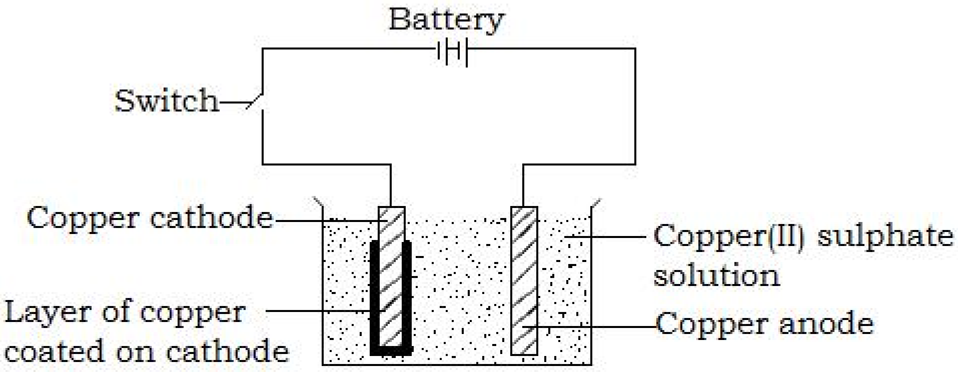
ii. The discharge of the hydroxyl ions at the anode disturbs the ionic equilibrium of water; therefore more water ionizes to restore this equilibrium. The excess hydrogen ions produced combines with the undischarged sulphate ions forming sulphuric acid which makes the solution around the anode acidic.

b) The nature of electrode

Different electrodes for a given electrolyte may cause different products to be formed at the electrodes.

4. Electrolysis of Copper(II) sulphate solution using copper electrodes

Set up of the apparatus



Ions present

$$CuSO_4(aq) \longrightarrow Cu^{2+}(aq) + SO_4^{2-}(aq)$$

$$H_2O(aq) \rightleftharpoons H^+(aq) + OH^-(aq)$$

At the cathode

Both copper ions and hydrogen ions migrate to the cathode but copper ions are discharged in preference to hydrogen ions since it's below it in the electrochemical series.

Equation

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

At the anode

Both sulphate and hydroxide ions migrate to the anode but non is discharged. Instead, the copper anode goes into solution i.e. it dissolves to form copper ions. This process is called **electrode ionization**. Such a process in favored in this case as it requires less energy than the discharge of ions.

Equation

$$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$$

Overall equation

$$Cu^{2+}(aq) + Cu(s) \longrightarrow Cu^{2+}(aq) + Cu(s)$$

Note

i. During the experiment, the anode loses mass and the cathode gains mass. The loss in mass at the anode equals the gain in mass at the cathode. The change in mass at either electrode is proportional to the quantity of electricity passed through the electrolyte.

ii. The intensity of the blue color of copper (II) sulphate remains constant as the process is a mere transfer of the copper ions from anode to cathode. i.e. the copper from the anode goes into solution as ions to replace the lost copper ions at the cathode.

iii. Overall concentration of the electrolyte remains constant.

5. Electrolysis of dilute sodium chloride using a mercury cathode and graphite/platinum anode

Ions present

$NaCl(aq) \longrightarrow Na^+(aq) + Cl^-(aq)$

$H_2O(aq) \rightleftharpoons H^+(aq) + OH^-(aq)$

At the anode

Both the chloride and hydroxyl ions migrated to the anode but the hydroxyl ions being lower than the chloride ions in the reactivity series, the hydroxyl ions are discharged by losing electrons in preference to chloride ions. Therefore bubbles of colorless gas that relights a glowing splint is observed at the anode.

Equation

$$4OH^-(aq) - 4e^- \longrightarrow 2H_2O(l) + O_2(g)$$

At the cathode

Both sodium and hydrogen ions move to the cathode. In this case, sodium ions are discharged despite the fact that hydrogen ions are lower than it in the electrochemical series. This is because the process requires less energy and the sodium atoms produced form an amalgam with mercury.

Equation at cathode

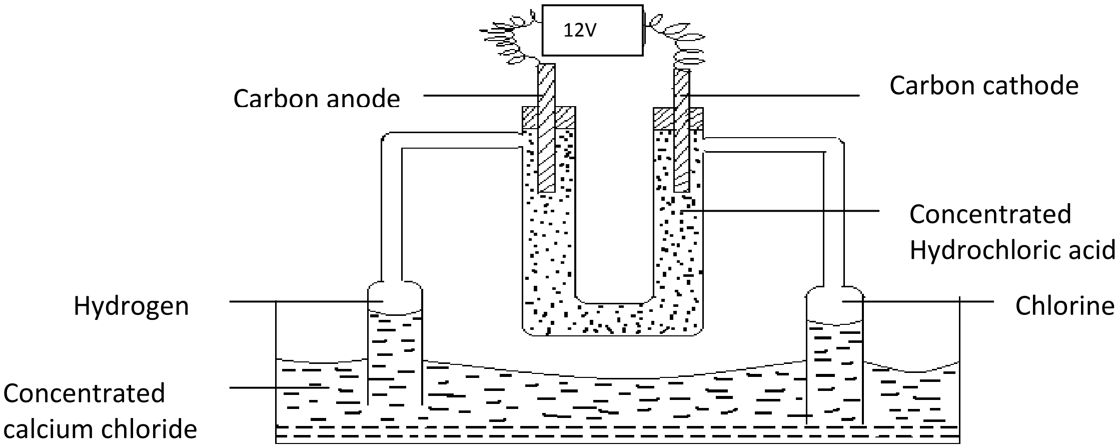
$$Na^+(aq) + e^- \longrightarrow Na(s)$$

c) Concentration

Increase in concentration of ions tends to promote its chance of being discharged. E.g. if concentrated Hydrochloric acid is used.

6. Electrolysis of concentrated hydrochloric acid using graphite electrodes

The electrolysis is done in the set up of apparatus below.



Ions present

$$HCl_{(aq)} \longrightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

$$H_2O(aq) \rightleftharpoons H^+(aq) + OH^-(aq)$$

At the cathode

Hydrogen ions migrate to the cathode where they are discharged forming bubbles of a colorless gas that burns with a pop sound (hydrogen gas).

Equation

$$2H^+(aq) + 2e^- \longrightarrow H_2(g)$$

At the anode

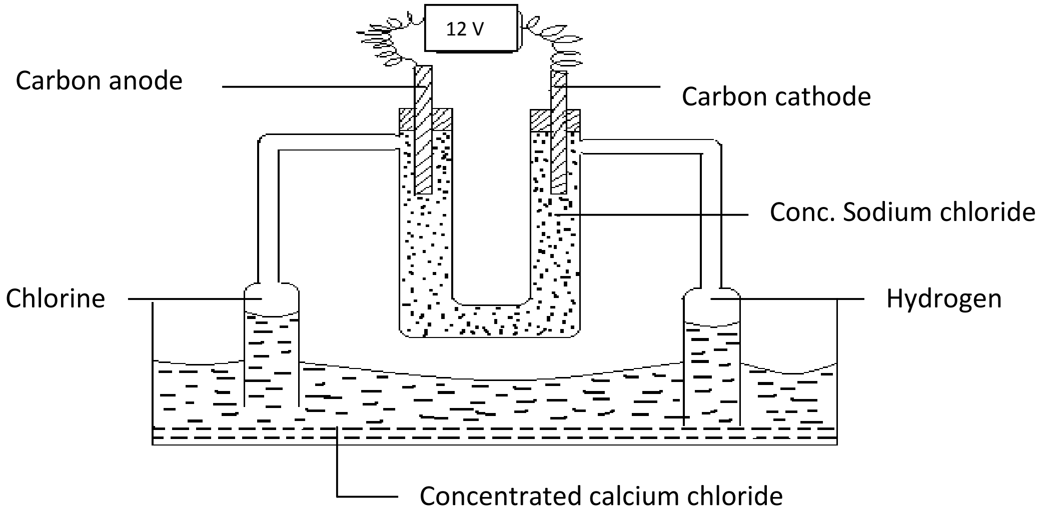
Both the chloride and the hydroxyl ions migrate to the anode but the hydroxyl ions despite them being lower than chloride ions in the electrochemical series are not discharged. Instead the chloride ions are discharged since they are present in a much higher concentration. Therefore, a green-yellow gas that bleaches damp litmus paper is observed indicating that the gas is chlorine gas.

Equation

$2Cl^-(aq) - 2e^- \longrightarrow Cl_2(g)$

7. Electrolysis of concentrated sodium chloride solution using graphite electrodes

Set up of the apparatus



Ions present

$NaCl(aq) \longrightarrow Na^+(aq) + Cl^-(aq)$

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

At the anode

Both the chloride and hydroxyl ions migrate to the anode but chloride ions being present in a much higher concentration are discharged in preference to the hydroxyl ions. There fore, bubble of a green yellow gas that bleaches damp litmus paper (chlorine) is observed.

Equation

$$2Cl^-(aq) - 2e^- \longrightarrow Cl_2(g)$$

At the cathode

Both sodium and hydrogen ions migrate to the cathode. Hydrogen being lower in the electrochemical series are discharged in preference to sodium ions. Therefore, bubbles of a colorless gas that burns with a pop sound is produced at the cathode.

Equation

$$2H^+(aq) + 2e^- \longrightarrow H_2(g)$$

LAWS OF ELECTROLYSIS

The laws of electrolysis express the quantitative results of electrolysis and were first stated by faraday. These laws assert that, the amount of substance (in moles) liberated during electrolysis depends on,

i) The time of passing the steady current

ii) The magnitude of the steady current passed

iii) The charge on the ion of the element

Faraday's first law of electrolysis

It states that, the amount (mass) of substance liberated/deposited at electrodes during electrolysis is directly proportional to the quantity of electricity passed.

The quantity of electricity passed is the product of time (seconds) and current (ampere).

$$\underset{\text{(Coulombs)}}{\text{Quantity of electricity}} = \underset{\text{(amperes)}}{\text{Current}} \times \underset{\text{(seconds)}}{\text{Time}}$$

i.e. $Q=It$ where Q is quantity of electricity(C), I is current(s) and t is time(s)

A **Coulomb(C**) is the quantity of electricity passed when a current of one ampere flows in one second.

A **Faraday** is the quantity of electricity needed to discharge or deposit one mole of ions or substance. A faraday is equivalent to one mole of electrons

$1F=96500C$

$1F=1$ mole of electrons

Examples

1. Calculate the quantity of electricity used when a current of 4A is passed through an electrolyte for 45 minutes.

Solution

$Q=It$

Given that I=4A

$t= (45x60)\ s \quad =2700s$

$Q= (4x2700) =10800C$

2. A steady current of 2.5A was passed through sodium sulphate solution for 1 hour, 40 minutes and 20 seconds. Calculate
   - i) The quantity of electricity passed
   - ii) The number of Faraday used (1f= 96500C)

Solution

i) $Q=It$

Given that I=2.5 A

$t= (1x60x60+4x60+20)\ s =6020s.$

$Q=2.5x6020 =15050C$

ii) 96500C is equivalent to 1 Faraday

1C is equivalent to (1/96500) Faradays

15050C is equivalent to (1x15050/96500)C

=0.156Faradays.

Faraday's second law of electrolysis

It states that, the mass or number of moles of a substance deposited at the electrode is inversely proportional to the charge on its ion. i.e. when the same quantity of electricity is passed through solutions of different electrolytes, the relative number of moles of the elements deposited are inversely proportional to the charges on the ions of each of the elements respectively.

For monovalent elements like sodium, silver and potassium. 1 mole of electron is needed to liberate 1 mole of the element.

$Ag^{+}(aq) + e^{-} \longrightarrow Ag(s)$

1 mole of Ag^{+} requires 1 mole of electron to be discharged.

For divalent elements such as copper, iron (II), calcium and zinc. 2 moles of electrons are needed to deposit 1 mole of the substance.

$Cu^{2+}(aq) + 2e^{-} \longrightarrow Cu(s)$

2 moles of electrons are required to deposit 1 mole of copper.

For trivalent elements such as iron (III) and aluminium 3 moles of electrons are needed to deposit 1 mole of the element.

$Al^{3+}(aq) + 3e^{-} \longrightarrow Al(s)$

3moles of electrons are required to deposit 1 mole of aluminium.

Examples

1. A current of 12A is passed through a solution of aluminium chloride for 3 minutes. Calculate the quantity of electricity required to deposit one mole of aluminium.

Solution

$Al^{3+}(aq) + 3e^- \longrightarrow Al(s)$

3 moles of electrons deposit 1 mole of aluminium. But 1 Faraday=1 mole of electrons

Therefore According to equation. 1 mole of Al is deposited by 3F i.e.
(3x96500)C= 289500C.

2. What is the amount of copper deposited when 96500C of electricity is passed through a solution of copper(II)sulphate

Solution

$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

2 moles of electrons deposit 1 mole of copper. Therefore 2 F deposits 1 mole of copper

(2x96500)C deposits 1 mole of copper
1C deposits (1x1/193000) moles of copper
96500C deposits (96500x1x1/193000) moles of copper
=0.5 moles of copper.

General applications of Faraday's laws of electrolysis

Faraday's laws of electrolysis can be used to determine,

- Quantity of electricity passed through an electrolyte for a specific time.
- The mass (moles) or volume of substances liberated/ deposited during electrolysis.
- The relative atomic mass of substances

- The charge on the ions discharged.

Examples of calculations

Determination of quantity of electricity passed and mass /volume of substances liberated during electrolysis

1. A current of 0.5 A flows for 1 hour, 40 minutes and 20 seconds through copper(II) sulphate solution. Determine

i) The quantity of electricity passed

ii) The mass of copper deposited

iii) The volume of oxygen liberated

(1 mole of gas occupies 24000cm^3 r.t.p, Cu=64, O=16, 1F=96500C)

Solutions

i) Q=It

I=0.5A, t= (1x60x60+4x60+20) s =6020s.

Q= (0.5X 6920) C =3010C

ii) $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

2 moles of electrons deposit 1 mole of copper. Therefore 2 Faradays deposit 1mole of copper

(2x96500)C deposits 64g of copper

1C deposits (1x64/193000) g of copper

3010C deposits (3010x1x64/193000) g of copper

=0.998g of copper

i) $4OH^-(aq) - 4e^- \longrightarrow 2H_2O(l) + O_2(g)$

From the equation,

4 moles of electrons liberate 1 mole of oxygen. Therefore, 4 Faradays liberate 1mole of oxygen

(4x96500)C liberates 24000cm^3 of oxygen
1C liberates (1x24000/386000) cm^3 of oxygen
3010C liberates (3010x1x24000/386000)cm^3 of oxygen
=187.2cm^3

2. A current of 0.5 A flows through nickel (II) sulphate solution for 6 minutes and 30 seconds. Determine,

i) The mass of nickel deposited
ii) The volume of gas produced at the anode.
(Ni=59, 1 mole of gas occupies 22.4 dm^3 at s.t.p., 1 f =96500C)

Solution

i) Q=*It*
Given that I=0.5 A
t= (60x60+30)s =390s.

Q=0.5x390 =195C

$$Ni^{2+}(aq) + 2e^- \longrightarrow Ni(s)$$

2 moles of electrons deposit 1 mole of nickel. Therefore 2 Faraday deposits 1mole of nickel

(2x96500)C deposits 59g of nickel
1C deposits (1x59/193000) g of nickel
195C deposits (195x1x59/193000) g of nickel
=0.06g of nickel

ii) $4OH^-(aq) - 4e^- \longrightarrow 2H_2O(l) + O_2(g)$

From the equation,

4 moles of electrons liberate 1 mole of oxygen. Therefore, 4 Faradays liberate 1mole of oxygen

(4x96500)C liberates 22.4dm^3 of oxygen
1C liberates (1x22.4/386000) dm^3 of oxygen
195C liberates (195x1x195/386000)dm^3 of oxygen
=0.0113dm^3 of oxygen was evolved.

Exercise

1. Calculate the volume of oxygen and hydrogen produced when a current of 2A is passed through dilute sulphuric acid for 20 minutes.(Answer Oxygen=0.139 litres, hydrogen=0.278 litres)
2. Calculate the mass of copper deposited on the cathode when current of 3A is passed through copper(II) sulphate solution for 25 minutes.(Answer is 1.49g)
3. 3A was passed through a cell for 20minutes containing dilute sulphuric acid. Find the volume of hydrogen at 25˚C and 120kp.(Answer is 385cm^3 NOT 417.8cm^3)

4. The same amount of current was passed through molten sodium chloride and through cryolite containing aluminiun oxide. If 4.60 g of sodium were liberated in one cell, what is the mass of aluminium liberated in the other cell? (Answer=1.8g)

Determination of relative atomic mass of substances

The RAM of a substance discharged during electrolysis can be estimated if a steady current passed through its solution, time taken and mass deposited are known.

Example

1. A current of 0.25 A flowing for 13 minutes through a nickel (II) Chloride solution deposits 0.059 g of nickel. Determine the RAM of nickel.

Solution

Q=*It*

Given that I=0.25 A

t= (13x60)s =780s.

Q=(0.25x780) =195C

195C deposits 0.059 g of nickel

1C deposits (0.059/195) g of nickel

But from the equation

$$Ni^{2+}(aq) + 2e^- \xrightarrow{\hspace{3cm}} Ni(s)$$

(2 x96500)C deposits (0.059x193000/195)g of nickel.

=58.4 g of nickel

There fore, the RAM of nickel is 58.4.

Determination of the charge on the ion discharged

For the charge on the ion of an element to be determined, the amount (moles) of the element discharged and number of moles of electrons must be known.

$$\text{Amount of element discharged/moles} = \frac{\text{Number of moles of electrons}}{\text{Number of charge on one ion of the element}}$$

Example

1. During electrolysis, it was calculated that 0.2 faradays were needed to deposit 0.1 mo of an element X at the cathode. What is the charge on the ion of X.

Solution

Given that: Faraday is equivalent to moles of electrons=0.2

Moles of X deposited=0.1

$$\text{Amount of element discharged/moles} = \frac{\text{Number of moles of electrons}}{\text{Number of charge on one ion of the element}}$$

From

0.1=0.2/Charge on X

Charge on X=0.2/0.1 =2.

2 When a metal of relative atomic mass 207 is deposited by electrolysis, a current of 0.0600A flowing for 66 minutes increases the mass of the cathode by 0.254g.

a) Find,
  i) The number of moles deposited
  ii) The number of moles of electrons that have passed
b) Deduce the number of units of charge on the cation of the metal

Solution

Given that: RAM=207, I=0.060A, t= (66x60) s =3960s, mass of metal deposited=0.254g

i) Moles=Mass deposited /relative atomic mass =0.253/207 =0.00123 moles.

ii) Quantity of electricity =It =(0.0060x3960)C =237.6C

96500C liberates 1 mole of electron

1 C liberates (1/96500) moles of electrons

237.6 liberates (237.6x1/96500) moles of electrons

=0.00246 moles

Therefore, moles of electrons that have passed is 0.00246 moles.

iii) From

$$\text{Amount of element discharged/moles} = \frac{\text{Number of moles of electrons}}{\text{Number of charge on one ion of the element}}$$

0.00123=0.00246/Charge

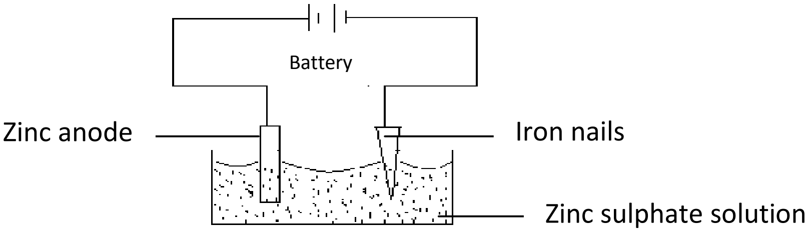
Charge=0.00246/0.00123 =2.

Application of electrolysis

a) Electroplating

This is the coating of one metal by metal using electricity. In electroplating, the metal to be coated is made the cathode and the plating metal is made the anode. The electrolyte must contain ions of the metal used for coating e.g. solution of a salt of the metal used as a coating on the object. Electroplating prevents rusting and improves on the appearance of metals. For an effective electroplating; a clean cathode must be used; a steady current must flow; a steady temperature must be maintained and the concentration of the electrolytes must also be steady.

Setup used to electroplate iron nail with zinc.

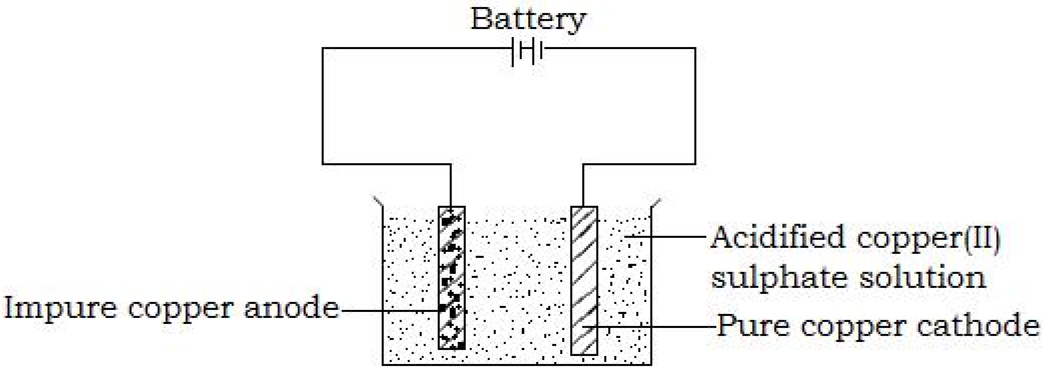


b) Purification of metals.

Metals such as copper and zinc may be purified by electrolysis.

Example: purification of copper

Set up of apparatus



In this case, the impure metal is made the anode and the pure metal the cathode. The electrolyte is usually an acidified solution of a salt containing the metal to be purified. In this case acidified copper (II) sulphate solution is used. During electrolysis, the impure copper anode dissolves and their fore losses mass as copper (II) ions (Cu^{2+})are formed.

$$Cu(s) - 2e^- \longrightarrow Cu^{2+}(aq)$$

The copper (II) ions (Cu^{2+}) move to the pure copper cathodes where they are deposited as a brown layer of pure copper.

$$Cu^{2+} + 2e^- \longrightarrow Cu(s)$$

The impurities (anode sludge) fall to the bottom of the cell.

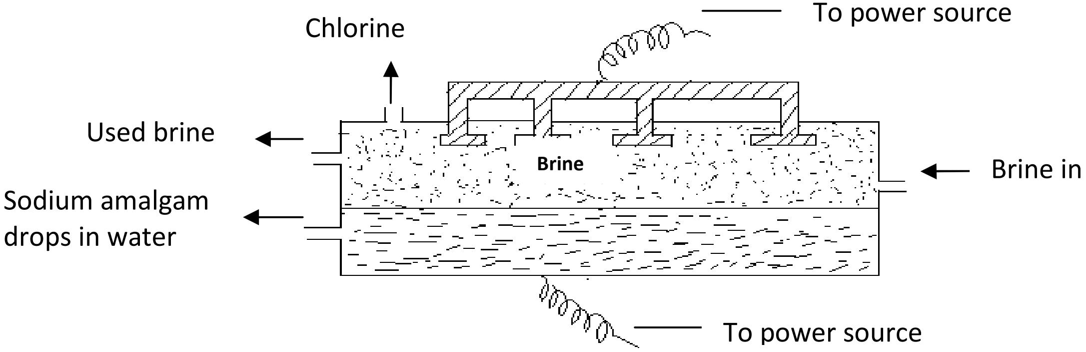
c) **Extraction of reactive metals** like sodium, and aluminium.

d) **Manufacture of chemicals such as sodium hydroxide and chlorine**

Manufacture of sodium hydroxide by electrolysis

Sodium hydroxide is manufactured industrially by the electrolysis of concentrated sodium chloride (Brine) using carbon anode and a layer of mercury as the cathode. The cell is called a **mercury cathode cell**.

The mercury cathode cell



When current is passed through the electrolyte, decomposition takes place as follows.

$$NaCl\,(aq) \longrightarrow Na^+(aq) + Cl^-(aq)$$

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

At the cathode

Both sodium and hydrogen ions move towards the cathode. Because of the nature of the mercury electrode, it will influence the discharge of sodium ion (Na^+)in preference to hydrogen ions (H^+).

$$Na^+(aq) + e^- \longrightarrow Na(s)$$

The sodium formed combines with the mercury to form sodium amalgam. The sodium amalgam is dropped in water to form sodium hydroxide, hydrogen gas and mercury.

Equation

$$Na/Hg(l) + H_2O(l) \longrightarrow NaOH(aq) + H_2\,(l) + Hg(l)$$

The mercury is recycled back, and sodium hydroxide evaporated to dryness and used for various purposes such as manufacture of soap, paper and rayon.

At the anode

Both chloride and hydroxyl ions move to the anode. Due to the high concentration of the chloride ions, it is discharged in preference to hydroxyl ions, therefore forming chlorine gas at the anode.

$$2Cl^-(aq) - 2e^- \longrightarrow Cl_2(g)$$

e) Anodization and dyeing of aluminium

Aluminium objects may be coated with a thin layer of aluminium oxide using electricity, a process known as anodization. During the process, the alluminium object is made the anode and the electrolyte is dilute suphuric acid. The aluminium oxide coating is important in preventing corrosion.

Sample questions on electrolysis

1. *Explain briefly what is meant by the following terms: conductor, non conductor, electrolyte, non electrolyte, electrolysis, electrodes, strong electrolyte and weak electrolyte, giving an example of each. Explain the ionic theory in relation to electrical conduction. How possible is it that Solid lead(II) bromide does not conduct electricity yet molten lead(II) bromide conducts.*
2. *Describe experiments to demonstrate the products formed in the electrolysis of solutions of (a) sulphuric acid (b) sodium sulphate (c) copper(II) sulphate.*
3. *State faradays laws of electrolysis. Describe carefully what happens when copper(II) sulphate solution is electrolysed between (a) platinum and (b) copper electrodes and when sodium chloride solution is electrolysed between (a) platinum and (b) carbon electrodes. Mention two areas where electrolysis is applied.*
4. *What mass of silver and what volume of oxygen at s.t.p. would be liberated in electrolysis by 9650C of electricity? (**Ans. 10.8g ; 560cm^3**)*
5. *How many moles of electrons are required to produce by electrolysis: (a) 27 grams of aluminium (b) 8 grams of oxygen. (**Ans. (a) 3; (b) 1**)*
6. *An element X has a relative atomic mass of 88. When a current of 0.5 amp was passed through the fused chloride of X for 32 minutes 10 seconds, 0.44g of X was deposited at the cathode. (1 faraday = 96500C). (a) calculate the number of faradays needed to liberate 1 mole of X. (b) write the formula for the X ion (c) write the formula for the hydroxide of X. (**Ans.(a)=2 (b)=X^{2+} (c)** **$X(OH)_2$**).*
7. *Explain how you can prepare by electrolysis a sample of pure copper from impure copper.*
8. *Describe carefully what happens when copper(II)sulphate solution is electrolysed between platinum electrodes and write equations for the reaction at both electrodes. Calculate the mass of each product of electrolysis if the current was stopped after the passage of 0.01 Faraday.*
9. *Describe five applications of electrolysis.*

MOLE CONCEPT

Introduction

Mole concept deals with determining or counting the number of particles. Since the number of particles is in large quantities, it becomes very difficult to deal with large numbers; therefore, these particles are placed in groups called moles. One mole of a substance has 6.02×10^{23} particles. The particles of substances grouped into moles can be molecules, atoms, ions, electrons, radicals, protons or any other specified particles. For example 1 mole of magnesium atoms contains 6.02×10^{23} atoms, 1 mole of magnesium ions contains 6.02×10^{23} ions, 1 mole of $H_2SO_4(aq)$ molecules contains 6.02×10^{23} molecules of $H_2SO_4(aq)$.

A **mole** is the amount of substance which contains 6.02×10^{23} particles. Or
A mole is the amount of substance that contains the same number of particles as the number of particles in 12grams of carbon-12.
The number 6.02×10^{23} is called Avogadro's number or constant and it is denoted by letter L.

Molar Mass

Molar mass is the mass of one mole of a substance. It is equal to the relative atomic mass expressed in grams. The relative atomic masses of all elements have already been established. Examples are given below.

Atoms	**Relative atomic mass**	**Molar mass**
Hydrogen	1	1g
Carbon	12	12g
Oxygen	16	16g
Sodium	23	23g
Magnesium	24	24g
Sulphur	32	32g
Aluminium	27	27g
Copper	64	64g
Silver	108	108g
Lead	207	207g
Chlorine	35.5	35.5g
Calcium	40	40g

Potassium	39	39g
Tin	119	119g
Zinc	65	65g
Nickel	59	59g
Nitrogen	14	14g
Iron	56	56g
Phosphorus	31	31g

Formula mass or molecular mass

This is the mass of one mole of a compound. It is obtained by adding the relative atomic masses of the atoms present in a compound. The formula mass of a compound is equal to the relative formula mass expressed in grams.

Examples

Calculate the formula masses of the following compounds.

a) Water ,H_2O (H=1,O=16)
   =(1x2) + 16
   =18g
b) Oxygen molecule, O_2 (0=16)
   =2x16
   =32g
c) Sodium sulphate, Na_2SO_4 (Na=23,S=32,0=16)
   =(2x23)+32+(4x16)
   =46+32+64
   =142g
d) Copper (II) sulphate crystals, $CuSO_4.5H_2O$ (Cu=64,S=32,O=16,H=1)
   =64+32+(4x16)+5(2x1+16)
   =250g

Exercise

Calculate the relative formula masses of the following compounds.

a) $FeSO_4.7H_2O$ (Fe=56,S=32,O=16,H=1)

b) $(NH_4)_2SO_4$ (N=14,H=1,S=32,O=16)

c) Al_2O_3 (Al=27,O=16)

d) $Ca(HCO_3)_2$ (Ca=40,H=1,C=12,O=16)

Calculating number of particles

This is based on the relationship that one mole of a substance contains 6.02×10^{23} particles.

Examples

1. Calculate the number of particles in the following compounds.

a) Water (H_2O)

Number of molecules in 1 mole of H_20 = 6.02×10^{23} molecules

1mole of H_2O contains 2 moles of H atoms, therefore, number of H atoms in 1 mole of H_2O = $(2X6.02\times10^{23})$ =$12.04x10^{23}$ atoms.

Number of O atoms in 1 mole of H_2O =6.02×10^{23}.

b) 2 moles of oxygen molecules,($2O_2$)

1 mole of oxygen molecules contains 6.02×10^{23} molecules

2 moles of oxygen molecules contains $(2x6.02\times10^{23})$= $12.04x10^{23}$ molecules

1 mole of oxygen molecule contains 2 moles of oxygen atoms

2 moles of oxygen molecules contains (2x2)= 4 moles of oxygen atoms

1 mole of oxygen atom contains 6.02×10^{23} atoms

4 moles of oxygen atoms contains $(4x6.02\times10^{23})$= $24.08x10^{23}$ atoms.

c) Number of hydrogen ions in 2 mole of H_2SO_4

1 mole of H_2SO_4 contains 2 moles of H^+

2 moles of H_2SO_4 contains (2x2)= 4 moles of H^+

1 mole of H^+ contains 6.02×10^{23} ions

4 moles of H^+ contains $(4x6.02\times10^{23})$= $24.08x10^{23}$ H^+ ions

d) Total number of ions in 1 mole of $(NH_4)_2SO_4$

$(NH_4)_2SO_4(aq) \longrightarrow 2NH_4^+(aq) + SO_4^{2-}(aq)$

1 mole of $(NH_4)_2SO_4$ contains 3 moles of ions

1 mole of ion contains 6.02×10^{23} ions

3 moles of ions contains $(3x6.02\times10^{23})= 18.06x10^{23}$ ions

2. Calculate the number of atoms in the following
   a) 0.25 moles of calcium

   1 mole of calcium contains 6.02×10^{23} atoms

   0.25 moles of calcium contains $(0.25x6.02x10^{23}/1)$ atoms

   $=1.5x10^{23}$ atoms

   b) 8 moles of sulphur

   1 mole of sulphur contains 6.02×10^{23} atoms

   8 moles of sulphur contains $(8x6.02x10^{23}/1)$ atoms

   $=4.8x10^{23}$ atoms

   c) O.4 g of oxygen atoms

   1 mole of oxygen atom contains 6.02×10^{23} atoms

   16g of oxygen contains 6.02×10^{23} atoms

   1g of oxygen contains $(1x6.02x10^{23}/16)$ atoms

   0.4g of oxygen contains $(0.4x1x6.02x10^{23}/16)$ atoms

   $=1.5 \text{ x } 10^{24}$ atoms

Exercise

1. Calculate the following (Al=27,H=1,O=16,S=32,Ca=40,Na=23,Hg=201,Cl=35.5)
   a) Number of atoms in 2 moles of sodium
   b) Number of molecules in 5 moles of hydrogen (H_2)
   c) Number of ions in 1 mole of $Al_2(SO_4)_3$
   d) Number of hydroxyl ions in 2 moles of $Ca(OH)_2$
2. Calculate the number of particles in the following
   a) 0.1 moles of sodium atoms
   b) 0.5 moles of chlorine atoms

c) O.3 moles of calcium atoms

3. Use the value of 6.02×10^{23} mol^{-1} for the Avogadro constant to find the number of atoms in

a) $2.0x\ 10^{-3}$g of calcium

b) $5.0x10^{-6}$g of argon

c) $1.00x10^{-10}$g of mercury

Converting number of particles to masses

Example

1. Calculate the mass of sodium with $1.5x10^{22}$ sodium atoms.

(Na=23, L=6.02×10^{23} atoms)

6.02×10^{23} atoms is contained in 1 mole of sodium

6.02×10^{23} atoms is contained in 23g of sodium

1 atom is contained in($1x23/6.02x10^{23}$)g of sodium

1.5×10^{22} atoms is contained in ($1.5x10^{22}\ x1x23/6.02x10^{23}$)g of sodium

=0.575g of sodium

2. How many grams of calcium contain

a) $6.02x10^{23}$ atoms

6.02×10^{23} atoms is contained in 1 mole of calcium

6.02×10^{23} atoms is contained in 40g of calcium

1 atom is contained in ($1x40/6.02x10^{23}$)g of calcium

6.02×10^{23} atoms is contained in ($6.02x10^{23}\ x1x40/6.02x10^{23}$)g of calcium

=40g of calcium

b) $1.5x10^{23}$ atoms

6.02×10^{23} atoms is contained in 1 mole of calcium

6.02×10^{23} atoms is contained in 40g of calcium

1 atom is contained in $(1x40/6.02x10^{23})$g of calcium

1.5×10^{23} atoms is contained in $(1.5x10^{23}\ x1x40/6.02x10^{23})$g of calcium

=10g of calcium

c) $3.0x10^{23}$ atoms

6.02×10^{23} atoms is contained in 1 mole of calcium

6.02×10^{23} atoms is contained in 40g of calcium

1 atom is contained in $(1x40/6.02x10^{23})$g of calcium

3.0×10^{23} atoms is contained in $(3.0x10^{23}\ x1x40/6.02x10^{23})$g of calcium

=20g of calcium

Converting masses to moles

Relationships

For atoms, 1 mole is equivalent to Relative atomic mass (RAM) in grams. For example: 1 mole of Cu=64g; 1 mole of S = 32g; 1 mole of H= 1g.

For molecules, 1 mole is equivalent to relative molecular mass (RMM) or relative formula mass (RFM). For example: 1 mole of H_2O= 18g; 1 mole of $CaCO_3$= 100g; 1 mole of H_2SO_4= 98g.

Examples

1. Calculate the number of moles of in;

a) 4g of oxygen molecule (O_2)

(0=16, S=32)

RMM= (2X16) =32

32g is contained in 1 mole of O_2

1 g is contained in ($^1/_{32}$) moles of O_2

4g is contained in ($^{4x1x1}/_{32}$) moles of O_2

=0.125 moles of O_2

b) 160g of sulphur (S)

RAM= 32

32g is contained in 1 mole of S

1 g is contained in ($^1/_{32}$) moles of S

160g is contained in ($^{160x1x1}/_{32}$) moles of S

=5.0 moles of S

2. Calculate the number of moles in the following molecules

a) 2 g of Calcium oxide

RMM of CaO= 40+16 =56

56g is contained in 1 mole of CaO

1 g is contained in ($^1/_{56}$) moles of CaO

2g is contained in ($^{2x1}/_{56}$) moles of CaO

=0.036 moles of CaO

b) 4 g of sodium hydroxide

RMM of NaOH= 23+16+1 =40

40g is contained in 1 mole of NaOH

1 g is contained in ($^1/_{40}$) moles of NaOH

2g is contained in ($^{2x1}/_{40}$) moles of NaOH

=0.1 moles of NaOH

Summary

Number of moles = $\frac{given\ mass}{RAM}$, for atoms

Number of moles= $\frac{given\ mass}{RMM}$, for molecules

Exercise

Given (O=16,Ca=40,C=12,S=32,H=1,Cu=64,Ag=108,Mg=24)

Calculate the number of moles in

a) 21.6 g of silver

b) 12g of magnesium

c) 6g of ammonia

d) 88g of carbondioxide

e) 22.2g of calcium chloride

Converting moles to masses

Examples

Given (O=16,Ca=40,C=12,S=32,H=1,Cu=64)

Calculate the mass in

a) 0.23 moles of sodium

1 mole of sodium weighs 23g

0.23 moles of sodium weighs ($^{0.23x23}/_{1}$)g

=5.29g

b) 7.1 moles of chlorine molecules

RMM of Cl_2

=35.5x2 =71

1 mole of chlorine weighs 71g

7.1 moles of chlorine weighs($^{7.1x71}/_{1}$)g

=504.1g of chlorine

c) 0.1 moles of potassium carbonate

RMM of K_2CO_3

=(39X2)+12+(16X3) =138

1 mole of potassium carbonate weighs 138g

0.1 moles of chlorine weighs($^{0.1x138}/_{1}$)g

=13.8g of chlorine

Exercise

Calculate the mass of

a) 0.1 moles of sodium atom
b) 0.3 moles of chlorine molecules
c) 0.05 moles of calcium atoms
d) 0.05 moles of copper(II)carbonate
e) 0.2 moles of lead(IV)oxide

Calculations on percentage composition by mass

From the formula of a compound, we can calculate the percentage by mass of each elements in a compound.

Percentage composition of an element = $\frac{Mass\ of\ an\ element}{Formula\ mass}$ x 100

Examples

Given (O=16,Ca=40,C=12,S=32,H=1,Cu=64)

Calculate the percentage composition by mass of

a) Oxygen in calcium carbonate

Formula mass of $CaCO_3$ =40+12+(3x16) =100g

Mass of oxygen= (3x16) =48g

Percentage of oxygen= $\frac{48}{100}$ x100

=48%

b) Water in $CuSO_4.5H_2O$

Formula mass of $CuSO_4.5H_2O$ =64+32+(4x16)+5(1x2+16) =250g

Mass of water= 5(1x2+16) =90g

Percentage of oxygen= $\frac{90}{250}$ x100

=36%

c) Nitrogen in nitrogen dioxide

Formula mass of NO_2 = 14+(2x16) =46g

Mass of nitrogen= 14g

Percentage of oxygen= $\frac{14}{46}$ x100

=30.4%

d) Oxygen in baking powder, $NaHCO_3$

Formula mass of $NaHCO_3$ =23+1+12+(3x16) =84g

Mass of oxygen= (3x16) =48g

Percentage of oxygen= $\frac{48}{84}$ x100

=57.14%

Exercise

1. Calculate the percentage of nitrogen in each of the following

a) Ammonium chloride (NH_4Cl)

b) Sodium nitrate ($NaNO_3$)

c) Ammonium sulphate, $(NH_4)_2SO_4$

(N=14,H=1,Cl=35.5,Na=23,O=16,S=32)

2. Calculate the percentage of water of crystallization in sodium carbonate crystals, $Na_2CO_3.10H_2O$
3. A metal sulphate, $X_2(SO_4)_3$ contains 28% by mass of metal X. Determine the Relative Atomic Mass of X and the Relative Molecular Mass of $X_2(SO_4)_3$

(X=56, RMM= 390)

Empirical and Molecular formulae

Empirical formula

This is the simplest formula of a compound which expresses the ratio in which different atoms present in one molecule exists.

Example of empirical and molecular formulae

Compound	**Molecular formula**	**Empirical formula**
Water	H_2O	H_2O
Ammonia	NH_3	NH_3
Ethene	C_2H_4	CH_2
Benzene	C_6H_6	CH
Glucose	$C_6H_{12}O_6$	CH_2O

Steps in calculating empirical formulae

1. Write down the symbols of the elements present
2. Write down the percentage composition or composition by mass below the symbols
3. Find the number of moles of each element by dividing the percentage composition or mass by Relative Atomic Mass
4. Find the mole ratio of the elements by dividing the moles with the smallest number
5. Write down the empirical number.

If the mole ratio is in fractions;

1. Round off to the nearest whole number if it is very close to the whole number.
2. Multiply by a small number that converts the fraction to a whole number if the fraction is not close to a whole number.

Molecular formula

Is a formula that shows the actual number of each atoms present in one molecule of a compound. The molecular formula is a multiple of the empirical formula, so, from the empirical formula, the molecular formula can be determined.

Molecular formula= (Empirical formula)n = Molecular mass
n is number to be determined

Calculations on empirical and molecular formulae

Examples

1. a) Calculate the empirical formula of a compound containing 80% carbon and 20% hydrogen.
   b) If the molar mass of the compound is 30g, determine its molecular formula.

Solution

a) Elements present	C	H
Percentage composition	80	20
Number of moles	$^{80}/_{12}$ 6.7	$^{20}/_{1}$ 20
Divide by the smallest	$^{6.7}/_{6.7}$	$^{20}/_{6.7}$
Mole ratio	1	3

The empirical formula is CH_3

b) (Empirical formula)n= Molar mass

$(CH_3)n=30$
$(12+3)n=30$
$15n=30$

n=2

The molecular formula is therefore $(CH_3)_2 = C_2H_6$

2. Calculate the empirical formula of a compound containing 28% of iron, 24% sulphur and the rest being oxygen.(Fe=56,S=32,0=16)

Solution

Percentage composition of oxygen=100-(28+24)

=100-52

=48%

Elements present	Fe	S	O
Percentage composition	28	24	48
Number of moles	${}^{28}/_{56}$	${}^{24}/_{32}$	${}^{48}/_{16}$
	0.5	0.75	3
Divide by the smallest	${}^{0.5}/_{0.5}$	${}^{0.75}/_{0.5}$	${}^{3}/_{0.5}$
Mole ratio	1	1.5	6
	2x (1	1.5	6)
	2	3	12

The empirical formula is $Fe_2S_30_{12}$ or $Fe_2(SO_4)_3$

3. A hydrocarbon contains 85.7% carbon and its relative molecular mass is 28. Work out its molecular formula.

Solution

Percentage of hydrogen 100%-85.7% =14.3%

Elements present	C	H
Percentage composition	85.7	14.3
Number of moles	$^{85.7}/_{12}$	$^{14.3}/_{1}$
	7.14	14.3
Divide by the smallest	$^{7.14}/_{7.14}$	$^{14.3}/_{7.14}$
Mole ratio	1	2

The empirical formula is CH_2

(Empirical formula)n= Molar mass

$(CH_2)n=28$

$(12+2)n=28$

$14n=28$

$n=2$

The molecular formula is therefore $(CH_2)_2 = C_2H_4$

4. Calculate the empirical formula of a compound that contains 52g of zinc, 9.6g of carbon and 38.4g of oxygen.(Zn=65,C=12,0=16)

Elements present	Zn	C	O
Composition by mass	52	9.6	38.4
Number of moles	$^{52}/_{65}$	$^{9.6}/_{12}$	$^{38.4}/_{16}$
	0.8	0.8	2.4

Divide by the smallest	$^{0.8}/_{0.8}$	$^{0.8}/_{0.8}$	$^{2.4}/_{0.8}$
Mole ratio	1	1	3

The empirical formula is $ZnCO_3$

Exercise

1. A compound X consists of carbon 40%, hydrogen 6.7% and the rest being oxygen. If the RMM is 60, determine its molecular formula.(C=12,H=1,0=16)(**Ans.** CH_2O)
2. A hydrocarbon is made up of 92.3% carbon and has molecular formula of 78g. Calculate its empirical and molecular formula.(**Answer CH**)
3. Calculate the empirical formula of the compound formed when 1.8g of carbon forms 2.4g of a hydrocarbon. (**Answer** CH_4)
4. Calculate the molecular formula of a hydrocarbon with empirical formula CH_2 and molecular mass of 28g. (**Answer** C_2H_4)
5. Calculate the empirical formula of a salt with the following composition, copper 25%, sulphur 12.8%, oxygen 25.6% and water 36.0% (**Answer** $CuSO_4.5H_2O$)
6. Calculate the empirical formula of a hydrated salt with the following composition, sodium 16.09%, carbon 9.20%, oxygen 16.78% and water 62.93% (**Answer** $Na_2C_2O_3.10H_2O$)
7. Find the empirical formulae of the compounds formed in the reactions described below.

a) 10.800g magnesium form 18.000g of an oxide (**Answer=MgO**)
b) 3.400g calcium form 9.435g of a chloride (**Answer=**$CaCl_2$)
c) 3.528g iron form 10.237g of a chloride.(**Answer=**$FeCl_3$)

8. Calculate the empirical formulae of the compounds from which the following analytical results were obtained.

a) 27.3%C, 72.7%O (**Answer=**CO_2)
b) 53.0%C, 47.0%O (**Answer=**C_3O_2)
c) 29.1%Na, 40.5%S,30.4%O (**Answer=**$Na_2S_2O_3$)
d) 32.4%Na, 22.5%S, 45.0%O (**Answer=**Na_2SO_4)

Calculation of masses from equations

Moles and mole ratios can be used to calculate the amount of substances reacting and products formed. This requires that a correctly balanced equation is written. Such an equation is known as a **stoichiometric equation**. **Stoichiometry** is the relationship between amounts of reactants and products in a chemical reaction.

A stoichiometric equation is an equation in which the reactants and products are correctly balanced.

Steps involved in the calculation

1. Write down a balanced equation for the reaction
2. Write down the moles of substances that concerns the question
3. Convert the moles into grammes

Examples

1. Calculate the mass of iron (II) sulphide formed by heating 64g of sulphur with excess iron filling.(S=32,Fe=56)

   Solution

   Molecular mass of FeS= 56+32 =88g

   Equation for reaction

   $Fe(s) + S(s) \longrightarrow FeS(s)$

   1 mole of sulphur forms 1 mole of Iron (II)sulphide

   32g of S forms 88g of FeS

   1 g of S forms ($^{1x88}/_{32}$)g of FeS

   64g of S forms ($^{64x1x88}/_{32}$)g of FeS

   =176g of FeS

2. What is the mass of magnesium required to form 55g of magnesium oxide.

   Solution

   Molecular mass of MgO= 24+16 =40g

   Equation for reaction

   (2x40)g of MgO is formed by (2x24)g of Mg

   $2Mg(s) + O_2(s) \longrightarrow 2MgO(s)$

2 moles of magnesium oxide is formed 2 mole of magnesium

(2x40)g of MgO is formed by (2x24)g of Mg

80g of MgO is formed by 48g of Mg

1 g of MgO is formed by ($^{1x48}/_{80}$)g of Mg

55g of MgO is formed by ($^{55x1x48}/_{80}$)g of Mg

=33g of Mg

3. a) Calculate the mass of copper formed when 3.2g of copper(II)oxide is completely reduced to the metal by hydrogen gas.

   b)How many grams of water was produced

   c)Calculate the mass of hydrogen used in the experiment

Solution

a) $CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$

1 mole of CuO forms 1 mole of Cu

(64+16)g of CuO forms 64g of Cu

80g of CuO forms 64g of Cu

1 g of CuO forms ($^{1x64}/_{80}$)g of Cu

3.2g of CuO forms ($^{3.2x1x64}/_{80}$)g of Cu

=2.6g of Cu

b) 1 mole of CuO forms 1 mole of H_2O

(64+16)g of CuO forms (1x2+16)g of H_2O

80g of CuO forms 18g of H_2O

1 g of CuO forms ($^{1x18}/_{80}$)g of H_2O

3.2g of CuO forms ($^{3.2x1x18}/_{80}$)g of H_2O

=0.7g of H_2O

c) 1 mole of CuO reacts with 1 mole of H_2

(64+16)g of CuO reacts with (1x2)g of H_2

80g of CuO reacts with 2g of H_2

1 g of CuO reacts with ($^{1x2}/_{80}$)g of H_2

3.2g of CuO reacts with ($^{3.2x1x2}/_{80}$)g of H_2

=0.08g of H_2

Exercise

1. A solution of 8.1g of NaOH was neutralized by hydrochloric acid. Calculate the mass of sodium chloride produced when the solution was evaporated to dryness. (C=12, Na=23, O=16, H=1, Cl=35.5,) (**Answer =11.85g**)
2. Calculate the mass of residue left when 2.40g of sodium hydrogen carbonate is decomposed by heat. (**Answer =1.51g**)
3. Calculate the loss in mass when 100g of calcium carbonate is heated to constant mass. (Ca=40, C=12, 0=16) (**Answer =44g**)
4. 76.5g of calcium hydrogen carbonate was heated strongly. What was the mass of carbon dioxide formed? (**Answer =20.78g**)
5. What mass of sodium oxide would be made from 1.5 g of sodium? (**Answer=2.02g**)

MOLECULAR THEORY OF GASES

The behavior of gases when subjected to temperature and pressure changes can be expressed in two simple laws i.e Boyle's law and Charles law. These laws apply to ideal gases (i.e. gases with negligible intermolecular forces of attraction and negligible volume)

Boyle's law

It states that *the volume of a given mass of a gas is inversely proportional to its pressure at constant temperature*.

Mathematically,

$V \alpha \frac{1}{P}$ Where P=pressure, V= volume,

$V = \text{Constant} \times \frac{1}{P}$

$V = \frac{\text{Constant}}{P}$

$PV = \text{Constant}$,

An ideal gas at a pressure P_1 occupying volumes V_1, can be made occupy volume V_2 at a pressure P_2

Therefore $P_1V_I=P_2V_2$

Charles' law

It states that, *the volume of a given mass of gas is directly proportional to its absolute temperature at a constant pressure.*

Mathematically,

$V \alpha T$

$V = \text{Constant} \times T$

$\frac{V}{T} = \text{Constant}$

Relating Volumes and absolute temperatures of an ideal gas at different values, we conclude that

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

The temperature at which the volume of gases would be theoretically zero is the **absolute temperature** or **absolute zero.** The temperatures on the absolute temperature scale/Kelvin scale are measured in kelvins and obtained by adding 273.15 to the temperature on the Celsius scale. i.e. 273.15K=0˚C

For example if the temperature is 100˚C, on the Kelvin scale it will be (100+273.15) =373.15K.

The ideal gas equation

This equation is obtained by combining Boyle's law ($PV = \text{Constant}$) and Charles' law ($\frac{V}{T} = \text{Constant}$), we obtain

$\frac{PV}{T}$ = Constant

This is often written as

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

This is the **equation of state for an ideal gas**. It enable one to calculate the effect of a change in temperature and pressure on the volume of a gas. One cannot compare gas volumes unless they are stated at the same temperature and pressure. Gas volumes are usually compared at 0˚C (273.15K) and 1 atmosphere(760mmHg). These conditions are referred to as standard temperature and pressure (stp). Sometimes volumes are quoted at room temperature and pressure(rtp): 20˚C and 1 atmosphere(760mmHg).

Examples

1. A certain gas occupies 211cm^3 at 18˚C and 740 mmHg pressure. What volume will it occupy (still gaseous) at -20˚C and 770mmHg pressure?

Solution

Given

P_1=740mmHg, T_1=(18+273)K =291K and V_1=211cm^3

P_2=770mmHg, T_2=(-20+273)K =2253K and V_2=?

Applying

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}, \quad \frac{740x211}{291} = \frac{770xV_2}{253} \quad V_2=\frac{740x211x253}{770x291}\text{ cm}^3 = 176\text{cm}^3$$

2. A certain mass of a gas occupies 146cm^3 at 18˚C and 738 mmHg pressure. Calculate its volume at s.t.p.

Solution

Given

P_1=738mmHg, T_1=(18+273)K =291K and V_1=146cm^3

P_2=760mmHg, T_2=273K and V_2=?

Using

$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$, $\frac{738x146}{291} = \frac{760xV_2}{273}$ $V_2=\frac{738x146x273}{760x291}$ $cm^3 = 133cm^3$

Gay-Lussac's law

It's the third law describing behavior of gases when involved in chemical reactions

The law states that; *when gases react, they do so in volumes which bear a simple ratio to one another and to the volume of the product if gaseous provided the temperature and pressure remains constant.*

Gay-Lussac studied chemical reactions between gases and noticed that there is always a very simple ratio between the volumes of gases that react together. For instance

1. $2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$
   2 volumes of hydrogen combine with 1 volume of oxygen to form 2 volumes of steam
2. $H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$
   1 volume of hydrogen combines with 1 volume of chlorine to form 2 volumes of hydrogen chloride gas
3. $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$
   1 volume of nitrogen combines with 3 volumes of hydrogen to form 2 volumes of ammonia gas

Examples

1. $100cm^3$ of hydrogen where burnt in excess chlorine. Determine
   a) The volume of chlorine used
   b) The volume of gaseous product formed

Solution

Equation

a) $H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$

1 volume of hydrogen combines with 1 volume of chlorine

$100cm^3$ of H_2 combines with $(^{1x100}/_1)$ cm^3 of Cl_2

$=100\ cm^3$ of chlorine

b) Form the equation, 1 volume of H_2 produces 2 volumes of HCl

1 volume of H_2 produces with 2 volume of HCl

$100cm^3$ of H_2 produces $(^{2x100}/_{1})$ cm^3 of HCl

$=200\ cm^3$ of HCl

2. $80cm^3$ of oxygen are exploded with 80 cm^3 of hydrogen, determine the volume of the unused gas (residual gas) at room temperature

Solution

Equation

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

2 volumes of hydrogen combines with 1 volume of oxygen

1 volumes of hydrogen combines with $(^{1}/_{2})$ volume of oxygen

$80cm^3$ of hydrogen combines with $(^{80}/_{2})$ cm^3 of oxygen

$=40\ cm^3$ of oxygen

Unused gas (oxygen)= original volume of oxygen-used volume

$=(80-40)\ cm^3 = 40\ cm^3$ of oxygen

3. $48cm^3$ of methane was exploded with 212 cm^3 of oxygen at a certain temperature and pressure. The mixture was then allowed to attain room temperature and pressure. Determine the volume of each of the remaining gases.

Solution

Equation

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g)\quad 2H_2O(g)$$

1 volume of methane combines with 2 volumes of oxygen

$48cm^3$ of methane combines with $(^{2x48}/_{1})$ cm^3 of oxygen

$=96\ cm^3$ of oxygen

All the methane was used up and (212-96) =116 cm^3 of oxygen remained unused.

Also, from the equation

1 volume of methane produces 1 volume of carbondioxide

$48cm^3$ of methane produces $(^{1x48}/_{1})$ cm^3 of carbondioxide

=48 cm^3 of carbondioxide

N.B. At room temperature, the 2 volumes of steam produced condensed to water.

The remaining gases are unused oxygen + carbondioxide formed

=(116+48) cm^3

=164 cm^3

Exercise

1. What is the volume of oxygen that will be required for complete combustion of 100 cm^3 of carbonmonoxide? What is the volume of the product formed.(All volumes at the same temperature and pressure.)
2. 200 cm^3 of carbonmonoxide were burnt in 80 cm^3 of oxygen and the mixture was allowed to attain the original room temperature and pressure. What would be the volume of each of the remaining gases.
3. 50 cm^3 of oxygen were added to a 40 cm^3 mixture of hydrogen and nitrogen. After explosion and cooling to the original temperature,the residual gas occupied 45 cm^3. What was the percentage of oxygen in the original mixture.(Answer is 75%)
4. Calculate the volume of oxygen required for the complete combustion of 200 cm^3 of ethane (C_2H_6). What was the volume of gas product formed.

Avogadro's law

The law states that *equal volume of gases at the same temperature and pressure contains the same number of molecules.*

Avogadro's law gives an interpretation to Gay-Lussac's law in terms of molecules of gases. Consider the following examples

1. $2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$

   2 volumes of hydrogen combine with 1 volume of oxygen to form 2 volumes of steam

   Is interpreted as

   2 molecules of hydrogen combine with 1 molecule of oxygen to form 2 molecules of steam

2. $H_2(g) + Cl_2(g) \xrightarrow{\hspace{2cm}} 2HCl(g)$

   1 volume of hydrogen combines with 1 volume of chlorine to form 2 volumes of hydrogen chloride gas

   Is interpreted as

   1 molecule of hydrogen combines with 1 molecule of chlorine to form 2 molecules of hydrogen chloride gas

This law is quite important because it enables us to change from a statement about volumes of gases to the same statement about moles of gases and vise versa.

The relationship between vapor density and Relative Molecular Mass can be deduced from Avogadro's law and is expressed as

2 x Vapor Density= Relative Molecular Mass

I.e. Relative Molecular Mass is twice the value of vapor density.

It follows from Avogadro's law that if equal volumes of gases contain equal numbers of molecules then the volume occupied by one mole must be the same for all gases. It is called the **gas molar volume**.

Gas molar volume

The molar gas volume is the volume occupied by one mole of a gas. It is the same for all gases under the same conditions of temperature and pressure. Values for molar gas volumes are given in the table below.

Condition	**Value for molar gas volume**
Standard temperature and pressure(stp)	22.4l or 22.4dm^3 or 22400cm^3
Room temperature and pressure (rtp)	24l or 24dm^3 or 24000cm^3

Example

1. Determine the number of moles in the following gaseous volumes at stp
   a) 1.2 dm^3 of nitrogen
   b) 300cm^3 of ammonia

Solution

a) 22.4 dm^3 is occupied by 1 mole of nitrogen

1 dm^3 is occupied by ($^1/_{22.4}$) moles of nitrogen

1.2 dm^3 is occupied by ($^{1.2x1}/_{22.4}$) moles of nitrogen

= 0.05 moles of nitrogen at stp

b) 22400 cm^3 is occupied by 1 mole of ammonia

1 cm^3 is occupied by ($^1/_{22400}$) moles of ammonia

300 cm^3 is occupied by ($^{300x1}/_{22400}$) moles of ammonia

= 0.013 moles of ammonia at stp

2. Determine at rtp the volume and mass of
   a) 0.04 moles of hydrogen
   b) 0.2 moles of carbondioxide

Solution

a) 1 mole of hydrogen occupies 24l at rtp

0.04 moles of hydrogen occupies ($^{0.04x24}/_1$) l at rtp

=0.96 l of hydrogen at rtp

RMM for H_2 =(1x2)= 2

1 mole of hydrogen weighs 2g

0.04 moles of hydrogen weighs ($^{0.04x2}/_1$)g

=0.08g of hyrogen

b) 1 mole of carbondioxide occupies 24l at rtp

0.2 moles of carbondioxide occupies ($^{0.2x24}/_1$) l at rtp

=4.8 l of carbondioxide at rtp

RMM for CO_2 =12+(16x2)= 44

1 mole of carbondioxide weighs 44g

0.2 moles of carbondioxide weighs ($^{0.2x44}/_{1}$)g

=8.8g of carbondioxide

3. Calculate the RMM of Y given that 0.8 g of Y occupies 560cm^3 at stp.

Solution

560 cm^3 of Y weighs 0.8g

1 cm^3 of Y weighs ($^{0.8}/_{560}$) moles of nitrogen

22400 cm^3 of Y weighs ($^{22400x1x0.8}/_{560}$) moles of nitrogen

=32g

The RMM of Y is 32.

Calculation of masses and volumes

Examples

1. 0.2 moles of ammonia gas measured at stp were passed over copper(II)oxide. The equation for the reaction is

$2NH_3(g) + 3CuO(s) \xrightarrow{\qquad} 3Cu(s) + N_2(g) + 3H_20(l)$

Calculate

a) The number of moles of copper(II)oxide used

b) The mass of copper formed

c) The volume of ammonia used at stp

(Cu=64,O=16,N=14,H=1, 1 mole of a gas occupies 22.4 l at stp)

Solution

1.

a) From the equation,

2 moles of NH_3 reacts with 3 moles of CuO

1 mole of NH_3 reacts with ($^{3}/_{2}$) moles of CuO

0.2 moles of NH_3 reacts with ($^{3x0.2}/_{2}$) moles of CuO

=0.3 moles of CuO

b) From the equation

2 moles of NH_3 produce 3 moles of Cu

2 moles of NH_3 produce (3x64)g of Cu

1 mole of NH_3 produces $(^{192}/_{2})$ moles of Cu

0.2 moles of NH_3 produces $(^{192x0.2}/_{2})$ moles of Cu

=19.2g of Cu

c) 1 mole of NH_3 occupies 22.4l at stp

0.2 moles of NH_3 occupies $(^{22.4x0.2}/_{1})$l at stp

=4.48 l of NH_3 at stp

2. 2.5 g of $CuCO_3$ were heated to constant mass. Determine
   a) The mass of the residue
   b) The volume of gas produced at stp

   (Cu=64,O=16,C=12, 1 mole of a gas occupies 24 dm^3 at stp)

Solution

$CuCO_3(s) \xrightarrow{\hspace{2cm}} CuO(s) + CO_2(g)$

a) From the equation

1 mole of $CuCO_3$ produces 1 mole of CuO

(64+12+16x3)g of $CuCO_3$ produces (64+16)g of CuO

124g of $CuCO_3$ produces 80g of CuO

1 g of $CuCO_3$ produces $(^{80}/_{124})$ g of CuO

2.5g of $CuCO_3$ produces $(^{2.5x80}/_{124})$ g of CuO

=1.61g of CuO (residue)

b) From the equation

1 mole of $CuCO_3$ produces 1 mole of CO_2

(64+12+16x3)g of $CuCO_3$ produces 24dm^3 of CO_2

124g of $CuCO_3$ produces 24dm^3 of CO_2

1 g of $CuCO_3$ produces $(^{24}/_{124})$ dm^3 of CO_2

2.5 g of $CuCO_3$ produces $(^{2.5x24}/_{124})$ dm^3 of CO_2

=0.484 dm^3 of CO_2

3. From the equation

$2Fe(s) + 3Cl_2(g) \xrightarrow{\hspace{2cm}} 2FeCl_3(s)$

Calculate the

a) Volume of chlorine at stp required to react with 8 g of iron

b) Mass of iron(III)chloride formed

Solution

a) From the equation

2 moles of Fe react with 3 moles of Cl_2

(2x56)g of Fe react with (3x22.4)dm^3 of Cl_2

112g of Fe react with 67.2dm^3 of Cl_2

1 g of Fe react with($^{67.2}/_{112}$)dm^3 of Cl_2

8 g of Fe react with($^{67.2x8}/_{112}$)dm^3 of Cl_2

=4.8 dm^3 of Cl_2 at stp

b) From the equation

2 mole of Fe produce 2 mole of $FeCl_2$

(2x56)g of Fe produces 2x(56+35.5x2)g of $FeCl_2$

112g of Fe produces 325g of $FeCl_2$

1 g of Fe produces ($^{325}/_{112}$) g of $FeCl_2$

8 g of Fe produces ($^{8x325}/_{112}$) g of $FeCl_2$

=23.01g of $FeCl_2$

Exercise

1. What volume of hydrogen at stp would be used if 40g of nitrogen combined with it to form ammonia.
2. 0.1 mole of lead nitrate is completely decomposed on heating according to the reaction

   $$2Pb(NO_3)_2(s) \longrightarrow 2PbO(s) + 4NO_2(g) + O_2(g)$$

   a) What volume at stp of nitrogen dioxide and oxygen were evolved

   b) What was the mass of the residue left
3. 1.4 litres of ammonia gas measured at stp where passed over hot copper(II)oxide. The equation for the reaction is

   $$3CuO(s) + 2NH_3(g) \longrightarrow 3Cu(s) + N_2(s) + 3H_2O(l)$$

   Calculate

   a) The number of moles of ammonia used

   b) The number of moles of nitrogen gas used

   c) The mass of copper formed

4. 3.1g of a carbonate XCO_3 are heated to constant mass; 2.0g of the metal oxide are formed.. when heated in a stream of hydrogen for sufficient time, the oxide is reduced to 1.6g of pure metal.
   a) Give the equation for the two reactions
   b) Calculate the minimum volume of hydrogen at stp needed to reduce the oxide
   c) Calculate the volume of carbondioxide at stp produced when the carbonate was completely decomposed.
5. Calculate the mass of sulphur deposited when 8.4 dm3 of chlorine oxidizes hydrogen sulphide.

Volumetric analysis

In an investigation to determine the nature of matter in a substance, a chemist focuses on two main questions:

a) What are the components of the substance?
b) What amount of each component is present in the substance?

In an attempt to answer question (a) and (b), a chemist carries out **qualitative** and **quantitative analysis** respectively. Most of the reactions which a chemist carries out take place in solution.

Consider the reaction below

$$A(aq) + B(aq) \longrightarrow C(aq) + D(aq)$$

If we are to determine the volume of A needed to completely react with a given amount or volume of B, the answer is provided practically through volumetric analysis.

In volumetric analysis, quantities of substances (often acids or alkalis) are estimated by analytical processes involving measurements of volumes of solutions using pipettes, burettes and measuring cylinders (for approximate measurement). Weighing may also be involved. Most of the work in volumetric analysis is based upon molar (M) solutions.

Standard and Molar solutions

A standard solution is a solution of known concentration. Examples of standard solutions are; solution containing 12g of sodium chloride in one litre of a solution; a solution containing 2 moles of solute in 1dm^3 e.t.c. The substance that is used to prepare a standard solution is known as a **primary standard**.

A Molar solution is a solution that contains one mole of a substance in a solution of one litre. In other words, it is a solution containing one mole of solute in one litre.

Other related terms are;

Concentration; this is the amount of solutes in a given volume of solution.

Molarity; this is the number of moles of solute in one litre of a solution. The unit is mol/dm^3 or mol/l. The molarity of a solution is commonly denoted by letter M. E.g. 0.2M NaOH which mean 1 litre of a solution containing 0.2 moles of NaOH.

1 litre(1 l)=1cubic decimetre (1dm^3) = 1000 cubic centimetre (1000cm^3)

Calculations on molarity and masses

Examples

1. Calculate the molarities of the following solution given (Na=23,O=16 H=1,C=12,Cl=35.5,S=32)
   a) 13.5g of copper(II)chloride in 1dm^3 of solution
   b) 4.0g of sodium hydroxide in 400cm^3 of solution
   c) 53g of anhydrous sodium carbonate in 2 dm^3 of solution

Solution

a) RMM of $CuCl_2$

=64+(35.5x2)

=135

135g is contained in 1 mole of $CuCl_2$

1 g is contained in ($^1/_{135}$) moles of $CuCl_2$

13.5g is contained in ($^{13.5x1}/_{135}$) moles of $CuCl_2$

=0.1M of $CuCl_2$

b) 400cm^3 of solution contains 4.0g of NaOH

1cm^3 of solution contains ($^{4.0}/_{400}$)g of NaOH

1000cm^3 of solution contains ($^{1000x4.0}/_{400}$)g of NaOH

=10g/l of NaOH (concentration in grams/litre)

RMM of NaOH

=23+16+1

=40

40g is contained in 1 mole of NaOH

1 g is contained in ($^{1}/_{40}$) moles of NaOH

10 g is contained in ($^{10x1}/_{40}$) moles of NaOH

=0.25M NaOH

c) 2dm^3 of solution contains 53g of Na_2CO_3

1dm^3 of solution contains ($^{53}/_{2}$)g of Na_2CO_3

=26.5g/dm^3 of Na_2CO_3 (concentration in grams/litre)

RMM of Na_2CO_3

=23x2+12+16x3

=106

106g is contained in 1 mole of Na_2CO_3

1 g is contained in ($^{1}/_{106}$) moles of Na_2CO_3

10 g is contained in ($^{10x1}/_{106}$) moles of Na_2CO_3

=0.25M Na_2CO_3

In general, $Molarity = \frac{concentration\ (g/l)}{molar\ mass}$

When the concentration of a solutr in grams per litre and the RMM are known then the molarity can be calculated from the above expression.

N.B. The use of formula is not so much recommended and workings should be from first principle.

2. Calculate the mass of the named substance needed to make
   a) 0.1 dm^3 of 2M sodium sulphate solution
   b) 1 l of 0.25M sodium hydroxide solution
   c) $25cm^3$ of 0.1M potassium carbonate solution
   d) $500cm^3$ of 0.05M sodium carbonate solution

Solution

a) $1dm^3$ of solution contains 2 moles of Na_2SO_4

$0.1dm^3$ of solution contains ($^{0.1x2}/_1$) moles of Na_2SO_4

=0.2 moles of Na_2SO_4

RMM for Na_2SO_4

=23x2+32+16x4

=142

1 mole of Na_2SO_4 weighs 142g

0.2 moles of Na_2SO_4 weighs ($^{0.2x142}/_1$))g

=28.4g

b) $1l$ of solution contains 0.25 moles of NaOH

RMM for NaOH

=23+16+1

=40

1 mole of NaOH weighs 40g

0.25 moles of NaOH weighs ($^{0.25x40}/_1$))g

=10g

c) 1000cm^3 of solution contains 0.1moles of K_2CO_3

1 cm^3 of solution contains ($^{0.1}/_{1000}$) moles of K_2CO_3

25 cm^3 of solution contains ($^{25x0.1}/_{1000}$) moles of K_2CO_3

=0.0025 moles of K_2CO_3

RMM of K_2CO_3

=39x2+12+16x3

=138

1 mole of K_2CO_3 weighs 138g

0.0025 moles of K_2CO_3 weighs ($^{0.0025x138}/_1$))g

=0.345g

Calculating number of moles of ions in standard solutions

Examples

1. Calculate the number of moles of hydrogen ions in 25cm^3 of a 0.2 M sulphuric acid.
2. Calculate the number of moles of potassium ions in 35cm3 of 0.12 M potassium carbonate solution.

Solution

1. 1000cm^3 of solution contains 0.2moles of H_2SO_4

1 cm^3 of solution contains ($^{0.2}/_{1000}$) moles of H_2SO_4

25 cm^3 of solution contains ($^{25x0.2}/_{1000}$) moles of H_2SO_4

=0.005 moles of H_2SO_4

From the equation of ionization of H_2SO_4

$$H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{-2}(aq)$$

1 mole of H_2SO_4 produces 2 moles H^+

0.005 moles of H_2SO_4 produces ($^{0.005x2}/_1$) moles H^+

=0.01 moles H^+

2. 1000cm^3 of solution contains 0.12 moles of K_2CO_3

   1 cm^3 of solution contains ($^{0.12}/_{1000}$) moles of K_2CO_3

   35 cm^3 of solution contains ($^{35x0.12}/_{1000}$) moles of K_2CO_3

   =0.0042 moles of K_2CO_3

From the equation of ionization of K_2CO_3

$K_2CO_3(aq) \longrightarrow 2K^+(aq) + CO_3^{-2}(aq)$

1 mole of K_2CO_3 produces 2 moles K^+

0.0042 moles of K_2CO_3 produces ($^{0.0042x2}/_{1}$) moles K^+

=0.0084 moles K^+

Titration

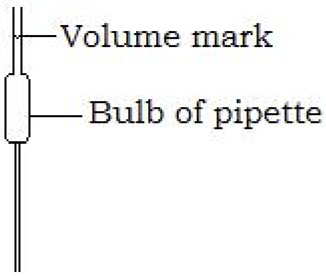
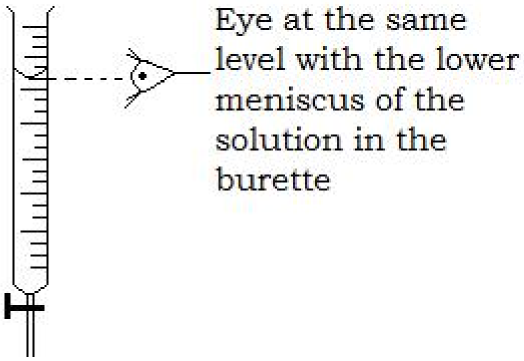
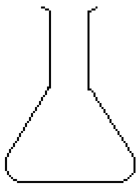
This is a method of volumetric analysis in which a solution (usually a standard solution) is added from a burette to another solution (usually whose concentration is unknown) until the reaction is complete. An indicator changes color immediately the reaction is complete or when the **end point** is reached. Most titrations at this level are acid-base titrations.

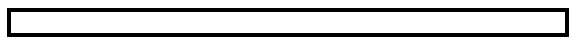
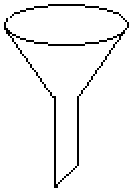
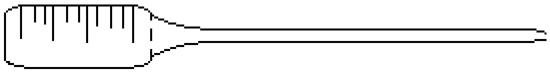
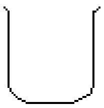
Common laboratory indicators and color changes

Indicator	**Color in acidic solution**	**Color in alkaline solution**
Phenolphthalein	Colorless	Purple
Methyl orange	Red/pink	Yellow

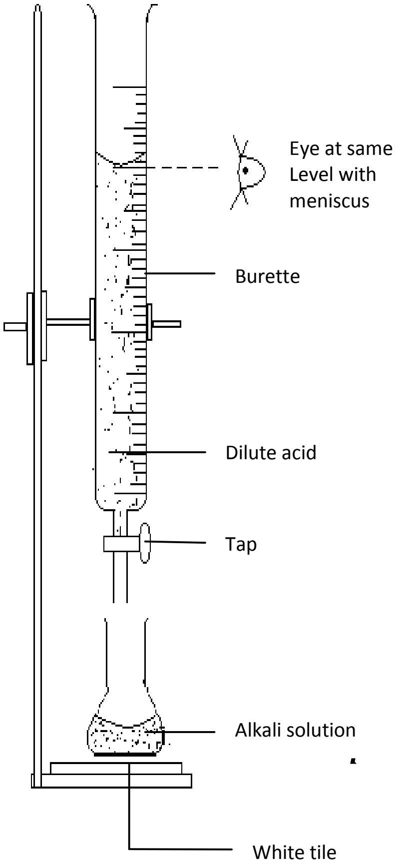
Litmus	Red	Blue

Apparatus commonly used in titration

Apparatus	Use
Pipette 	This is used for measuring and transferring fixed volumes of solutions. The volume of solution (usually whose concentration is not known) is measured accurately and transferred into a conical flask using a pipette. The solution transferred by the pipette is the **aliquot**. The common pipettes used are of volumes 20.0cm^3 and 25.0cm^3 which are usually written on them. The reading of a pipette is recorded to one decimal place e.g. 20.0 cm$^{3.}$
Burette 	This is used to transfer accurately a precise volume of solution. The **titrant** is filled into a burette using a filter funnel to prevent wastage of solutions. The reading of a burette is recorded to two decimal places e.g. 0.00cm^3, 18.40cm^3 e.t.c. For accurate reading of volumes of liquid in burettes and pipettes, the eye must be at the same level with the meniscus of the liquid as illustrated besides.
Conical flask 	It is a container in which the reaction between the acid and the alkali take place. The titrant is run from a burette into a conical flask containing a known volume of the solution (usually whose concentration is not known).
White tile	It enables one to observe indicator color

	changes more clearly. A piece of white paper can be used in place of a white tile.
Filter funnel 	Enables the solution to be poured directly into the burette without wastage
Dropper 	For addition of the indicator to the solution in the conical flask
Beaker 	For holding and pouring out solutions

Procedure for acid-base titration



1. Wash the pipette with distilled water then a little of the solution it is to measure. Use the pipette to deliver either 20.0cm^3 or 25.0cm^3 of the alkali into a clean conical flask. Add a few drops (2 or 3 drops) of indicator.
2. Wash the burette with distilled water then a little of the acid solution and run out the acid through the tap. Fill the burette above the 0 cm^3 mark and run a little of the acid out to bring the meniscus of the acid to the 0 cm^3 mark or slightly below it. Take the burette reading as V_1 cm^3.
3. Arrange the apparatus as shown on the left hand side. Run the acid solution from the burette drop wise. Use your left hand to open the tap and your right hand to

swirl the conical flask (unless you are left handed). Stop when the indicator just changes color. This is the end point the titration.

4. Take the burette reading again V_2 cm^3. Subtract (V_2- V_1) cm^3 to get the "**titre**" (i.e. the volume of the acid needed to neutralize the known volume of alkali).
5. Repeat the titration. Obtain an average titre. From this volume you can calculate the unknown concentration.

N.B The first titration is regarded as a trial run (rough titration) and may not be very accurate, therefore the value may not be used in computing the average volume. Values used in calculating the average volume must be close to each other

Specimen readings

Example

Neutralization of 0.1M NaOH solution with a solution of HCl

Volume of pipette used =25.0cm^3

Number of titration	1	2	3
Final burette reading/ cm^3	14.80	30.00	15.00
Initial burette reading/ cm^3	0.00	15.00	0.00
Volume of acid used/ cm^3	14.80	15.00	15.00

Value used to calculate average volume of acid used: 15.00 cm^3 and 15.00 cm^3

Average volume of acid used=$\frac{15.00+15.00}{2}$ =15.00 cm^3

Calculate

a) The number of moles of sodium hydroxide that reacted
b) The number of moles of hydrochloric acid that reacted
c) The molarity of the hydrochloric acid (i.e concentration in mol/litre)

Method of calculation

1. Write the equation for the reaction that took place. This gives you the mole ratio of reaction between the acid and the alkali.

   $HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$

   1 mole of HCl neutralizes 1 mole of NaOH

2. Work out the number of moles of the standard solution. In this case it is the alkali (NaOH) as its concentration is known (25.O cm3 contains of 0.1M NaOH).

a) Moles of NaOH that reacted

1000cm^3 of solution contains 0.1moles of NaOH

1 cm^3 of solution contains ($^{0.1}/_{1000}$) moles of NaOH

25 cm^3 of solution contains ($^{25x0.1}/_{1000}$) moles of NaOH

=0.0025 moles of NaOH

3. Work out the number of moles of the acid that reacted by relating the number of moles of the alkali to the mole ratio of reaction between the acid and alkali.

b) From the equation of reaction

1 mole of NaOH reacts with 1 mole of HCl

0.0025 moles of NaOH reacts with ($^{0.0025}/_{1}$) moles of HCl

=0.0025 moles of HCl

(since the mole ratio of the reaction of the HCl : NaOH is 1:1, so, the number of moles of HCl = 0.0025 moles)

4. Now work out the molarity of the acid

c) 15cm^3 of solution contains 0.0025moles of HCl

1cm^3 of solution contains ($^{0.0025}/_{15}$) moles of HCl

1000cm^3 of solution contains ($^{1000x0.0025}/_{15}$) moles of HCl

=0.167 mol/l of HCl

The molarity of the HCl is 0.167M

5. From the molarity, you can proceed and work out the concentration in grams/litre, if the Molar mass is known.(e.g. for 0.167M HCl in the above case)

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

   1 mole of HCl weighs 36.5g

   0.167 moles of HCl weighs ($^{0.167x36.5}/_{1}$)g

   =6.1g/l

Other examples

1. 25cm^3 of sulphuric acid of concentration 0.15mol/dm^3 neutralized 31.2cm^3 of potassium hydroxide solution. Find the concentration of the KOH solution in mol/l and in grams/litre.

Solution

Write the equation

Equation for the reaction

H_2SO_4 (aq) + 2KOH(aq) $\longrightarrow$ K_2SO_4(aq) + H_2O(l)

1 mole of acid reacts with 2 moles of alkali

Calculate the number of moles of the acid (standard solution) that reacted

1000cm^3 of solution contains 0.15moles of H_2SO_4

1 cm^3 of solution contains ($^{0.15}/_{1000}$) moles of H_2SO_4

25 cm^3 of solution contains ($^{25x0.15}/_{1000}$)moles of H_2SO_4

=0.00375 moles of H_2SO_4

Relate the number of moles of acid to the mole ratio of the reaction to find the number of moles of the alkali that reacted

1 mole of H_2SO_4 reacts with 2 moles of KOH

0.00375moles of H_2SO_4 reacts with ($^{0.00375x2}/_{1}$) moles of KOH

=0.0075 moles of KOH

Calculate the molarity (concentration in mol/dm^3) of the alkali

31.2cm^3 of solution contains 0.0075moles of KOH

1cm^3 of solution contains ($^{0.0075}/_{31.2}$) moles of KOH

1000cm^3 of solution contains ($^{1000x0.0075}/_{31.2}$) moles of KOH

=0.24 mol/l of KOH

The concentration of the KOH is 0.24 mol/l of KOH

Proceed and calculate the concentration in g/l

Molar mass of KOH= (39+16+1)= 56g

1 mole of KOH weighs 56g

0.24 moles of KOH weighs ($^{0.24x56}/_{1}$)g

=13.44g

The concentration of the KOH is 13.44g/l

2. Determine the molarity of acids in the following solutions.
    a) 16.0cm^3 of 1.5M sodium hydroxide neutralized by 20.0cm^3 of hydrochloric acid
    b) 25.0cm^3 of 0.2M ammonia solution neutralized by 20.0cm^3 of nitric acid

Solution

a) Equation of reaction

$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(aq)$

1000cm^3 of solution contains 1.5moles of NaOH

1 cm^3 of solution contains ($^{1.5}/_{1000}$) moles of NaOH

16 cm^3 of solution contains ($^{16x1.5}/_{1000}$)moles of NaOH

=0.024 moles of NaOH

From the equation

1 mole of NaOH reacts with 1 mole of HCl,

0.024 moles of NaOH reacts with ($^{0.024x1}/_{1}$)moles of HCl

=0.024 moles of HCl

20.0cm^3 of solution contains 0.024moles of HCl

1cm^3 of solution contains ($^{0.024}/_{20}$) moles of HCl

1000cm^3 of solution contains ($^{1000x0.024}/_{20}$) moles of HCl

=1.2 M HCl

The concentration of the HCl is 1.2M

Equation for reaction

b) $HNO_3(aq) + NH_4OH(aq) \longrightarrow NH_4NO_3(aq) + H_2O(aq)$

1000cm^3 of solution contains 0.2moles of NH_4OH

1 cm^3 of solution contains ($^{0.2}/_{1000}$) moles of NH_4OH

25cm^{3}of solution contains ($^{25x0.2}/_{1000}$)moles of NH_4OH

=0.005 moles of NH_4OH

From the equation

1 mole of NH_4OH reacts with 1 mole of HNO_3

0.005moles of NH_4OH reacts with ($^{0.005x1}/_{1}$)moles of HNO_3

=0.005 moles of HNO_3

20.0cm^3 of solution contains 0.005moles of HNO_3

1cm^3 of solution contains ($^{0.005}/_{20}$) moles of HNO_3

1000cm^3 of solution contains($^{1000x0.005}/_{20}$)moles of HNO_3

=0.25 M HNO_3

The concentration of the HNO_3 is 0.25M

3. 5.0g of a mixture of sodium chloride and anhydrous sodium carbonate were made up to 200cm^3 of aqueous solution. 25cm^3 of this solution required 40cm^3 of 0.1M of hydrochloric acid for neutralization. What is the percentage by mass of the anhydrous sodium carbonate in the mixture?

Solution

Equation for the reaction

$2HCl(aq) + Na_2CO_3(aq) \longrightarrow NaCl(aq) + H_2O(aq) + CO_2(g)$

1000cm^3 of solution contains 0.1moles of HCl

1 cm^3 of solution contains ($^{0.1}/_{1000}$) moles of HCl

40cm^3 of solution contains ($^{40x0.1}/_{1000}$)moles of HCl

=0.004 moles of HCl

From the equation

2 mole of HCl reacts with 1 mole of Na_2CO_3

0.004 moles of HCl reacts with ($^{0.004x1}/_{2}$)moles of Na_2CO_3

=0.002 moles of Na_2CO_3

25.0cm^3 of solution contains 0.002moles of Na_2CO_3

1cm^3 of solution contains ($^{0.002}/_{25}$) moles of Na_2CO_3

1000cm^3 of solution contains ($^{1000x0.002}/_{25}$) moles of Na_2CO_3

=0.08 M Na_2CO_3

The molarity of the Na_2CO_3 is 0.08M

RMM of Na_2CO_3

=23x2+12+16x3

=106

1 mole of Na_2CO_3 weighs 106g

0.08 moles of Na_2CO_3 weighs ($^{0.08x106}/_{1}$) g

=8.48g of Na_2CO_3

The mass of Na_2CO_3 in 1 litre is 8.48g

200cm^3 of solution contains 5.0g of mixture

1cm^3 of solution contains ($^{5.0}/_{200}$)g of mixture

1000cm^3 of solution contains ($^{5.0x1000}/_{200}$)g of mixture

=25g of the mixture.

The mass of a mixture of sodium carbonate and sodium chloride in 1 litre is 25g.

Percentage mass of $Na_2CO_3 = \frac{mass\ of\ sodium\ carbonate}{mass\ of\ mixture} x\ 100$

$= \frac{8.48}{25} x100$

=34%

4. 3.5g of a mixture of K_2CO_3 and K_2SO_4 were made up to 250cm^3 of aqueous solution. 25cm^3 of the solution required 24.6cm^3 of 0.1 HCl for complete neutralization.
   a) Write the equation for the reaction
   b) Determine the percentage by mass of potassium sulphate in the mixture.

Solution

a) $2HCl(aq) + K_2CO_3(aq) \longrightarrow 2KCl(aq) + H_2O(l) + CO_2(g)$

b) 1000cm^3 of solution contains 0.1moles of HCl

1 cm^3 of solution contains $(^{0.1}/_{1000})$ moles of HCl

24.6cm^3 of solution contains $(^{24.6x0.1}/_{1000})$moles of HCl

=0.00246 moles of HCl

From the equation

2 mole of HCl reacts with 1 mole of K_2CO_3

0.00246 moles of HCl reacts with $(^{0.00246x1}/_{2})$moles of K_2CO_3

=0.00123 moles of K_2CO_3

25.0cm^3 of solution contains 0.00123moles of K_2CO_3

1cm^3 of solution contains $(^{0.00123}/_{25})$ moles of K_2CO_3

1000cm^3 of solution contains $(^{1000x0.00123}/_{25})$ moles of K_2CO_3

=0.0492 M K_2CO_3

The molarity of the K_2CO_3 is 0.0492M

RMM of K_2CO_3

=39x2+12+16x3

=138

1 mole of K_2CO_3 weighs 138g

0.0492 moles of K_2CO_3 weighs $(^{0.0492x138}/_{1})$ g

=6.79g of K_2CO_3

The mass of K_2CO_3 in 1 litre is 6.79g

250cm³ of solution contains 3.5 g of mixture

1cm³ of solution contains $(^{3.5}/_{250})$g of mixture

1000cm³ of solution contains $(^{3.5x1000}/_{250})$g of mixture

=14g of the mixture.

The mass of a mixture of potassium carbonate and potasium sulphate in 1 litre is 14g.

Mass of K_2SO_4= mass of mixture-mass of K_2CO_3

=(14-6.79)g= 7.21g

Percentage by mass of K_2SO_4=$\frac{7.21}{14}$x 100

=51.5%

Calculating number of moles of water of crystallization

The water of crystallization present in compounds does not take part in the reaction

1. Calculate the number of moles of water of crystallization in oxalic acid crystals, $H_2C_2O_4.xH_2O$ from the following data.
   5.0g 0f the crystals were made up to 250cm3 of aqueous solution and 25cm3 of this solution required 15.9 cm3 of 0.5M NaOH solution to neutralize it.

Solution

Equation for the reaction

$2NaOH(aq) + H_2C_2O_4(aq) \longrightarrow Na_2C_2O_4(aq) + 2H_2O(l)$

1000cm³ of solution contains 0.5moles of NaOH

1 cm³ of solution contains $(^{0.5}/_{1000})$ moles of NaOH

15.9 cm^3 of solution contains ($^{15.9x0.5}/_{1000}$)moles of NaOH

=0.00795 moles of NaOH

From the equation

2 mole of NaOH reacts with 1 mole of $H_2C_2O_4$,

0.00795 moles of NaOH reacts with ($^{0.00795x1}/_{2}$) moles of $H_2C_2O_4$

=0.003975 moles of $H_2C_2O_4$

25.0cm^3 of solution contains 0.003975moles of $H_2C_2O_4$

1cm^3 of solution contains ($^{0.003975}/_{25}$) moles of $H_2C_2O_4$

1000cm^3 of solution contains ($^{1000x0.003975}/_{25}$) moles of $H_2C_2O_4$

=0.159 M $H_2C_2O_4$

The concentration of the $H_2C_2O_4$is 0.159mol/dm^3

250cm^3 of solution dissolves 5 g of oxalic acid

1cm^3 of solution contains ($^{5}/_{250}$)g of oxalic acid

1000cm^3 of solution contains ($^{5x1000}/_{250}$)g of oxalic acid

=20g/l of the oxalic acid.

The mass of a oxalic acid in 1 litre is 20g.

RMM of the oxalic acid

0.159 moles of $H_2C_2O_4.xH_2O$ weighs 20g

1 mole of $H_2C_2O_4.xH_2O$ weighs($^{20}/_{0.159}$)g

=125.79g

The RMM of $H_2C_2O_4.XH_2O$ is 125.79≈ 126

There fore

$H_2C_2O_4.xH_2O$ = 126

1x2+12x2+16x4+ x(1x2+16) =126

90+18x =126

18x=126-90

18x=36

X=2.

The formula of the oxalic acid is $H_2C_2O_4.2H_2O$

2. 0.465g of a hydrated form of sodium carbonate exactly reacts with 75cm3 of 0.10M hydrochloric acid. Calculate the number of moles of water of crystallization present in one mole of the hydrated salt.

Solution

Equation for the reaction

$2HCl(aq) + Na_2CO_3(aq) \longrightarrow NaCl(aq) + H_2O(aq) + CO_2(g)$

1000cm^3 of solution contains 0.1moles of HCl

1 cm^3 of solution contains $(^{0.1}/_{1000})$ moles of HCl

75cm^3 of solution contains $(^{75x0.1}/_{1000})$ moles of HCl

=0.0075 moles of HCl

From the equation

2 mole of HCl reacts with 1 mole of Na_2CO_3

0.0075 moles of HCl reacts with $(^{0.0075x1}/_{2})$ moles of Na_2CO_3

=0.00375 moles of Na_2CO_3

RMM of Na_2CO_3

=23x2+12+16x3

=106

1 mole of Na_2CO_3 weighs 106g

0.00375 moles of Na_2CO_3 weighs $(^{0.00375x106}/_{1})$ g

=0.3975g of Na_2CO_3

Mass of water contained=(0.465-0.3975)g=0.0675g

The mole ratio of $Na_2CO_3 : H_2O$ is obtained by dividing each by the molar mass of the compound i.e. mole ratio of $Na_2CO_3 : H_2O = \frac{0.3975}{106} : \frac{0.0675}{18} = 1:1$

The formula is $Na_2CO_3.H_2O$

Calculating basicity of acids

Basicity of an acid is the number of hydrogen ions that can be produced by one molecule of the acid upon complete ionization. It is not the number of hydrogen atoms in one molecule of an acid.

Example

1. Calculate the basicity of an acid H_nZ if 15cm^3 of 0.1M acid is completely neutralized by 9cm^3 of 0.5 M potassium hydroxide.

Number of moles of the acid

1000cm^3 of solution contains 0.1moles of H_nZ

1 cm^3 of solution contains $({}^{0.1}/_{1000})$ moles of H_nZ

15cm^3 of solution contains $({}^{15x0.1}/_{1000})$ moles of H_nZ

=0.0015 moles of H_nZ

Number of moles of KOH

1000cm^3 of solution contains 0.5moles of KOH

1 cm^3 of solution contains $({}^{0.5}/_{1000})$ moles of KOH

9cm^3 of solution contains $({}^{9x0.1}/_{1000})$ moles of KOH

=0.0045 moles of KOH

Equation of reaction

$H_nZ(aq) + nKOH(aq) \longrightarrow K_nZ(aq) + nH_2O(l)$

Reaction ration, $\frac{number\ of\ moles\ of\ acid}{number\ of\ moles\ of\ base} = \frac{1}{n}$

$$\frac{0.0015}{0.0045} = \frac{1}{n}$$

$$0.0015n=0.0045$$

$$n= 3$$

Exercise

1. 11.2 g of potassium hydroxide were made up to 1 litre of aqueous solution. 25cm^3 of this solution required 24.9cm^3 of 0.1 M of an acid H_nX. Dertermine the value of n in the acid(i.e the basicity of the acid) (**Answer=2**)
2. 1.5g of sodium hydroxide contained in 250cm^3 of solution was used to titrate 0.1M hydrochloric acid. What volume of acid would be needed to react with 20cm^3 of the alkali? (**Answer=30cm^3**)

3. 8.50g of a sample of iron required just 75 cm^3 of 3.00M hydrochloric acid to dissolve it and give a neutral solution. Calculate the percentage purity of the sample of iron.(**Answer=74.1%)**
4. Determine the number of moles of nitrate ions in 1 litre solution,if 6.62g of Lead(II)nitrate are made up to $200cm^3$ of aqueous solution.(**Answer=0.2 moles**)
5. Kyagulanyi and Atim were playing and accidentally poured an acid that was on the table in the drinking water for students. Samanya and Dickens, young analytical chemists titrate the water and find that $10.0dm^3$ of water are needed to neutralize $10.0cm^3$ of a $0.010moldm^3$ solution of sodium hydroxide. What is the concentration of the hydrogen ions in the water.(**Answer 0.00005M**)
6. Sodium carbonate crystals (27.823g) were dissolved in water and made up to 1 litre. 25.0 cm^3 of the solution were neutralized by 48.8cm3 of hydrochloric acid of concentration 0.10M. Find n in the formula $Na_2CO_3.nH_2O$.(**Answer n=10**)

Sample questions on mole concept

1. 25 cm^3 of 0.1M NaOH(aq) *required $12.5cm^3$ of Sulphuric acid solution for complete reaction. Calculate the.*
  i) Molarity of the acid
  ii) Concentration of the acid in g/l.

2. $25cm^3$ of a solution containing 5.3g/l of a metal carbonate of formula, M_2CO_3 *neutralized $25cm^3$ of 0.1M Hydrochloric acid. Calculate the*
  i) Molarity of the carbonate solution
  ii) Formula mass of the metal carbonate
  iii) Relative atomic mass of the metal (M) in the metal carbonate.

3. $25cm^3$ of impure sodium hydroxide solution containing 5.0g of impure sodium hydroxide per litre completely reacted with $20cm^3$ of 0.1M Hydrochloric acid. Calculate the;
  I) molarity of Sodium Hydroxide solution
  II) Percentage impurity in the sample of sodium Hydroxide.

4. 1.00g of a mixture of calcium carbonate was carefully shaken with $200cm^3$ of distilled water, and then filtered. $20cm^3$ of the filtrate required $8cm^3$ of 0.125M Hydrochloric acid for neutralization. Calculate

i) *The mass of Calcium carbonate in the sample*

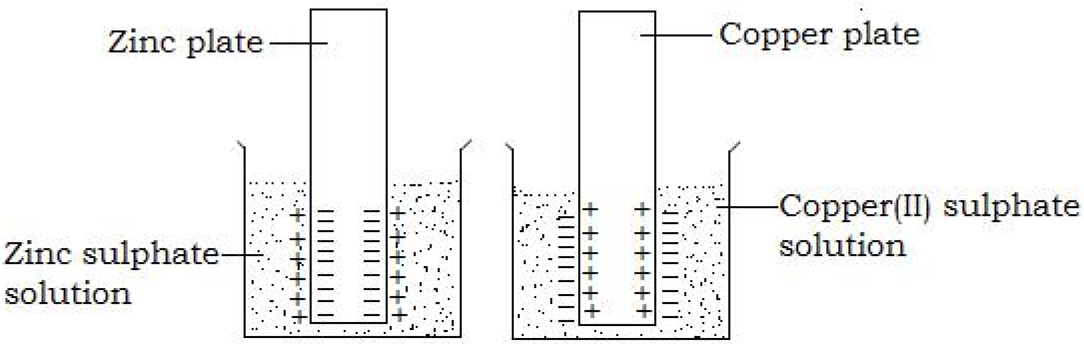
ii) *Percentage purity of sodium carbonate in the mixture*

5. *25cm^3 of sodium carbonate required 27.25cm^3 of a solution containing 4g of Hydrochloric acid per litre. Calculate the concentration of sodium carbonate in grams per litre of solution.*
6. *25cm^3 of a solution containing 12.0g of sodium hydroxide per litre, required 15cm^3 of a solution containing 18.25g/l of acid, HA. Calculate the formula mass of the acid.*
7. *20cm^3 of 0.12M Potassium hydroxide required 24 cm^3 of a solution of a dibasic acid,* H_2X *containing 1.225g per 250cm^3 of solution. Calculate the formula mass of the acid.*
8. *32.5cm^3 of a solution containing 6.1 g/l of an acid,* HXO_3, *required 25cm^3 of a solution containing 5.0g/l of sodium hydroxide solution. Calculate the relative atomic mass of the acid.*
9. *20cm^3 of a solution containing 25.8g/l of a dibasic hydrated acid,* $H_2X.nH_2O$ *neutralized 10 cm^3 of 0.8M sodium Hydroxide solution. If the relative formula mass of* $H_2X = 90$, *determine the value of n*
10. *20cm^3 of nitric acid solution containing 6.3g/l of solution required 25 cm^3 of a metal carbonate,* M_2CO_3, *solution containing 5.52 g/l of solution for neutralization. Calculate the relative mass of metal* M.

ELECTROCHEMICAL CELLS

In electrochemical cells, electricity is produced from chemical reactions i.e. chemical energy is converted to electric energy. It consists of two half cells or electrodes and at each electrode/half cell, an element is in contact with a solution of its ions.

For example, consider zinc plate dipped into a solution containing zinc ions and copper plate dipped in a solution of copper ions as below.



In such systems, there are to tendencies:

i) The metal (zinc) tends to dissolve to form ions. The electrons remain behind on the metal surface (zinc surface) and so it becomes negatively charged because of the excess electrons.

Equation

$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$

So, there exists a potential difference between the metal and its solution and this is called the **electrode potential.** The electrode potential measures the reactivity of the metal.

ii) The metal ions (copper ions) gain electrons and get deposited on the metal surface.

Equation

$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

The metal surface there fore loses electrons and becomes positively charged. There will be a potential difference between the metal surface and its solution.

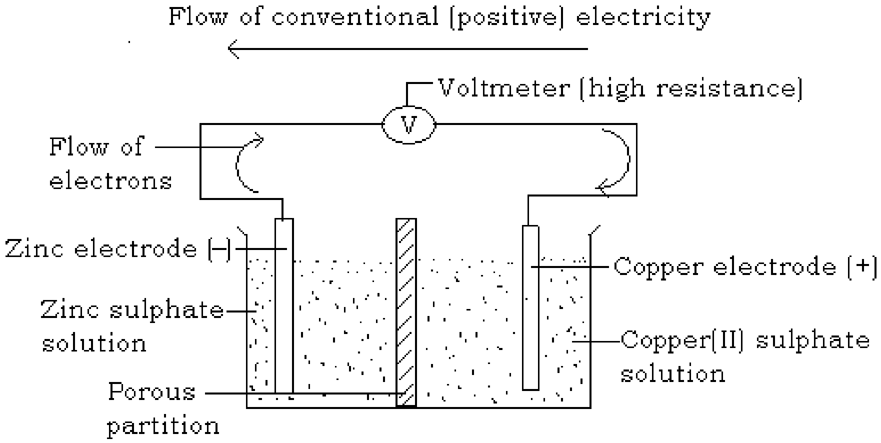
The above reactions depend on the reactivity of the metals. Metals which are more reactive e.g. zinc have a greater tendency to dissolve as ions and a smaller tendency to be deposited as metal. While metals that are less reactive like copper have greater tendencies to be deposited as metal and small tendency to dissolve and form ions. Zinc there fore acquires a more negative potential than copper.

The tendency of a metal to dissolve when in solution of its own ions can be measured by joining two half cells to from an electrochemical cell/galvanic or voltaic cell.

Connection between the two half cells

i) The two half cells can be connected by a porous partition which allows conduction of electricity by movement of ions across without physically mixing the solutions.

For example the Daniel cell (a primary cell) which consists of zinc and copper metals in solutions of their ions combined as shown below.



The zinc electrode is on the **left hand side** as it has a more negative potential and copper on the **right hand side** as it has a more positive potential. Current flows from the positive electrode to the negative electrode and electrons flow from negative to positive electrode.

Reaction at the electrodes

At Zinc (negative) electrode

Zinc dissolves from the zinc electrode forming zinc ions and electrons are released.

$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$

At copper (positive) electrode

Copper ions gain electrons and copper is deposited on the electrode.

$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

Overall reaction

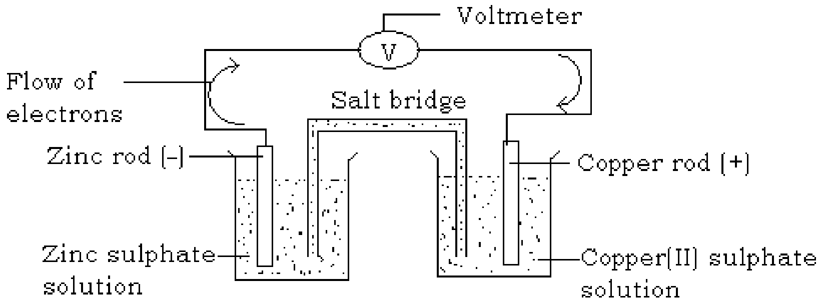
$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$

The cell notation of the reaction is $Zn(s)/Zn^{2+}(aq) /\!/ Cu^{2+}(aq)/ Cu(s)$

$/\!/$ Represents the separation of the two half cell by a porous medium.

Since at the positive electrode copper ions are being converted to copper solid, the blue colour of the solution fades and changes to colourless.

ii) The two half cell can be connected using a salt bridge. A salt bridge is a bent glass tube containing sodium or potassium salts from strong acids. E.g. potassium chloride and sodium nitrate or pieces of filter paper wetted with these salts.



The reactions at the electrodes are the same as the above.

The cell notation is is Zn(s)/Zn^{2+}(aq)//Cu^{2+}(aq)/ Cu(s)

// represents the separation of the two half cells by a salt bridge.

The force with which the electrons flow between the two half cells creates a potential difference between them. This force is called the **electromotive force (E.m.f.).**

$E.m.f_{(cell)} = E_{(right)} - E_{(left)}$

$E_{(right)}$ = electrode potential at right hand side
$E_{(left)}$ = electrode potential at left hand side

The lead-acid battery (secondary cell-rechargeable cell)

The lead-acid battery is used in vehicles to power the starter motor. It consists of six 2V cells connected in series. Each cell consists of two lead plates dipping into a 30% solution of sulphuric acid. One of the plates is coated with lead (IV) oxide, PbO_2. The lead plate and the lead (IV) oxide- coated plate have different electrode potentials so when they are connected, a current flows between them. The electrode processes are:

Positive plate: $2PbO_2(s) + 2H^+(aq) + SO_4^{2-}(aq) + 2e^- \longrightarrow PbSO_4(s) + 2H_2O\ (l)$

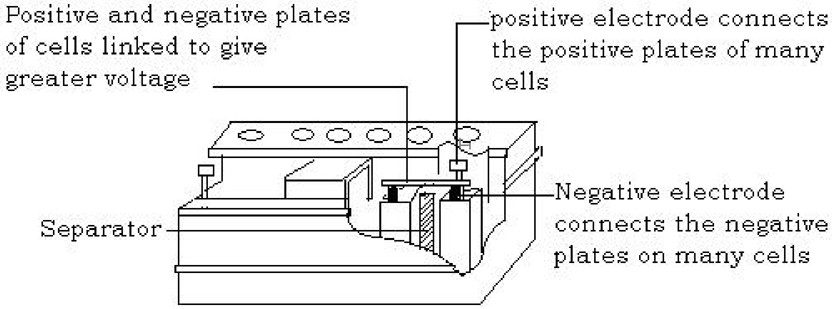
Negative electrode: $Pb(s) + SO_4^{2-}(aq) \longrightarrow PbSO_4(s) + 2e^-$

When PbO_2 and Pb have been converted into $PbSO_4$, there is no difference between the plates, and the cell can no longer give a current. The battery must be charged.

The battery is charged by passing a direct electric current through it. This reverses the electrode processes in the cells to produce lead and lead (IV) oxide. The battery can again give a current. When the vehicle is in motion, it drives a generator which charges the battery. If there is too much stopping and starting, the battery loses its charge and becomes flat until it is recharged.

During discharge, sulphuric acid is formed. During charge, it is used up. Since sulphuric acid is much denser than water, the state of the battery can be assessed by measuring the density of the battery acid with a hydrometer.

Part of a lead-acid battery



REDOX REACTIONS AND IONIC EQUATIONS

These are reactions in which both reductions and oxidations take place simultaneously.

Oxidation

Can be defined as;

i) Addition of oxygen to a substance

ii) Removal of hydrogen from a substance

iii) Loss of electrons from a substance

NB. Oxidizing agents are electron acceptors. Therefore non metals are oxidizing agents.

An oxidizing agent: is a substance which accepts electrons.

An example of an oxidation reaction is conversion of metal atoms to ions by loss of electrons.

Na(s) to Na^+: $Na\ (s) - e^- \longrightarrow Na^+$ (Oxidation reaction)

Mg(s) to Mg^{2+}: Mg(s) - $2e^- \longrightarrow Mg^{2+}$(aq) (Oxidation reaction)

Reduction

Can be defined as;

i) Removal of oxygen from a substance

ii) The addition of hydrogen to a substance

iii) Gain of electrons by a substance

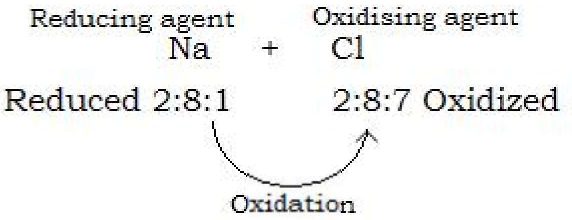
NB. Reducing agents are electron donors; therefore all metals are reducing agents.

A reducing agent: is a substance which donates electrons.

An example of reduction reaction is conversion of non metal atoms to ions by gain of electrons.

Example. $Cl + e^- \longrightarrow Cl^-$

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



Cl + $e^- \longrightarrow Cl^-$ (Reduction)

↑

Electron acceptor

$\longrightarrow$

↑

Electron donor

Na^+ (Oxidation)

Oxidation numbers

Oxidation numbers are combining powers of substances with oxygen. The state of oxidation in an element in a given compound is indicated by its oxidation number

Simple rules when calculating oxidation numbers

1. All elements in a free state [uncombined state] have an oxidation number of zero.

2. In the case of simple ions, the element has an oxidation number with the same size and sign of the charge on the ion. For example Cu^{2+} has an oxidation number of +2 and S^{2-} has an oxidation number of -2.

3. The sum of all oxidation numbers of the elements in the compound is zero.

Example

Calculate or determine the oxidation numbers of stated elements in the following

i) S in SO_2

ii) S in SO_3

iii) H in H_2O

iv) Mn in $MnO_4{}^-$

v) N in $NO_3{}^-$

vi) Mn in MnO_2

Solutions

i) Sulphur in Sulphur dioxide

$$SO_2 = 0$$

$$S + (-2X2) = 0$$

$$S - 4 = 0$$

$$\Rightarrow \quad S = {}^{+}4$$

ii) Sulphur in Sulphur trioxide

$$SO_3 = 0$$

$$S + (-2X3) = 0$$

$$S - 6 = 0$$

$$\Rightarrow \quad S = {}^{+}6$$

iii) Hydrogen in water

$$H_2O = 0$$

$$H\,X\,2 + -2 = 0$$

$$2H - 2 = 0$$

$$\Rightarrow \quad H = {}^{+}1$$

iv) Manganese in Permanganate ion

$$MnO_4{}^- = -1$$

$$Mn + (-2X4) = -1$$

$$Mn - 8 = -1$$

$$\Rightarrow \quad Mn = {}^{+}7$$

v) Nitrogen in Nitrate

$$NO_3{}^- = -1$$

$$N\ X\ (-2X3) = -1$$

$$N - 6 = -1$$

⇨ $N = {}^{+}5$

vi) Manganese in Manganese (IV) Oxide

$$Mn\ _2 = 0$$

$$Mn + (-2X2) = 0$$

$$Mn - 4 = 0$$

⇨ $Mn = {}^{+}4$

Exercise

Calculate or determine the oxidation numbers of stated elements in the following;

i) Sulphur in Sulphite ion
ii) Hydrogen in Hydrochloric acid
iii) Phosphorous in phosphate
iv) Carbon in Carbon dioide
v) Carbon in cabon monoxide
vi) Sulphur in Sulphur dioxide
vii) Oxygen in Hydroxyl ion
viii) Chromium in Dichromate ion
ix) Copper in Copper (II) oxide
x) Manganese in Permanganate (VII) ions

Examples

State what is taking place and write half reaction equation in each case.

NB. Oxidation is an increase in oxidation number and reduction is a decrease in the oxidation number.

i) The conversion of hydrogen ions (H+) to hydrogen molecules (H2)

$$2H^{+}(aq) + 2e^{-} \longrightarrow H_2\ (g)$$

This is reduction because H^{+} ions has gained electrons

ii) Conversion of Iron (II) ions (Fe^{2+}) to Iron (III) ions Fe^{3+}

$$Fe^{2+}(aq) - e^{-} \longrightarrow Fe^{3+}(aq)$$

This is oxidation because Fe^{2+} *ions has lost electrons*

iii) Conversion of Copper (II) ions Cu^{2+} to Copper, Cu

$Cu^{2+}(aq) + 2e^{-} \longrightarrow Cu(s)$

This is reduction because Cu^{2+} *ions has gained electrons*

iv) Conversion of Zinc to Zinc ions, Zn^{2+}

$Zn(s)\ -2e^{-} \longrightarrow Zn^{2+}(aq)$

This is oxidation because Zn has lost electrons

Other redox reactions

1. Reactions of metals with acids

Example

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

$Zn(s) + 2H^{+}, \cancel{2Cl^{-}} \longrightarrow Zn^{2+}, \cancel{2Cl^{-}} + H_{2(g)}$

(Ionic equation) $Zn(s) + 2H^{+}(aq) \longrightarrow Zn^{2+}(aq) + H_2(aq)$

$Zn(s)\ -2e^{-} \longrightarrow Zn^{2+}$ (oxidation half reaction)……………………………(i)

$2H^{+}\ (aq) + 2e^{-} \longrightarrow H_{2(g)}$ (Reduction half reaction)……………………(ii)

Redox overall equations: $Zn(s) + 2H^{+}\ (aq) \longrightarrow Zn^{2+}\ (aq) + H_2\ (aq)$ i.e (i) + (ii)

2. Displacement reactions as redox reactions

(a) Metals

A displacement reaction is one in which one element takes the place of another element in a chemical reaction. Example, metals high up in the reactivity series [more reactive metals] tend to display those below them from their solutions [compounds]. In this case, the more reactive metals act as reducing agents [electron donors] and their reducing power decreases as reactivity decreases.

Example

1. When zinc dust is added to a solution of copper [II] sulphate, the copper [II] ions are displaced from a solution to form a brown precipitate[solid] and the blue colour of the solution gradually fades

$Zn(s) + CuSO_4 (aq) \longrightarrow ZnSO_4 (aq) + Cu(s)$

(Blue solution) (Colourless) (Brown)

$Zn(s) + Cu^{2+}, \cancel{SO_4^{2-}} (aq) \longrightarrow Zn^{2+}(aq) + \cancel{SO_4^{2-}} (aq) + Cu(s)$

$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$

i) $Zn(s) - 2e^{-} \longrightarrow Zn^{2+}(aq)$ (oxidation half reaction)......................(i)

ii) $Cu^{2+}(aq) + 2e^{-} \longrightarrow Cu(s)$ (Reduction half reaction).................(ii)

Overall redox equation: $Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$

Example

Show that the following reactions are redox reactions. In each case write half reactions and overall redox reaction and state the observation(s) made if any.

i) Iron in Copper (II) sulphate
ii) Magnesium and Zinc (II) nitrate
iii) Silver and Copper (II) Sulphate
iv) Zinc and Lead (II) nitrate

b) Displacement reactions of halogens

Halogens are displaced from their solutions by other halogens which are higher in the electro chemical series [more reactive halogens]. The order of reactivity of halogens is Flourine >Chlorine>Bromine>Iodine>Astatine.

For example

When chlorine gas is passed through an aqueous colourless solution of potassium iodide, a brown colouration develops immediately. The brown colouration is due to iodine liberated as iodide ions are displaced by chlorine from solution example;

$$Cl_2\ (g) + 2KI\ (aq) \longrightarrow 2KCl\ (aq) + I_2\ (l)$$

$$Cl_2\ (g) + \cancel{2K^+} + 2I^-(aq) \longrightarrow \cancel{2K^+}\ (aq) + 2Cl^-(aq) + I_2(aq)$$

$$Cl_2\ (g) + 2I^-(aq) \longrightarrow 2Cl^-(aq) + I_2(aq)$$

Cl_2 (g) + $2e^-$(aq) $\longrightarrow$ $2Cl^-(aq)$ (Reduction half reaction)..........(i)

$2I^-$ (aq) - $2e^-$(aq) $\longrightarrow$ I_2 (l) (Oxidation half reaction)...............(ii)

Overall Redox equation: $Cl_2\ (g) + 2I^-(aq) \longrightarrow 2\ Cl^-(aq) + I_2(aq)$

Chlorine in the above case acts as the oxidizing agent and the oxidizing power of the halogens decreases with their decreasing reactivity.

IONIC EQUATIONS

The ionic equation is used to describe the chemical reaction while also clearly indicating which of the reactants and/or products exist primarily as ions in aqueous solution.

Rules in writing ionic equations

1. Start with a balanced molecular equation.

A molecular equation is one that shows the chemical formulas of all reactants and products but does not expressly indicate their ionic nature.

For example,

$$2\ Na_3PO_4\ (aq) + 3\ CaCl_2\ (aq) \longrightarrow 6\ NaCl\ (aq) + Ca_3(PO_4)_2\ (s)$$

2. **Break all soluble strong electrolytes (compounds with (aq) beside them) into their ions**
    - indicate the correct formula and charge of each ion
    - indicate the correct number of each ion
    - write (aq) after each ion
3. **Bring down all compounds with (s), (l), or (g) unchanged**

$6\ Na^+(aq) + 2PO_4^{3-}(aq) + 3Ca^{2+}(aq) + 6Cl^-(aq) \longrightarrow 6Na^+\ (aq) + 6Cl^-\ (aq) + Ca_3(PO_4)_2(s)$

How did I get this equation?

Consider each reactant or product separately:

1 mole of Na_3PO_4 contains 3 moles of Na^+ and 1 mole of PO_4^{3-}. Since the balanced equation shows that two moles of sodium phosphate are involved in the reaction, a total of 6 moles (2 x 3) of Na^+ and 2 moles (2 x 1) of PO_4^{3-} are formed. Notice that the subscript "4" in the formula for the phosphate ion is not used when determining the number of phosphate ions present. That particular subscript is part of the formula for the phosphate ion itself.

$2\ Na_3PO_4\ (aq) \longrightarrow 6\ Na^+\ (aq) + 2\ PO_4^{3-}\ (aq)$

1 mole of $CaCl_2$ contains 1 mole of Ca^{2+} and 2 moles of Cl^-. Remember, the subscript "2" indicates the number of chloride ions. Cl_2 is elemental chlorine. You will never have a diatomic chlorine ION (i.e. Cl_2^- or Cl_2^{2-}) in aqueous solution. Since the balanced equation shows that 3 moles of calcium chloride are involved in the reaction a total of 3 moles (3 x 1) of Ca^{2+} and 6 moles (3 x 2) of Cl^- are formed.

$3\ CaCl_2\ (aq) \longrightarrow 3\ Ca^{2+}\ (aq) + 6\ Cl^-\ (aq)$

1 mole of NaCl contains 1 mole of Na^+ and 1 mole of Cl^- . Since the balanced equation shows that 6 moles of NaCl are produced by the reaction, 6 moles (6 x 1) of Na^+ and 6 moles (6 x 1)of Cl^- will be formed.

$6\ NaCl\ (aq) \longrightarrow 6\ Na^+\ (aq) + 6\ Cl^-\ (aq)$

Since calcium phosphate is an insoluble solid (indicated by the (s) beside its formula), it will not form appreciable amounts of ions in water. It is brought down unchanged into the complete ionic equation.

4. Cross out the spectator ions that are present

Spectator ions are ions that are present in the reaction mixture but do not participate in it. They "sit around and watch the reaction take place" just like a spectator at a football game watches the players in the game but doesn't play the game himself. You can recognize spectator ions by looking for ions that are present on both sides of the equation. They will always have the same exact formula, charge, and physical state. They will also be present in exactly the same number on both sides of the equation.

$$\cancel{6\ Na^+(aq)} + 2PO_4^{3-}(aq) + 3Ca^{2+}(aq) + \cancel{6Cl^-(aq} \longrightarrow \cancel{6\ Na^+(aq)} + \cancel{6\ Cl^-(aq)} + Ca_3(PO_4)_2(s)$$

5. Write the "leftovers" as the net ionic equation

$$2\ PO_4^{3-}\ (aq) + 3\ Ca^{2+}\ (aq) \longrightarrow Ca_3(PO_4)_2\ (s)$$

The ions that appear in the final net ionic equations are called participants as they are the ones that are actively involved in the reaction.

Example 1:

Write the ionic equation for the word equation

Sodium chloride(aq) + silver nitrate(aq) → silver chloride(s) + sodium nitrate(aq)

Solution:

Step 1: Write the equation and balance it.

$NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

Step 2: Split the ions. (Only compounds that are aqueous are split into ions.)

$Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-$

Step 3: Cancel out spectator ions.

$$\cancel{Na^+(aq)} + Cl^-(aq) + Ag^+(aq) + \cancel{NO_3^-(aq)} \rightarrow AgCl(s) + \cancel{Na^+(aq)} + \cancel{NO_3^-}$$

Step 4: Write a balanced ionic equation

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

Example 2:

Write the ionic equation for the word equation

Sodium(s) + hydrochloric acid(aq) -> sodium chloride(aq) + hydrogen(g)

Solution:

Step 1: Write the equation and balance it.

$2Na(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2(g)$

Step 2: Split the ions. (Only compounds that are aqueous are split into ions.)

$2Na(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow 2Na^+(aq) + 2Cl^-(aq) + H_2(g)$

Step 3: Cancel out spectator ions.

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

Step 4: Write a balanced ionic equation

$2Na(s) + 2H^+(aq) \rightarrow 2Na^+(aq) + H_2(g)$

Exercise

Write the complete ionic and net ionic equations for each of the following reactions:

1. $3\ (NH_4)_2CO_3\ (aq) + 2\ Al(NO_3)_3\ (aq) \longrightarrow 6\ NH_4NO_3\ (aq) + Al_2(CO_3)_3\ (s)$
2. $2\ NaOH\ (aq) + H_2SO_4\ (aq) \longrightarrow Na_2SO_4\ (aq) + 2\ H_2O\ (l)$
3. $Mg\ (s) + 2\ HCl\ (aq) \longrightarrow MgCl_2\ (aq) + H_2\ (g)$
4. $Zn\ (s) + CuSO_4\ (aq) \longrightarrow ZnSO_4\ (aq) + Cu\ (s)$
5. $Na_2CO_3\ (aq) + 2\ HNO_3\ (aq) \longrightarrow 2\ NaNO_3\ (aq) + H_2O\ (l) + CO_2\ (g)$

QUALITATIVE ANALYSIS

This is the systematic examination of a chemical compound to find out the elements or ions present in it, type of bond or nature of reactivity. It is also referred to as property analysis.

In a practical experiment, a student is normally given one or more substances and he/she is expected to carryout a number of tests on the substance(s). He/she is further expected to record the observations and make relevant deductions about the nature of substance(s) given.

Preliminary test

The purpose of this test is to give general guidelines on the nature of substance being analyzed. Preliminary analysis is based on three major aspects namely:

- Appearance
- Flame colour
- Action of heat

Appearance

The appearance of substances can further be divided in to three categories i.e. physical state, colour and smell.

i) Physical state

Carbonates and oxides appear powdery whereas nitrates, sulphates, and chlorides appear crystalline.

ii) Smell

Some compounds especially ammonium compounds and sulphides posses a noticeable smell.

iii) Colour

Colour serves as a good guide as to what metallic ions may be present in a given substance. However it is important to note that a substance cannot be completely identified basing on its colour.

Summary of analysis based on appearance

Observation	**Deduction**
Crystalline substance	Probably a nitrate, sulphate, chloride or a hydrate (hydrated compound) present
Deliquescent substance	Probably a chloride, nitrate present
White solid or colourless solution	Probably potassium(K^+), sodium(Na^+), calcium(Ca^{2+}), magnesium(Mg^{2+}), aluminium(Al^{3+}), zinc(Zn^{2+}), lead(Pb^{2+}), ammonium ions(NH_4^+) present
Green substance	Possibly iron(II) (Fe^{2+}), copper(II) (Cu^{2+}) present
Blue substance	Probably copper(II) (Cu^{2+}) present
Brown/yellow substance	Probably iron(III) (Fe^{3+}), lead(II) oxide (PbO)(if the substance is a solid) present
Colourless gas with a pungent smell (choking smell) and turns moist red litmus paper blue	Possibly ammonium salt, (NH_4^+) present as ammonia gas is given off.
Colourless gas with a pungent smell (choking smell) and turns moist blue litmus paper red	Possibly sulphite (SO_3^{2-}) present
Smell of rotten eggs	Probably a sulphide (S^{2-}) present

Action of heat

Usually, a little of the substance is heated in a dry, hard glass test tube gently at first and then strongly until there is no further change.

Since the action of heat on solids leads to decomposition of the substance, there is need to always observe the gases given off (if any) and the residue left in the test tube after heating. Observations made when some solids are heated are given in the table below.

Observations	**Deductions**
Residue is yellow when hot and white when cold	Residue is zinc oxide (ZnO), Zinc salt probably present
Residue is brown/red when hot and turns yellow on cooling	Residue is lead(II) oxide (PbO), Lead(II) salt probably present
Residue is a black solid	Residue is copper(II) oxide (CuO) or iron(II) oxide (FeO)/triiron tetraoxide (Fe_3O_4). copper(II) (Cu^{2+}),iron(II) (Fe^{2+}) or Iron(III) (Fe^{3+}) salts probably present
White sublimate	Probably ammonium salt present
Colourless vapour condenses into a colourless liquid on cooler part of the test tube and turns white anhydrous copper(II) sulphate to blue	Substance is hydrated i.e. contains water of crystallization
Cracking sound heard (decripitation occurs)	Possibly lead(II) nitrate ($Pb(NO_3)_2$) present
Brown fumes which turns damp blue litmus paper red	Nitrogen dioxide gas given, nitrate (NO_3^-) possibly present
Colourless gas that turns lime water milky and damp blue litmus paper pink is given off	Carbon dioxide gas given off, A carbonate (CO_3^{2-}) or A hydrogen carbonate (HCO_3^-) probably present
A colourless gas that relights a glowing splint is given off	Oxygen is given off, probably a nitrate (NO_3^-) or an oxide (O^{2-}) is present
Colourless gas with a pungent smell (choking smell) and turns moist red litmus paper blue	Possibly ammonium salt, (NH_4^+) present as ammonia gas is given off.

Flame test

This is done by examining the colour of flame of the substance burning.

Flame colour	**Deduction**
Bright yellow flame	Sodium ion (Na^{+}) possibly present
Lilac (purple)	Potassium ion (K^{+}) possibly present
Red	Calcium ion (Ca^{2+}) possibly present
Blue-green	Copper(II) ion (Cu^{2+}) possibly present
Blue	Lead(II) ion (Pb^{2+}) possibly present

Solubility of salts

Since qualitative analysis is mainly carried out in aqueous solutions, there is need to know which salts are soluble in water and those that are in soluble.

Soluble	**Insoluble**
All salts of sodium, potassium and ammonium	
All nitrates	
All sulphates	Except lead(II) sulphate ($PbSO_4$), barium sulphate ($BaSO_4$) and calcium sulphate ($CaSO_4$)
All chlorides	Except silver chloride (AgCl), lead(II) chloride ($PbCl_2$) (only soluble in hot water)
Carbonates of potassium, sodium and ammonium	All carbonates except potassium, sodium and ammonium carbonates
Hydroxides of potassium, sodium and ammonium	All hydroxides except potassium, sodium and ammonium

Making a test solution

Procedure

Dissolve a little of the substance provided e.g. a spatula end full in about 5cm^3 of water. Shake the test tube well until the substance dissolves. The solution obtained is now your test solution.

If the solid is insoluble in water, you will be required to dissolve it in cold dilute nitric acid or dilute hydrochloric acid.

Detection of cations

The following cations for this level will be considered: NH_4^+, Al^{3+}, Cu^{2+}, Zn^{2+}, Mg^{2+}, Pb^{2+}, Fe^{2+}, Fe^{3+}, Ca^{2+}. These cations can be identified using

i) Sodium hydroxide solution, NaOH(aq)
ii) Ammonia solution/aqueous ammonia, NH_4OH or NH_3(aq)

The main reaction between these two reagents and the cations is the precipitation of insoluble metal hydroxides. Some of these metal hydroxides dissolve in excess of these reagents forming solutions of complex ions and some do not dissolve.

Identification using sodium hydroxide

Procedure

- To a small portion of the test solution, add a few drops of sodium hydroxide and shake the mixture. Check if precipitate is formed and note its colour.
- Add more sodium hydroxide solution until in excess and then shake the solution. Note whether the precipitate formed is soluble in excess alkali or insoluble in excess alkali.

NB. Incase no precipitate appears after adding sodium hydroxide solution, warm the mixture gently and test for ammonia.

Below is a summary of reactions that takes place for common cations with NaOH

Observation	Cation	Ionic equation for reaction
Paleblue precipitate; insoluble in excess alkali	Cu^{2+}	$Cu^{2+}(aq) + 2OH^-(aq) \longrightarrow Cu(OH)_2(s)$

Green precipitate; in soluble in excess alkali and turns brown on standing.	Fe^{2+}	Fe^{2+}(aq) + 2OH^-(aq) $\longrightarrow$ Fe(OH)$_2$(s)
Brown precipitate formed which is insoluble in excess alkali	Fe^{3+}	Fe^{3+}(aq) + 3OH^-(aq) Fe(OH)$_3$(s)
White precipitate soluble in excess alkali forming a colourless solution	Pb^{2+}, Zn^{2+}, or Al^{3+}	Al^{3+}(aq) + 3OH^-(aq) $\longrightarrow$ Al(OH)$_3$(s) Pb^{2+}(aq) + 2OH^-(aq) $\longrightarrow$ Pb(OH)$_2$(s) Zn^{2+}(aq) + 2OH^-(aq) $\longrightarrow$ Zn(OH)$_2$(s) **In excess** Al(OH)$_3$(s) + 3OH^-(aq) $\longrightarrow$ $Al(OH)_4^{3-}$(aq) (Aluminate ion) Pb(OH)$_2$(s) + 2OH^-(aq) $\longrightarrow$ $Pb(OH)_4^{2-}$(aq) (Plumbate ion) Zn(OH)$_2$(s) + 2OH^-(aq) $\longrightarrow$ $Zn(OH)_4^{2-}$(aq) (Zincate ion)
White precipitate insoluble in excess alkali	Ca^{2+} Mg^{2+}	Ca^{2+}(aq) + 2OH^-(aq) $\longrightarrow$ Ca(OH)$_2$(s) Mg^{2+}(aq) + 2OH^-(aq) $\longrightarrow$ Mg(OH)$_2$(s)
No observable change. Colourless gas with a pungent choking smell which turns damp red litmus paper blue on warming is given off	NH_4^+	NH_4^+(aq) + OH^-(aq) $\longrightarrow$ NH$_3$(g) + H$_2$O(l)

NB.Iron(II) (Fe^{2+}) turns to brown on standing due to air oxidation/aerial oxidation of iron(II) to iron(III) (Fe^{3+}).

Identification of cations using aqueous ammonia

The procedure remains the same as that of sodium hydroxide. Here, there is no heating even when precipitates are not formed.

Below is a summary of reactions that takes place for common cations

Observation	**Cation**	**Ionic equation for reaction**
Pale blue precipitate; soluble in excess alkali to form a deep blue solution	Cu^{2+}	$Cu^{2+}(aq) + 2OH^-(aq) \longrightarrow Cu(OH)_2(s)$ **In excess** $Cu(OH)_2(s) + 4NH_3(aq) \longrightarrow Cu(NH_3)_4^{2+}(aq) + 2OH^-(aq)$ (tetra amine copper(II) ion)
Green precipitate; in soluble in excess alkali and turns brown on standing.	Fe^{2+}	$Fe^{2+}(aq) + 2OH^-(aq) \longrightarrow Fe(OH)_2(s)$
Brown precipitate formed which is insoluble in excess alkali	Fe^{3+}	$Fe^{3+}(aq) + 3OH^-(aq) \longrightarrow Fe(OH)_3(s)$
White precipitate soluble in excess alkali forming a colourless solution	Zn^{2+},	$Zn^{2+}(aq) + 2OH^-(aq) \longrightarrow Zn(OH)_2(s)$ **In excess** $Zn(OH)_2(s) + 4NH_3(aq) \longrightarrow Zn(NH_3)_4^{2+}(aq) + 2OH^-(aq)$ (tetra amine zinc(II) ion)
White precipitate insoluble in excess alkali	Al^{3+}, Pb^{2+}	$Al^{3+}(aq) + 3OH^-(aq) \longrightarrow Al(OH)_3(s)$ $Pb^{2+}(aq) + 2OH^-(aq) \longrightarrow Pb(OH)_2(s)$
No observable change	NH_4^+, Ca^{2+}	

Confirmatory tests for cations

A confirmatory test is a specific sensitive reaction undergone by a particular ion. Since sodium hydroxide solution or ammonia cannot completely identify some cations, confirmatory tests are then used.

Confirmatory tests for lead(II) ions, Pb^{2+}

i) To a small portion of solution containing Pb^{2+} in a test tube, add 3 drops of potassium iodide solution.

Observation

Bright yellow precipitate is formed

Equation

Pb^{2+}(aq) + 2 I^-(aq) $\longrightarrow$ PbI_2(s)

(Bright yellow precipitate)

ii) To a small portion of solution containing Pb^{2+} in a test tube, add 3 drops of dilute sulphuric acid.

Observation

White precipitate is formed

Equation

Pb^{2+}(aq) + SO_4^{2-}(aq) $\longrightarrow$ $PbSO_4$(s)

(White precipitate)

iii) To a small portion of solution containing Pb^{2+} in a test tube, add 3 drops of dilute hydrochloric acid.

Observation

White precipitates are formed which dissolve upon boiling and recrystallises on cooling.

Equation

Pb^{2+}(aq) + 2Cl^-(aq) $\longrightarrow$ $PbCI_2$(s)

(White precipitate)

Confirmatory tests for copper(II) ions, Cu^{2+}

To a small portion of solution containing Cu^{2+} add ammonia solution drop wise until in excess.

Observation

Pale blue precipitates are formed which dissolve in excess ammonia solution forming a deep blue solution.

Equation

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

In excess

$Cu(OH)_2$(s)+4NH_3(aq) $\longrightarrow$ $Cu(NH_3)_4^{2+}$(aq)+ 2OH^-(aq)

(Tetraamine copper(II) ion)

Confirmatory tests for zinc ions, Zn^{2+}

To a small portion of solution containing Zn^{2+} add ammonia solution drop wise until in excess.

Observation

White precipitates are formed which dissolve in excess ammonia solution forming a colourless solution.

Equation

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

In excess

$Zn(OH)_2(s) + 4NH_3(aq) \longrightarrow Zn(NH_3)_4^{2+}(aq) + 2OH^-(aq)$

(Tetra amine zinc(II) ion)

Detection of anions

Major anions that are considered are: sulphate (SO_4^{2-}), carbonate (CO_3^{2-}), chloride (Cl^-) and nitrate (NO_3^-).

These ions can be divided into three major groups as:

- Those that react with (major) dilute acids to give off gases
- Those that react with concentrated sulphuric acid to liberate gases
- Those that react with neither dilute acids nor concentrated sulphuric acid

Anions that react with dilute acids

Test procedure

To the test solution, add 2-3 drops of dilute nitric acid or dilute hydrochloric acid. If no reaction occurs, warm the mixture gently, if effervescence occurs, smell the gas with care and identify it.

Observation	**Deduction**	**Possible reaction**
A colourless gas that turns lime water milky and damp blue litmus paper pink is evolved.	Carbondioxide given off, possibly CO_3^{2-} or HCO_3^- present	$CO_3^{2-}(aq) + 2H^+(aq) \longrightarrow CO_2(g) + H_2O(l)$
Colourless gas with an Irritating/pungent smell and turns damp blue litmus paper to red	Sulphur dioxide gas given off, possibly SO_3^{2-} is present	$SO_3^{2-}(aq) + 2H^+(aq) \longrightarrow SO_2(g) + H_2O(l)$

A colourless gas with the smell of rotten eggs	Hydrogen sulphide gas is given off, probably S^{2-} present	$S^{2-}(aq) + 2H^{+}(aq) \longrightarrow H_2S(g)$

Anions that react with concentrated sulphuric acid

The reaction is the same as above but in this case, they are more vigorous.

Test procedure

Add 1-2 cm^3 of concentrated sulphuric to the solid under test in a test tube. If no reaction occurs, warm the mixture but do not boil. Identify the anion present as per the following table.

Observation/gas evolved	**Possible anion**
A colourless gas that turns lime water milky and damp blue litmus paper pink is evolved.	Carbon dioxide given off, possibly CO_3^{2-} or HCO_3^- present
Colourless gas with an Irritating/pungent smell and turns damp blue litmus paper to red	Sulphur dioxide gas given off, possibly SO_3^{2-} is present
Greenish yellow gas with an irritating smell which turns wet blue litmus red then bleaches it and forms dense white fumes with concentrated ammonia	Chlorine gas given off, probably Cl^- is present
Brown fumes of a gas which turns wet blue litmus paper red is given off	Nitrogen dioxide gas evolved, NO_3^- probably present
A colourless gas with the smell of rotten eggs	Hydrogen sulphide gas is given off, probably S^{2-} present

Anions that do not react with acids

Under this, consideration will only be made for a sulphate.

Test for a sulphate (SO_4^{2-})

Procedure

To the solution under test, add an equal volume of dilute nitric acid followed by 2-3 drops of:

i) Barium nitrate solution

Observation

A dense white precipitate is formed (if a sulphate is present)

Equation

$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$

This serves as a confirmatory test for a sulphate.

NB. Barium chloride solution can be used in place of barium nitrate

ii) Lead(II) nitrate solution and boil the mixture

Observation

White precipitate is formed which does not dissolve on boiling

Equation

$Pb^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow PbSO_4(s)$

Confirmatory test for other anions

Chloride (Cl^-)

a) Silver nitrate test

To a small portion of solution containing Cl^-, add an equal volume of dilute nitric acid followed by 2-3 drops of silver nitrate solution.

Observation

A white precipitate appears which turns violet on exposure to light and the precipitate is readily soluble in aqueous ammonia.

Equation

$Ag^+(aq) + Cl^-(aq) \longrightarrow \underset{\text{(White precipitate)}}{AgCl(s)}$

b) Lead(II) nitrate test

Procedure

To a small portion of solution containing Cl^-, add an equal volume of dilute nitric acid followed by 2-3 drops of lead(II) nitrate solution and boil the mixture.

Observation

A white precipitate formed which dissolves on boiling and recrystallises on cooling. This confirms a chloride.

Equation

$Pb^{2+}(aq) + 2Cl^-(aq) \longrightarrow PbCl_2(s)$

(White precipitates which dissolves on boiling)

Carbonate (CO_3^{2-})

All carbonates and hydrogen carbonates give off carbon dioxide when they are treated with dilute acids. This serves as a temporary test for carbonates and hydrogen carbonates.

Apart from potassium, sodium and ammonium carbonates, all other carbonates decompose upon heating to give off carbon dioxide.

Test of a soluble carbonate

Procedure

To the test solution, add

i) Barium nitrate solution/barium chloride solution followed by excess dilute nitric acid.

Observations

A white precipitate is formed which dissolves in excess dilute nitric acid with effervescence.

Equation

$Ba^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow BaCO_3(s)$

ii) Add lead(II) nitrate solution followed by excess dilute nitric acid

Observation

A white precipitate is formed. The precipitate dissolves with effervescence in excess dilute nitric acid.

Equation

$Pb^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow PbCO_3(s)$

iii) Add 2-3 drops of silver nitrate solution followed by excess dilute nitric acid

Observation

A white precipitate appears that dissolves in excess dilute nitric acid

Equation

$2Ag^{+}(aq) + CO_3^{2-}(aq) \longrightarrow Ag_2CO_3(s)$

Nitrate (NO_3^-)

A part from potassium, sodium and ammonium nitrates, all other nitrates liberate brown fumes of nitrogen dioxide gas when heated strongly. The liberation of brown fumes serves as a temporary confirmatory test for a nitrate.

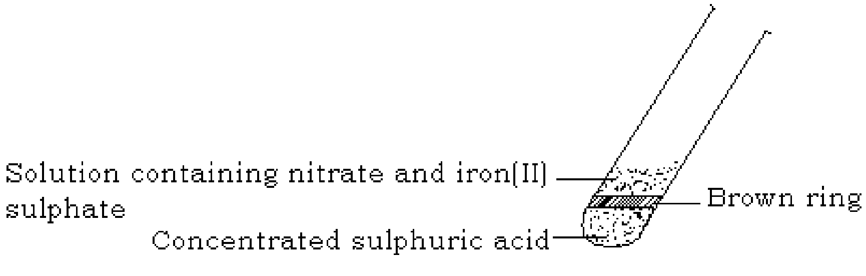
Brown ring test

Procedure

To the solution of a substance under test, add an equal volume of freshly prepared iron(II) sulphate solution. Then carefully add a few drops of concentrated sulphuric acid down the sides of a slanting test tube so that the acid sinks to the bottom and forms a separate layer.

Observation

A brown ring slowly forms where the acid meets the top layer. This confirms the presence of a nitrate.



Detection and identification of gases

Gases are normally given off when substances are heated or when reagents like acids are added to substances.

How do you know that a gas is being evolved?

i) Effervescence (rapid emission of bubbles) occurs

ii) Colour. A few gases posses noticeable colours e.g. brown fumes for nitrogen dioxide.

iii) Smell. Some gases have got distinctive smells and they are usually unpleasant. E.g. ammonia (choking smell) and hydrogen dioxide (smell of rotten eggs)

After identifying that a gas is given off, it can then be identified using a chemical test.

Common tests to identify gases

i) Support of combustion

a) Oxygen relights a glowing splint

b) For a burning splint, if the flame is blue in colour, then suspect hydrogen or carbon monoxide.

ii) Litmus paper

Damp/wet litmus paper is always used. Alkaline gases such as ammonia turn red litmus paper blue while acidic gases like carbon dioxide turn blue litmus pink/red.

iii) Lime water/calcium hydroxide

This serves as a confirmatory test for carbon dioxide gas, indicating the presence of a carbonate or hydrogen carbonate. Carbon dioxide turns lime water milky.

iv) Water vapour

This is seen as a colourless condensed liquid on the wall of the test tube when heating.

Summary of gases

Gas	Colour	Smell	Effect on damp litmus	Specific test
Water vapour	Colourless	No smell	No effect	Turns white anhydrous copper(II) sulphate to blue
Carbon dioxide	Colourless	No smell	Turns blue litmus paper to	Forms white precipitates with lime water

			pink	
Hydrogen sulphide	Colourless	Foul smell (rotten egg smell)	Turns blue litmus paper to pink	It turns potassium dichromate(VI) solution from golden yellow to green
Ammonia	Colourless	Choking smell	Turns red litmus paper to blue	Gives dense white fumes when in contact with hydrogen chloride
Nitrogen dioxide	Brown fumes	Irritating smell	Turns blue litmus paper to red	No test
Oxygen	Colourless	No smell	No effect	Relights a glowing splint
Chlorine	Green-yellow gas	Irritating smell	Turns blue litmus to red then bleaches it	No test
Sulphur dioxide	Colourless	Irritating smell	Turns blue litmus paper red then bleaches it	It turns potassium dichromate(VI) solution from golden yellow to green

Sample questions with solutions

1. You are provided with substance **X** that contains one cation and one anion. Carry out the following tests to identify the cation and anion in **X**. Identify any gases that may be evolved. Record your observations and deduction in the table below.

Test	**Observation**	**Deduction**
a) Appearance of **X**	X is a green crystalline substance	Transition metal ion. Fe^{2+}, or Cu^{2+}
b) Place a spatula endful of **X** in a dry test-tube then heat strongly till no further change occurs	X forms a black residue A colourless vapour that condenses to a colourless liquid is given off. The liquid turns white anhydrous copper (II)sulphate blue. A colourless gas with an irritating smell is evolved. The gas turns blue litmus red and forms misty fumes in air and dense white fumes with conc. Ammonia solution on a glass rod	CuO or FeO/Fe₃O₄ Cu^{2+} or Fe^{2+} or Fe^{3+}present Water of crystallization From hydrated salt or OH^- Hydrogen chloride gas evolved Cl^- probably present
c) To a spatula endful of **X** in a test tube add 5cm3 of water and shake. Divide the solution into 4 portions. i) To the first portion, add sodium hydroxide solution drop wise till in excess	X dissolves in water forming a pale blue solution Pale blue precipitate insoluble in excess NaOH(aq)	Cu^{2+} suspected Cu^{2+}present
ii) To second portion, add dilute ammonia solution dropwise till in excess	Blue ppt which dissolves in excess NaOH(aq) forming a deep blue solution	Cu^{2+} confirmed

iii) To the third portion, add a few drops of Lead (II) nitrate solution	White ppt observed	SO_4^{2-} or Cl^- present
iv) Use the fourth portion to carry a test of your own choice to confirm the anion in **X** **Test** To the fourth portion, add 2-3 drops of silver nitrate solution followed by dilute nitric acid	Intense white ppt insoluble in nitric acid	Cl^- confirmed

Identify the:

Cation in X: Cu^{2+}

Anion in X: Cl^-

2. You are provided with substance **Y** that contains two cations and two anions. Carry out the following tests to identify any gases that may be evolved. Record your observations and deductions in the table below

Test	**Observation**	**Deduction**
a) Place two spatula endful of **Y** in a test-tube, add 5cm^3 of water, shake vigorously and filter. Keep both filtrate and residue. Divide the filtrate into 3 parts. i) To the first part add excess sodium hydroxide solution then heat and smell	White crystalline –powdery substance Y partially dissolves forming a colourless filtrate and white residue. No observable colour change. But a colourless gas with an irritating pungent smell is evolved. The gas turns damp red litmus blue and forms dense white fumes with Conc. HCl (aq) on a glass rod	Non-transition metals such as Pb^{2+}, Zn^{2+}, NH_4^+, Al^{3+} possibly present NH_3 gas evolved NH_4^+ confirmed

ii) To the second part, add Lead (II) nitrate solution and heat	White ppt insoluble on heating	Cl^- absent SO_4^{2-} probably present
iii) Use the third part to carry out a test of your choice to confirm the anion in Y **Test:** To third part add 2-3 drops of Barium chloride solution followed by dilute hydrochloric acid	White ppt insoluble in the acid	SO_4^{2-} confirmed
b) Wash the residue with little distilled water then divide it into two parts. i) Strongly heat the first part until no further change occurs	White residue turns yellow when hot and then white on cooling. A colourless gas that turns blue litmus pink and lime water milky is evolved	CO_2 gas; CO_3^{2-} present
c) To the second part of the residue, add 3cm3 of dilute hydrochloric acid. Divide the mixture into 2 parts	The residue dissolves with rapid effervescence of a colourless gas that turns damp blue litmus pink and lime water milky. A colourless solution is formed	CO_2 gas; CO_3^{2-} confirmed
i) To the first part of the mixture, add sodium hydroxide solution drop wise till in excess	White ppt soluble in excess NaOH(aq) forming a colourless solution is observed	Zn^{2+}, Pb^{2+}, Al^{3+} present
ii) To the second part add ammonia solution drop wise till in excess	White ppt that dissolves in excess NH_3 (aq) forming a colourless solution is observed.	Zn^{2+} confirmed

Identify the:

Cations in Y : NH_4^+ and Zn^{2+}

Anions in Y: SO_4^{2-} and CO_3^{2-}

NITROGEN AND ITS COMPOUNDS

Nitrogen is in period 2 and group (V) of the periodic table of elements. It has atomic number of 7 and electronic configuration 2:5.

Occurrence

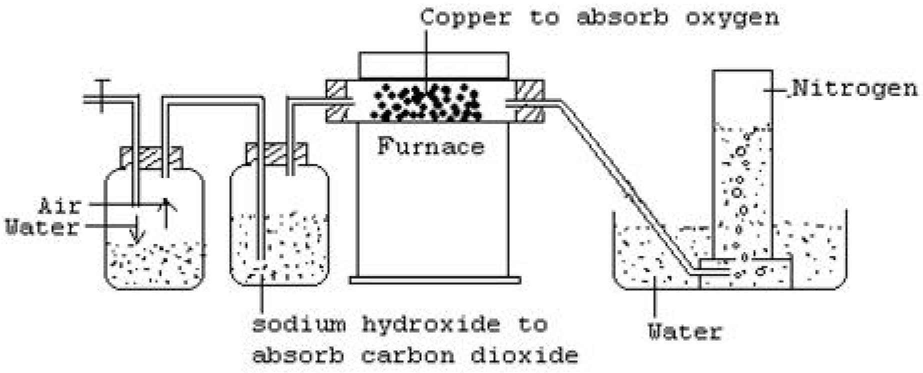
Nitrogen exists freely in the air as diatomic molecules and makes up to 78% of the air by volume.

It occurs in combined states as nitrites, nitrates and most of these are distributed every where in soil as ammonium sulphate, sodium nitrate, potassium nitrate and calcium nitrate,. It is also constituent of living matter of pants and animals.

Laboratory preparation of nitrogen

a) From air

Set up



Air is made to pass through a solution of concentrated sodium hydroxide to remove carbon dioxide gas

CO_2 (g) + 2NaOH (aq) $\longrightarrow$ Na_2CO_3 (aq) + H_2O (l)

The remaining air is then passed over heated copper to remove oxygen gas

2Cu(s) + O_2 (g) $\longrightarrow$ 2CuO(s)

The nitrogen gas can now be collected over water.

If the nitrogen is required dry, it is then passed over fused calcium chloride to remove water vapour or it can be passed through a U tube containing beads soaked in concentrated sulphuric acid to dry the nitrogen gas.

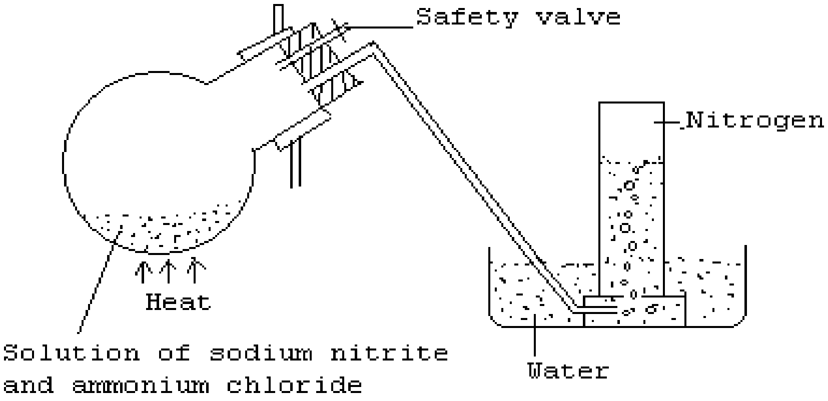
NB. The nitrogen obtained by this method is denser than ordinary nitrogen since it contains impurities like argon and other inert gases.

b) Preparation of nitrogen by action of heat on ammonium nitrite

The ammonium nitrite is formed by the reaction between sodium nitrite and ammonium chloride.

$NaNO_2(aq) + NH_4Cl(aq) \longrightarrow NaCl(aq) + NH_4NO_2(aq)$

On slight warming, the solution of ammonium nitrite decomposes to give nitrogen gas. The nitrogen gas produced can then be collected over water as shown.



$NH_4NO_2(aq) \longrightarrow N_2(g) + 2H_2O(l)$

c) Other methods

i) Action of chlorine on ammonia

$3Cl_2(g) + 8NH_3(g) \longrightarrow N_2(g) + 6NH_4Cl(s)$

ii) Passing ammonia over heated copper(II) oxide

$3CuO(s) + 12NH_3(g) \longrightarrow 3Cu(s) + 3H_2O(l) + N_2(g)$

iii) Reduction of oxides of nitrogen using copper

$$2NO_2(g) + 4Cu(s) \longrightarrow N_2(g) + 4CuO(s)$$

Industrial preparation of nitrogen

Nitrogen is obtained in the industry by fractional distillation of liquid air. Liquid air is fractionally distilled and nitrogen is obtained at a temperature of -196˚C (77 K at standard pressure). Oxygen with a higher boiling point (-183˚C) is left behind. The separated nitrogen is liquefied and stored in specially designed container ready for use. The nitrogen may also be sold as compressed air.

Properties of nitrogen

Physical properties

- It is a tasteless, colorless and odourless gas
- It is slightly soluble in water (almost insoluble)
- It is slightly denser than air
- It is a neutral gas i.e. neither acidic nor basic
- It does not support burning though a few metals burn in it.

Chemical properties

Nitrogen is generally an unreactive gas. This is because of the presence of the strong triple covalent bonds between its atoms in a molecule (N≡N). The triple covalent bonds are hard to break rendering nitrogen inert.

Some reactive metals like magnesium and calcium burn in nitrogen to form nitrides i.e. magnesium nitride and calcium nitride

$$3Mg(s) + N_2(g) \longrightarrow \underset{\text{(Green yellow)}}{Mg_3N_2(s)}$$

$$3Ca(s) + N_2(g) \longrightarrow \underset{\text{(Dark red)}}{Ca_3N_2(s)}$$

The nitrides formed dissolves in water to form an alkaline solution of calcium and magnesium hydroxides; a gas with a choking smell (ammonia) is also given off.

$Ca_3N_2(s) + 6H_2O(l) \longrightarrow 3Ca(OH)_2(aq) + 2NH_3(g)$

Nitrogen reacts with hydrogen under special conditions to form ammonia (Haber process).

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

Nitrogen is inert and does not burn though is forms a number of oxides like nitrogen monoxide (NO), nitrogen dioxide (NO_2) and dinitrogen oxide (N_2O).

Uses of nitrogen

- Nitrogen is used in the synthesis of ammonia gas.
- It is used in the manufacture of fertilizers like ammonium phosphate.
- The atmospheric nitrogen is fixed by thunder or bacteria in root nodules of leguminous plants to nitrates which can be used by plants.
- Because of its low boiling points, liquid nitrogen is used to cool materials to very low temperatures.
- It is used in the manufacture of nitric acid.

OXIDES OF NITROGEN

NITROGEN DIOXIDE (NO_2)

Laboratory preparation

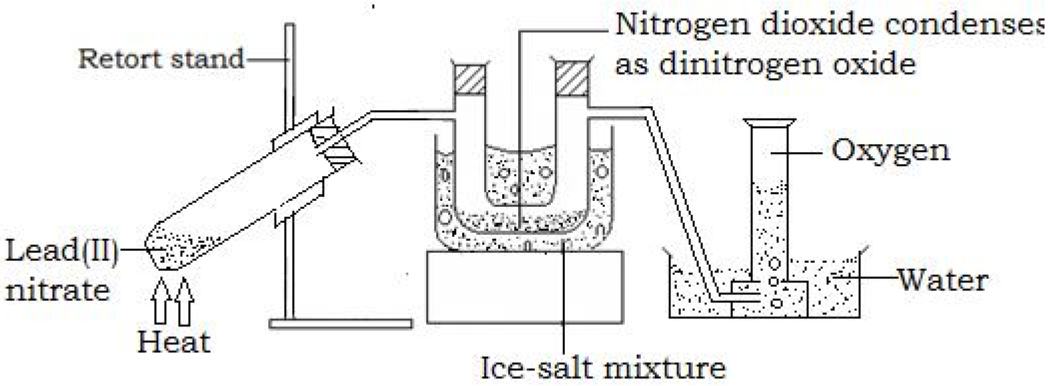
It is conveniently prepared in the laboratory by heating lead(II) nitrate crystals (this is because lead(II) nitrate forms crystals without water of crystallization which is not common to other metallic nitrates and would otherwise interfere with the preparation)

Nitrogen dioxide is finally obtained as brown fumes

$2Pb(NO_3)_2(s) \longrightarrow 2PbO(s) + 4NO_2(g) + 2O_2(g)$

NB.

1. The gas is passed through a freezing mixture of common salt and ice where it collects in a U tube as dinitrogen tetra oxide, N_2O_4(a pale green liquid which appears pale yellow when it is pure). The oxygen gas passes on and escapes as a colourless, harm less gas which can be collected if required over water as shown.
2. As the white lead(II) nitrate crystals are heated, they make a crackling sound and melt. This is known as decrepitation. The residue in the boiling tube is Lead(II) oxide which is yellow when it cools.



Properties of nitrogen dioxide

Physical properties

- It is reddish brown gas at room temperature
- It is soluble in water
- It is denser than air
- It is highly poisonous and produces nitric acid in the lungs when inhaled (it causes septic pneumonia)
- It has an irritating pungent smell
- It is easily liquefied

Chemical properties

1. Burning of metals. Metals e.g. magnesium and non metals like phosphorus burn even more highly in nitrogen dioxide. This suggests that nitrogen dioxide supports combustion and this will be so if the material reacting is hot enough to decompose nitrogen dioxide into oxygen (a gas that supports burning)

$$4Mg(g) + 2NO_2(g) \longrightarrow 4MgO(s) + N_2(g)$$

$$2P_4(s) + 10NO_2(g) \longrightarrow 2P_4O_{10}(s) + 5N_2(g)$$

2. Nitrogen dioxide oxidizes red hot metals e.g. copper and zinc to their oxides and nitrogen dioxide is itself reduced to nitrogen.

$$4Cu(s) + 2NO_2(g) \longrightarrow 4CuO(s) + N_2(g)$$

(Brown) (Black)

3. When the gas is heated above 150˚C, the brown colour of the gas tends to fade gradually giving a colourless gas which is a mixture of nitrogen monoxide and oxygen. This is called dissociation of nitrogen dioxide.

$$2NO_2(g) \rightleftharpoons \underbrace{2NO(g) + O_2(g)}$$

(Brown) (Colourless)

$$N_2O_4 \rightleftharpoons 2NO_2(g) \rightleftharpoons \underbrace{2NO(g) + O_2(g)}$$

(Green-impure)

(Pale yellow-pure) (Brown fumes) (Colourless)

4. Nitrogen dioxide dissolves in water to give a faint blue solution. This solution is a mixture of nitric acid and nitrous acid.

$$2NO_2(g) + H_2O(l) \longrightarrow HNO_3(aq) + HNO_2(aq)$$

The nitrous acid formed is immediately oxidized by atmospheric oxygen to form nitric acid.

$$2HNO_2(aq) + O_2(g) \longrightarrow 2HNO_3(aq)$$

NB. Since nitrogen on dissolving in water produces two acids, it is said to be a **mixed acid anhydride**.

3. Nitrogen dioxide neutralizes alkalis to give the corresponding nitrate and a nitrite together with water.

   $2NO_2(g) + 2NaOH(aq) \longrightarrow NaNO_3(aq) + NaNO_2(aq) + H_2O(l)$

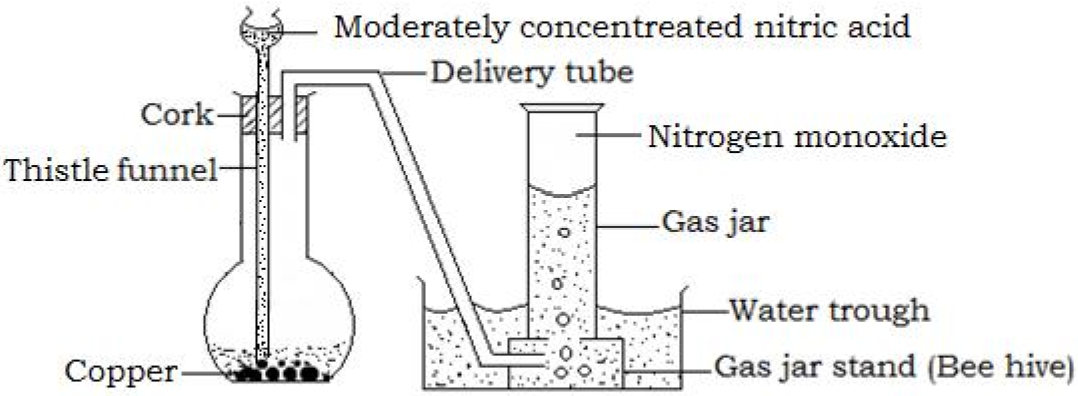
Uses of nitrogen dioxide

- It is used in the manufacture of plastics, explosives, nylon materials e.t.c.
- Used in the manufacture of nitric acid.

NITROGEN MONOXIDE (NO)

Laboratory preparation

Set up



Procedure

Place some copper turnings in a flask, add some water to cover it, then add moderately concentrated (50%) nitric acid (same volume as that of the water).

Observation

Vigorous effervescence occurs and the flask is filled with brown fumes. The brown fumes are nitrogen dioxide produced partly by the action of the acid upon the copper and partly by the oxidation of the main product, nitrogen monoxide by the oxygen of the air in the flask.

$2NO(g) + O_2(g) \longrightarrow 2NO_2(s)$

The brown fumes dissolve in water over which the nitrogen monoxide is collected as a colourless gas. A green solution of copper (II) nitrate is left in the flask.

$3Cu(s) + 8HNO_3(aq) \longrightarrow 3Cu(NO_3)_2(aq) + 4H_2O(l) + 2NO(g)$

Properties of nitrogen monoxide

1. It does not normally support combustion and a burning splint is extinguished when placed in a gas jar of nitrogen monoxide. However, it supports combustion of very hot substances like magnesium and phosphorus which provide temperatures hot enough to decompose the gas into oxygen. E.g

   $2Mg(s) + 2NO(g) \longrightarrow 2MgO(s) + N_2(g)$

   $2P_4(s) + 10NO(g) \longrightarrow 2P_4O_5(s) + 5N_2(g)$

2. It is a neutral gas and has no effect on litmus paper.
3. **Chemical test**

   When the gas is exposed to air, nitrogen monoxide immediately reacts with atmospheric oxygen to form brown fumes of nitrogen dioxide.

   $2NO(g) + O_2(g) \rightleftharpoons 2NO_2(s)$

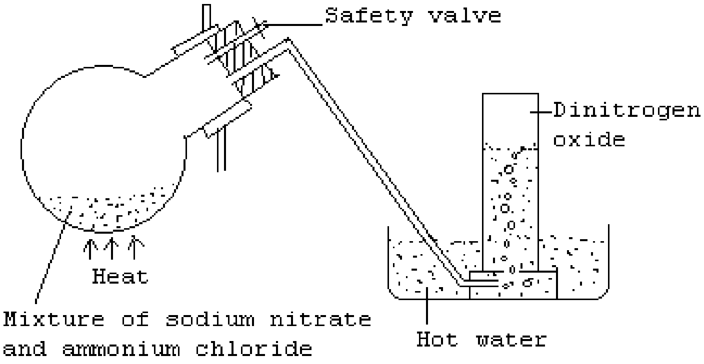
Alternative test

When the gas is bubble though iron(II) sulphate solution, the solution changes from pale green to dark brown or black. The dark brown of black coloration is due to formation of a compound, $(FeSO_4).NO$. When this compound is heated, pure nitrogen monoxide is formed.

DINITROGEN OXIDE, N_2O (LAUGHING GAS)

Laboratory preparation

Set up



Procedure

1. Grind a mixture of sodium nitrate and ammonium chloride.
2. Put the mixture in a round bottom flask and assemble the apparatus as shown in the diagram above.

NB. On heating the mixture, ammonium nitrate is formed which quickly decomposes to give dinitrogen oxide and water vapour.

$NaNO_3(s) + NH_4Cl(s) \longrightarrow NH_4NO_3(s) + NaCl(s)$

$NH_4NO_3(s) \longrightarrow N_2O(g) + 2H_2O(g)$

The dinitrogen oxide is then collected over hot water.

NB. It could also be prepared by heating ammonium nitrate directly but this method is dangerous since ammonium nitrate can be explosive when heated.

Physical properties of dinitrogen oxide

i) It is a colourless gas with a faint sweet smell.
ii) In it fairly soluble in cold water but insoluble in hot water.
iii) It is denser than air.
iv) It can be easily liquefied.
v) It is a neutral gas and has no effect on litmus paper.

Chemical properties of dinitrogen oxide

1. It supports combustion and relights a glowing splint that is hot enough to decompose it into oxygen and nitrogen.
2. Burning metals and non metals like magnesium and sulphur continue to burn in the gas giving their respective oxides.

$$N_2O(g) + Mg(s) \longrightarrow MgO(s) + N_2(g)$$
$$2N_2O(g) + S(s) \longrightarrow SO_2(g) + 2N_2(g)$$

Other metals like copper when they are very hot can be oxidized by the gas to their respective oxides.

$$N_2O(g) + Cu(s) \longrightarrow CuO(s) + N_2(g)$$

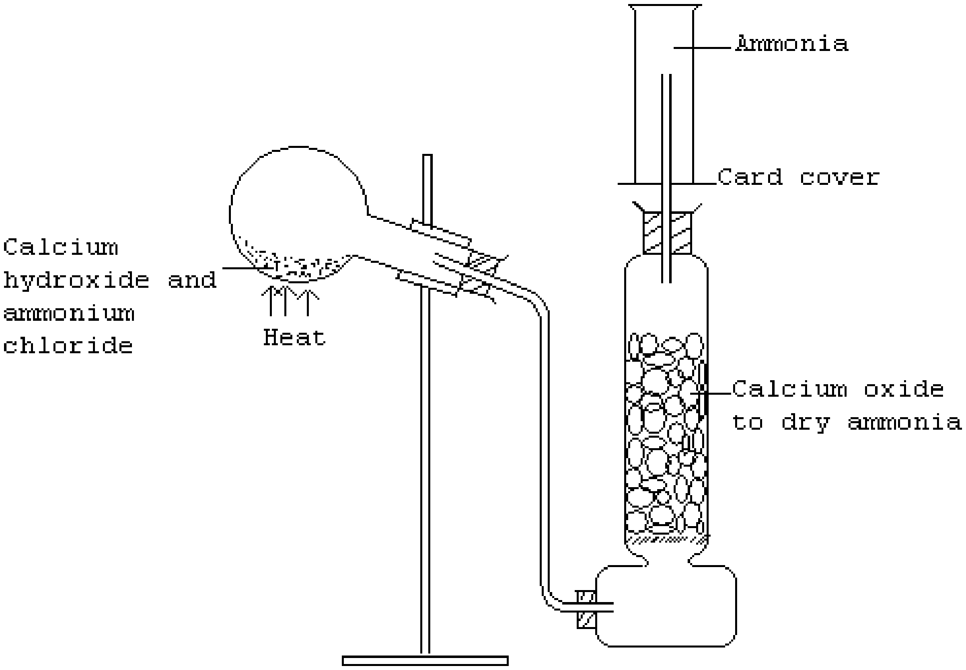
Use of dinitrogen oxide

It is used as an anaesthetic in less complex surgical operations like dentistry. It has an effect that makes a patient laugh hysterically and it is therefore referred to as "laughing gas".

AMMONIA

Laboratory preparation

Setup



Procedure

- Assemble the apparatus as shown in the diagram above
- Grind a mixture of ammonium chloride and calcium hydroxide and place it in a round bottom flask of a hard glass

NB. The neck of the flask should bend down wards and the flask should be in a slanting position because the formed water vapour will condense and if allowed to run back on the hot flask causes breakage

- Heat the mixture in the flask to evolve ammonia gas

Equation

$$Ca(OH)_2(s) + 2NH_4Cl(s) \longrightarrow 2NH_3(g) + CaCl_2(s) + 2H_2O(l)$$

- The gas is then dried by passing it over lumps of calcium oxide. The usual drying agents i.e. fused calcium chloride and concentrated calcium chloride react with the gas as below.

$$2NH_3(g) + H_2SO_4(l) \longrightarrow (NH_4)_2SO_4(s)$$

$$CaCl_2(s) + 4NH_3(g) \longrightarrow CaCl_2.4NH_3(s)$$

Since the gas is less dense than air and very soluble in water, it is collected by up ward delivery method.

Ammonia gas can be prepared in the laboratory by heating any ammonium salt with an alkali.

Properties of ammonia

Physical properties

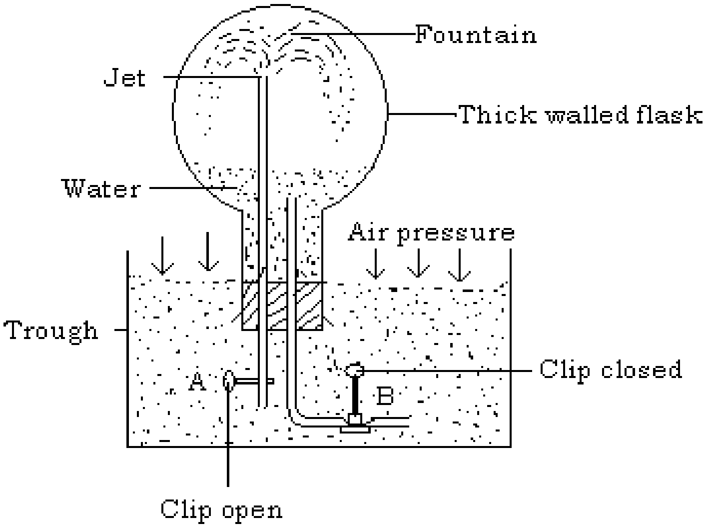
i) Ammonia is a colourless gas with a characteristic choking smell.
ii) It is less dense than air.
iii) It has a very low boiling point (-34°C) and liquefied under pressure.
iv) It is very soluble in water. It is in fact the most soluble gas known.
v) It turns moist red litmus paper blue. It is the only known alkaline gas.

Solubility of ammonia in water

The fountain experiment

The great solubility of ammonia is illustrated by the fountain experiment.

Set up



Procedure

- Fill a dry thick walled flask with ammonia gas and close it with a cork carrying tubes and clips as shown.
- Clamp the flask upside down and immerse the tubes with clips inside water which has been coloured with red litmus solution.
- Open clip B for a moment and close it. This allows few drops of water to enter in the flask. The water is made to run to the round end of the flask.
  The ammonia gas in the flask dissolves in the water (forming a blue solution since it is an alkaline gas). This greatly reduces the gas pressure inside the flask.
- Open clip A.

Observation

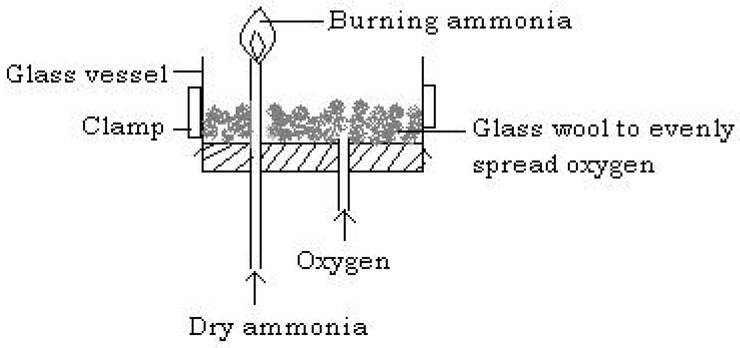
A fountain at once occurs as illustrated in the diagram. This will continue until the flask is full of water as it was with ammonia. When clip A was opened, water was forced into the flask because the atmospheric pressure from the outside was far much greater than the pressure inside the flask The fountain appears blue due to the alkaline nature of ammonia gas.

Chemical properties of ammonia

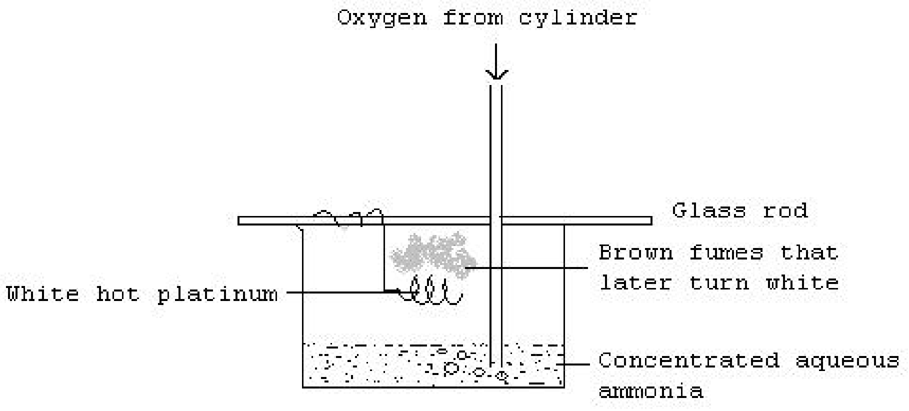
Reaction with air

When a lit splint is placed in a gas jar full of ammonia, it is extinguished showing that ammonia does not support combustion. On its own, ammonia does not burn but in plenty of oxygen, ammonia burns to produce nitrogen and water vapour.

$$4NH_3(g) + 3O_2(g) \longrightarrow 2N_2(g) + 6H_20(g)$$



In the presence of a catalyst e.g. platinum foil, ammonia it oxidized to nitrogen monoxide when it reacts with oxygen/air. In this case, the platinum foil continues to glow in the mixture of air and ammonia since the reaction is exothermic. Brown fumes of nitrogen dioxide are later seen as the nitrogen monoxide formed is oxidized in the presence of oxygen.



$4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(g)$

$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$

This reaction is known as **Ostwald's catalytic oxidation of ammonia**. It is the initial stage in the commercial preparation of nitric acid.

The fumes then later turn white due to the formation of ammonium nitrate according to the following reactions.

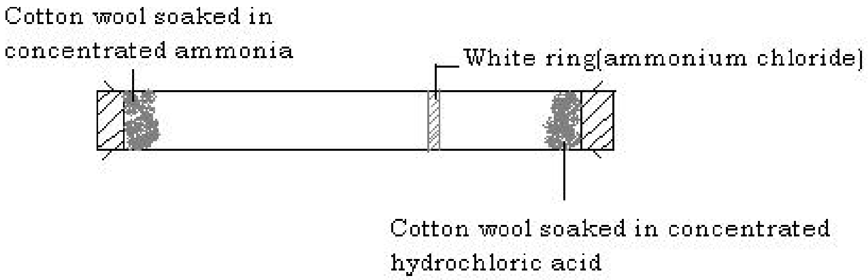
$2H_2O(g) + 4NO_2(g) + O_2(g) \longrightarrow 4HNO_3(g)$

$HNO_3(g) + NH_3(g) \longrightarrow NH_4NO_3(g)$

Reaction with hydrogen chloride gas

Ammonia reacts with hydrogen chloride to form dense white fumes of ammonium chloride which settle as white solids.

$NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s)$



Reaction with chlorine

When ammonia reacts with chlorine, ammonia reduces chlorine to hydrogen chloride and its self is oxidized to nitrogen.

$2NH_3(g) + 3Cl_2(g) \longrightarrow 6HCl(g) + N_2(g)$

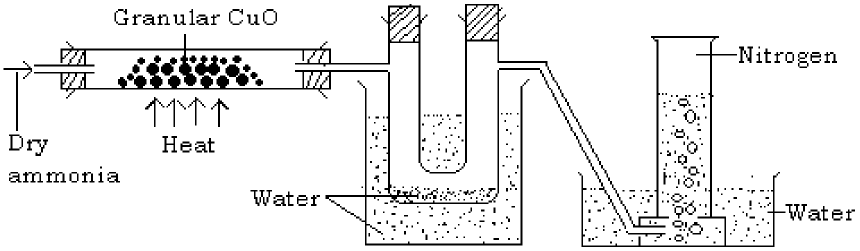
When ammonia is in excess, dense white fumes of ammonium chloride are formed.

$HCl(g) + NH_3(g) \rightleftharpoons NH_4Cl(s)$

Reaction with metal oxides

Ammonia reduces oxides of metals below iron in the reactivity series to respective metals and itself is oxidized to nitrogen and water. E.g when ammonia is passed over heated copper(II) oxide, the oxide changes from black to brown as copper metal is formed.

$3CuO(s) + 2NH_3(g) \longrightarrow 3Cu(s) + N_2(g) + 3H_2O(l)$



The colourless liquid collected in the U tube turns white anhydrous copper(II) sulphate blue indicating that it is water.

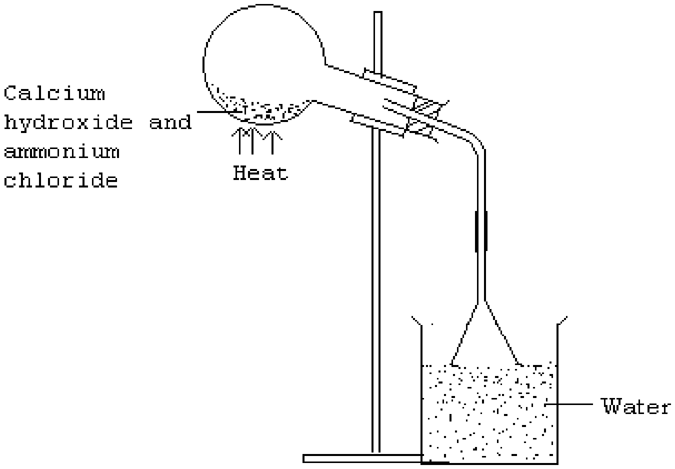
Reaction with water

Ammonia gas dissolves in water to form an alkaline solution of ammonium hydroxide.

$NH_3(g) + H_2O(l) \longrightarrow NH_4OH(aq)$

$NH_3(g) + (aq) \longrightarrow NH_3(aq)$

Setup for the preparation of ammonium solution



A mixture of ammonium chloride and calcium hydroxide is heated to produce ammonia gas which then dissolves in water to form the alkaline solution of ammonium hydroxide.

Precaution: The rim of the inverted funnel should just touch the surface of water. This prevents the water from being sucked into the flask.

Reaction with acids

Ammonia solution (aqueous ammonia) has many properties of typical alkaline solutions. E.g. it reacts with acids to form ammonium salts.

$NH_4OH(aq) + HCl(aq) \longrightarrow NH_4Cl(aq) + H_2O(l)$ $(NH_3(aq) + HCl(aq) \longrightarrow NH_4Cl(aq))$

$NH_4OH(aq)+H_2SO_4(aq) \longrightarrow (NH_4)_2SO_4(aq)+H_2O(l)$ $(NH_3(aq)+H_2SO_4(aq) \longrightarrow (NH_4)_2SO_4(aq))$

These salts can be crystallized out and are similar to ordinary metallic salts.

Reaction of aqueous ammonia with metallic ions

When aqueous ammonia is added to a solution of metallic salt, it forms precipitate of the insoluble metal hydroxide.

For example when a solution containing copper (II) ions e.g. copper (II) sulphate solution is put in a test tube and aqueous solution of ammonia added to it a little, pale blue precipitates of copper (II) hydroxide are formed.

$$CuSO_4(aq) + 2NH_4OH(aq) \longrightarrow (NH_4)_2SO_4(aq) + \underset{\text{(pale blue precipitate)}}{Cu(OH)_2(s)}$$

Ionic equation

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

When the ammonia solution is added until in excess, the pale blue precipitates dissolve giving a deep blue solution. The deep blue solution is due to the formation of a complex salt containing tetra amine copper (II) ion.

$$Cu(OH)_2(s) + 4NH_3(aq) \longrightarrow Cu(NH_3)_4^{2+}(aq) + 2OH^-(aq)$$

Examples of metallic ions that react with ammonium solution are;

$Fe^{2+}(aq) + 2OH^-(aq) \longrightarrow Fe(OH)_2(s)$

(Green precipitate)

$Fe^{3+}(aq) + 3OH^-(aq) \longrightarrow Fe(OH)_3(s)$

(Brown precipitate)

$Pb^{2+}(aq) + 2OH^-(aq) \longrightarrow Pb(OH)_2(s)$

(White precipitate)

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

(White precipitate)

$Al^{3+}(aq) + 3OH^-(aq) \longrightarrow Al(OH)_3(s)$

(White precipitate)

However, sodium hydroxide and ammonium hydroxide are different in that:

1. The pale blue precipitates of copper(II) hydroxide do not dissolve in excess sodium hydroxide but dissolves in excess ammonia solution forming a deep blue solution.
2. The amphoteric aluminium and lead ions dissolve in excess sodium hydroxide but not in excess ammonia solution.
3. Zinc dissolves in both excess sodium hydroxide and ammonia solution but the reactions are different.

$Zn(OH)_2(s) + 2OH^-(aq) \longrightarrow Zn(OH)_4^{2-}(aq)$

(zincate ion)

$Zn(OH)_2(s) + 4NH_3(aq) \longrightarrow Zn(NH_3)_4^{2+}(aq) + 2OH^-(aq)$

(tetrammine zinc ion)

Manufacture of ammonia (Haber process)

Large scale manufacture of ammonia is done through the Haber process by the synthesis of ammonia from its elements, hydrogen and nitrogen.

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + heat$

The raw materials for the process are hydrogen and nitrogen.

Process

Dry nitrogen and hydrogen in the ration of 1:3 by volume are mixed in the presence of a finely divided iron catalyst and the mixture heated.

Condition

The heating is carried out at a very high pressure of about 200-400 atmospheres and at a temperature of 400-500˚C in the presence of a catalyst (finely divided iron impregnated with alumina). The reaction is exothermic and reversible.

Note

1. The two gases (hydrogen and nitrogen) must first be purified. This is because the impurities may poison the catalyst. The impurities that must be removed include: water vapour; dust; carbon dioxide and sulphur dioxide.
2. The raw materials for the process are obtained from various sources e.g. the nitrogen is obtained from fractional distillation of liquid air, hydrogen from steam reforming of natural gas.

The ammonia gas produced is then liquefied and stored for future use.

Discussion conditions for the Haber Process

Effect of pressure

The formation of ammonia proceeds with a decrease in volume i.e. 4 volumes of reactants give 2 volumes of products.

$$\underbrace{N_2(s) + 3H_2(g)}_{\text{4 volumes}} \rightleftharpoons \underbrace{2NH_3(g)}_{\text{2 volumes}}$$

High pressure would increase the yield of ammonia since the reaction proceeds with a decrease in volume. There fore a pressure of 200-500 atm is used. At a high pressure, the reacting molecules collide more frequently thus increasing the rate of the reaction.

Effect of temperature

The formation of ammonia from its elements is an exothermic reaction and there fore low temperature will cause a better yield of ammonia. But at low temperature, the reaction will be slow thus a moderate temperature of 400-500˚C is used together with a catalyst.

Effect of a catalyst

To make the reaction proceed faster, a catalyst is used. The catalyst used is finely divided iron impregnated with alumina. The catalyst should be finely divided to increase on the surface area since the reaction occurs at the surface.

Effect of concentration

Since the reaction is reversible, we use in excess any of the reactants inorder to get a better yield of ammonia. However, in practice, nitrogen and oxygen are used in the ratio of 1:3 respectively.

Uses of ammonia

- It is used in the manufacture of fertilizers like ammonium phosphate and urea.
- Ammonia solution is used to soften hard water.
- Ammonia is used in the manufacture nitric acid.
- Liquid ammonia is used in large scale refrigerating plants such as in ships and ware houses.
- It is used in the manufacture of sodium carbonate in the Solvay process.

Ammonium salts

a) All common ammonium salts are white crystalline substances; soluble in water and are ionic compounds.

b) The sulphates, nitrates and chlorides dissolve in water to form an acidic solution.

$$NH_4Cl(s) + H_2O(l) \longrightarrow NH_4OH(aq) + HCl(aq)$$

The ammonium chloride dissolves to form a weak ammonium hydroxide solution and strong hydrochloric acid. The ammonium hydroxide is too weak to neutralize the effect of the strong hydrochloric acid and thus the resultant solution is acidic in nature.

c) All ammonium salts sublime except ammonium sulphate.

$$NH_4Cl(s) \longrightarrow NH_3(g) + HCl(g)$$

$NH_4NO_3(s) \longrightarrow N_2O(g) + 2H_2O(g)$

$(NH_4)_2CO_3(s) \longrightarrow CO_2(g) + H_2O(g) + 2NH_3(g)$

It is only ammonium sulphate that does not sublime

$(NH_4)_2SO_4(s) \longrightarrow H_2SO_4(g) + 2NH_3(g)$

d) All ammonium salts liberate ammonia gas upon heating with an alkali e.g. sodium hydroxide solution.

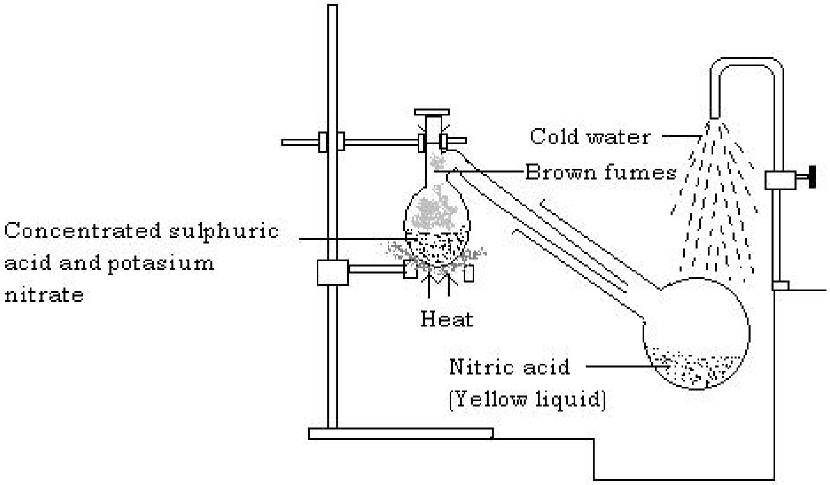
$NH_4^+ aq) + OH^-(aq) \longrightarrow NH_3(g) + H_2O(l)$

This reaction serves as a confirmatory test for the presence of ammonium ion.

NITRIC ACID

Laboratory preparation

Set up



Procedure

- Assemble the glass apparatus as shown in the diagram above.
- Put some potassium nitrate crystals or sodium nitrate crystals in the bulb of the retort with concentrated sulphuric acid.

- Heat the mixture gently and then collect the nitric acid in a water cooled receiver.

Observation

The potassium nitrate crystal gradually dissolves and effervescence occurs.

$KNO_3(s) + H_2SO_4(aq) \longrightarrow KHSO_4(s) + HNO_3(g)$

The nitric acid distills and collects in the cooled receiver as a yellow liquid (the yellow colour is due to the presence of dissolved nitrogen dioxide gas), while drops of the acid can be seen running down the bulb and neck of the retort. The brown fumes are nitrogen dioxide formed by slight decomposition of the nitric acid by heat.

$4HNO_3(l) \longrightarrow 2H_2O(l) + 4NO_2(g) + O_2(g)$

NB. The apparatus used must be glass because nitric acid quickly attacks other materials like cork and rubber tubing.

Any metallic nitrate when heated with concentrated sulphuric acid produces nitric acid.

$NaNO_3(s) + H_2SO_4(l) \longrightarrow NaHSO_4(aq) + HNO_3(aq)$

Industrial Preparation of nitric acid (The Ostwald's process)

Nitric acid is manufactured by the catalytic oxidation of ammonia and then dissolving the products in water.

Major steps

1. Ammonia from the Haber process is mixed with excess air and passed over a platinum (90%)/rhodium (10%) gauze catalyst. The catalyst is heated to red hot to start the reaction and since the reaction is exothermic, no heating is required once the reaction starts. Here ammonia is oxidized to colorless dinitrogen gas.

   $4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(l)$

2. The nitrogen monoxide formed is rapidly cooled and combines with the oxygen from the excess air to form nitrogen dioxide.

   $2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$

3. The nitrogen dioxide, in the presence of more air is absorbed in hot water to form the nitric acid.

   $2H_2O(l) + 4NO_2(g) + O_2(g) \longrightarrow 4HNO_3(aq)$

Properties of nitric acid

Physical properties

i) It is a colourless fuming liquid (if pure)

ii) It is corrosive just like other acids

iii) It turns blue litmus red and has no effect on red litmus paper

Chemical properties

a) Thermal decomposition

Upon heating, concentrated nitric acid decomposes to give off brown fumes of nitrogen dioxide, oxygen and water vapour.

$4HNO_3(aq) \longrightarrow 4NO_2(g) + O_2(g) + 2H_2O(l)$

b) Acid reactions of nitric acid

i) Dilute nitric acid is neutralized by alkalis to form a salt and water only.

$NaOH(q) + HNO_3(aq) \longrightarrow NaNO_3(aq) + H_2O(l)$

ii) Carbonates are decomposed by dilute nitric acid to form the corresponding nitrate, water and carbon dioxide.

$Na_2CO_3(aq) + 2HNO_3(aq) \longrightarrow 2NaNO_3(aq) + CO_2(g) + H_2O(l)$

iii) Hydrogen gas is liberated when very dilute nitric acid is made to react with highly electropositive metals like magnesium and zinc.

$Mg(s) + 2HNO_3(aq) \longrightarrow Mg(NO_3)_2(aq) + H_2(g)$

Other metals do not react with dilute nitric acid to produce hydrogen, they are simply oxidized to their corresponding nitrates and the nitric acid is reduced to nitrogen dioxide.

c) Oxidation reactions of nitric acid

Concentrated nitric acid is a powerful oxidizing agent i.e. it readily gives up its oxygen. It converts most metals to their corresponding nitrates and non metals such as carbon are oxidized to their oxides. In all cases, nitric acid is itself reduced to nitrogen dioxide. For example

i) Concentrated nitric acid oxidizes brown copper metal to copper (II) nitrate and the nitric acid is itself reduced to nitrogen dioxide.

$Cu(s) + 4HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$

However, when moderately concentrated nitric acid (50%) is used, nitrogen monoxide is instead formed.

$3Cu(s) + 8HNO_3(aq) \longrightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + 4H_2O(l)$

No heating is required for this reaction.

Concentrated nitric acid renders iron and aluminium passive due to formation of an oxide layer which serves as a protective coating and therefore prevents the metals from reacting any further.

ii) Hot concentrated nitric acid oxidizes carbon to carbon dioxide and itself reduced to nitrogen dioxide.

$C(s) + 4HNO_3(l) \longrightarrow CO_2(g) + 4NO_2(g) + 2H_2O(l)$

iii) Hot concentrated nitric acid oxidizes red hot phosphorus to phosphoric acid.

$P(s) + 5HNO_3(aq) \longrightarrow H_3PO_4(aq) + 5NO_2(g) + H_2O(l)$

iv) Concentrated nitric acid oxidizes sulphur to sulphuric acid it self reduced to nitrogen dioxide.

$S(s) + 6HNO_3(aq) \longrightarrow H_2SO_4(aq) + 6NO_2(g) + 2H_2O(l)$

v) When hydrogen sulphide is bubbled through a concentrated solution of nitric acid, a yellow deposit of sulphur is formed and nitric acid is reduced to brown fumes of nitrogen dioxide.

$H_2S(g) + 2HNO_3(aq) \longrightarrow S(s) + 2NO_2(g) + H_2O(l)$

vi) Concentrated nitric acid oxidizes acidified green iron(II) sulphate to brown/yellow iron(III) sulphate and itself reduced to nitrogen monoxide (colourless) which is then oxidized immediately in air to form nitrogen dioxide (brown fumes).

$FeSO_4(aq) + 3H_2SO_4(aq) + 2HNO_3(aq) \longrightarrow Fe_2(SO_4)_3(aq) + 2NO(g) + 4H_2O(l)$

$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$

Uses of nitric acid

- It is used in the manufacture of dyes.
- It is used in the manufacture of fertilizers e.g. ammonium nitrate.
- Used in the manufacture of explosives.

- Since it is an oxidizing agent, it is used in the manufacture of nylon.
- It is used in the refining of precious metals.

NITRATES

These are salts of nitric acid. All nitrates are soluble in water.

Action of heat on nitrates

All the nitrates decompose on heating. The thermal decomposition of metal nitrates depends upon the position of the metal in the reactivity series.

When potassium and sodium nitrates are heated, they melt into colourless liquid decompose to give pale yellow nitrites and oxygen gas. E.g.

$$2NaNO_3(s) \longrightarrow 2NaNO_2(l) + O_2(g)$$

(Pale yellow)

From calcium to copper, the nitrates are decomposed on heating to form oxides of the metal, brown fumes of nitrogen dioxide gas and oxygen gas (relights a glowing splint). E.g. when white crystals of lead(II) nitrate is heated, it produces a cracking sound, brown fumes of nitrogen dioxide and a colourless gas that relights a glowing splint (oxygen gas). A residue of lead(II) oxide (brown when hot; yellow when cold) is left.

$$Pb(NO_3)_2(s) \longrightarrow PbO(s) + NO_2(g) + O_2(g)$$

(White crystals) (Brown-hot: yellow-cold)

$$Mg(NO_3)_2(s) \longrightarrow MgO(s) + NO_2(g) + O_2(g)$$

(White) (white solids)

$$Zn(NO_3)_2(s) \longrightarrow ZnO(s) + NO_2(g) + O_2(g)$$

(White) (Yellow-hot: white-cold)

$$Cu(NO_3)_2(s) \longrightarrow PbO(s) + NO_2(g) + O_2(g)$$

(Blue) (Black)

From silver to gold, their nitrates decompose to give corresponding metals, nitrogen dioxide (brown fumes) and oxygen gas (colourless gas that relights a glowing splint).

$2AgNO_3(s) \longrightarrow 2Ag(s) + 2NO_2(g) + O_2(g)$

$Hg(NO_3)_2(s) \longrightarrow Hg(s) + 2NO_2(g) + O_2(g)$

Ammonium nitrate decomposes explosively when heated into dinitrogen oxide and water.

$NH_4NO_3(s) \longrightarrow N_2O(g) + 2H_2O(l)$

Test for nitrates

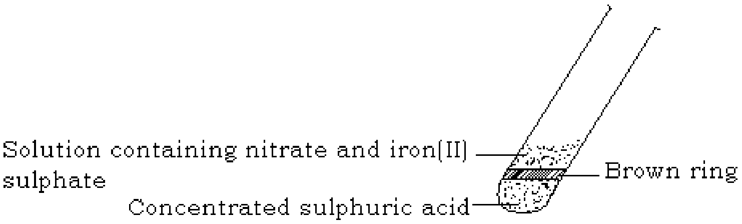
All nitrates irrespective of their position in the reactivity series under go the same reaction with iron(II) sulphate and concentrated sulphuric acid and the reaction is used as a test for all soluble nitrates.

Procedure

To the nitrate solution in a test tube, add an equal volume of freshly prepared iron(II) sulphate solution. The test tube is tilted and concentrated sulphuric acid is carefully poured down the side of the test tube.

Observation

A brown ring is formed at a junction between sulphuric acid and iron(II) sulphate. The brown ring is of a compound with the formula $FeSO_4.NO$.



The brown ring is formed when concentrated sulphuric acid reacts with nitrate ions to give nitric acid.

$H_2SO_4(aq) + NO_3^-(aq) \longrightarrow 2HNO_3(aq) + SO_4^{2-}(aq)$

The nitric acid formed oxidizes iron(II) to iron(III) and itself is reduced to nitrogen monoxide.

$Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + e^-$

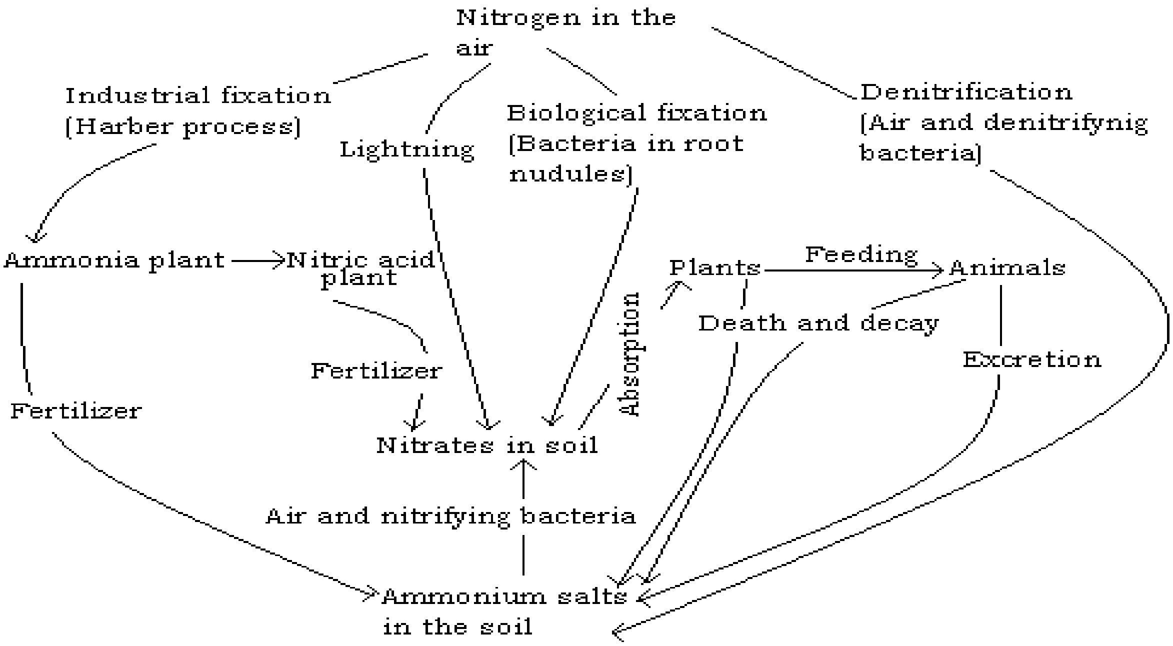
$4HNO_3(aq) \longrightarrow 4NO(g) + 2H_2O(l) + 3O_2(g)$

The nitrogen monoxide then combines with some of the iron(II) sulphate to form the complex $FeSO_4.NO$ (brown ring).

$FeSO_4(aq) + NO(g) \longrightarrow FeSO_4.NO(aq)$

The nitrogen cycle

This is a balance that exists between reactions that take nitrogen out of the air and out of the soil and reactions that put nitrogen into the air and into the soil. Nitrogen constitutes 78% by volume of air and it is an essential element in all living things.



Plants obtain nitrogen mainly in form of dissolved nitrates by absorption from the soil. When the nitrogen in the soil is not replenished, it leads to infertility and poor yield in crops. Animals obtain nitrogen by feeding on plants. Nitrogen is supplied to the soil through: death and decay of plants and animals; excretion by animals; nitrogen fixation by colonies of bacteria in root nodules and bacteria that live freely in the soil; lightning -causes some slight combination of nitrogen and oxygen which leads to passage of nitrogen into the soil as nitrates; industrial fixation through Haber process forming nitrogenous fertilizers such as ammonium sulphate. Denitrifying bacteria such as *Pseudomonas denitrificans* convert ammonium salts in the soil to atmospheric nitrogen

Sample question on Nitrogen and its compounds

1. *Outline briefly how a sample of nitrogen can be obtained from (a) air, (b) an ammonium salt.*

   Describe an outline of the industrial process showing how nitrogen is converted to ammonia. The reaction involved in this process is reversible. Indicate three ways in which the yield of ammonia can be made maximum. State two industrial uses of ammonia.
2. *Make a fully labeled drawing and give equation to show how you would prepare dry sample of ammonia in the laboratory starting from a named ammonium salt and a named alkali.*

   Giving equations and reaction conditions, outline how nitric acid is manufactured from ammonia.
3. *Describe the action of heat on ammonium chloride. By what reaction would you obtain (i) ammonia from ammonium chloride (ii) nitrogen from ammonia? How can you account for the fact that a solution of ammnia in water will turn litmus blue and give brown precipitate when mixed with a olution of iron(II) chloride?*
4. *Describe how ammonia is manufactured from its elements. State the source of each element. Outline three differences between nitrogen/hydrogen mixture and ammonia. How would you show that ammonia is a very soluble gas?*
5. *Explain how ammonia is converted into nitric acid on a large scale. Describe two reactions in which nitric acid (i) acts as an acid (ii) acts as an oxidizing agent.*

Outline some uses of nitric acid. Calculate the percentage of nitrogen in pure ammonium nitrate.

Describe with equations what happens when ammonia is passed (i) into dilute sulphuric acid, (ii) over heated copper(II) oxide.

6. *Nitric acid can be prepared in the laboratory by heating solid sodium nitrate with concentrated sulphuric acid. Make a labeled drawing of the apparatus you would use for the preparation and write equation for the reaction. Explain why sulphuric acid is used in this preparation rather than hydrochloric acid.*
7. *Describe an experiment to show how nitrogen dioxide is prepared in the laboratory. Describe the oxidation reaction of nitrogen dioxide, use equations to illustrate. Mention two uses of nitrogen dioxide.*
8. *With the aid of equations, describe the reactions of ammonia with: oxygen, hydrogen chloride, chlorine, water, metal oxides and metallic ions.*

CHLORINE AND ITS COMPOUND

The four non metals fluorine, chlorine, bromine and iodine make up a family of related elements called **halogens** meaning salt producers since they react with most metals to form electrovalent salt like compounds.

Preparation of chlorine

Chlorine can be generally prepared by the removal of hydrogen from hydrochloric acid i.e. oxidation of hydrochloric acid. This can be done by a substance containing oxygen (an oxidizing agent) that will combine with the hydrogen to form water. The oxygen for the oxidation of hydrochloric acid is provided by any of the following:

Potassium manganate(VII) or potassium permanganate ($KMnO_4$)

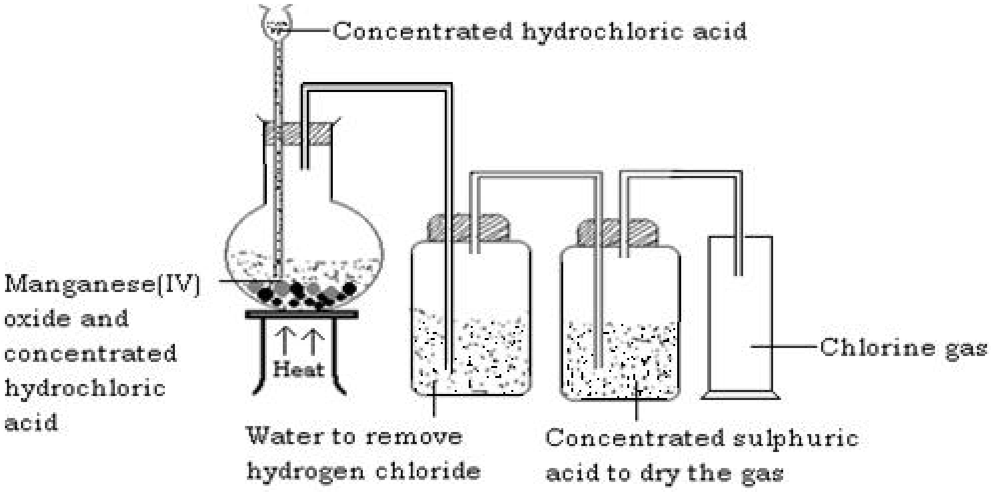
i) Manganese(IV) oxide (MnO_2)
ii) Lead(II) oxide (PbO_2)
iii) Tri lead tetraoxide (Pb_3O_4)

The common oxidizing agents used in the preparation of chlorine are potassium permanganate and manganese(IV) oxide.

Preparation of chlorine from hydrochloric acid by oxidation with manganese(IV) oxide

Procedure

- Place some manganese(IV) oxide into a flask. Lumps are preferably used as the powder is very reactive
- Fit the apparatus as shown below
- Pour concentrated hydrochloric acid down the thistle funnel
- Heat the mixture in the flask



Observation

Effervescence occurs evolving a greenish yellow gas (chlorine gas). Chlorine is evolved together with a small amount of hydrogen chloride gas (misty fumes) which is removed by passing it through the first bottle containing water.

The gas is then dried using concentrated sulphuric acid and collected by down ward delivery method since it is denser than air.

Equation

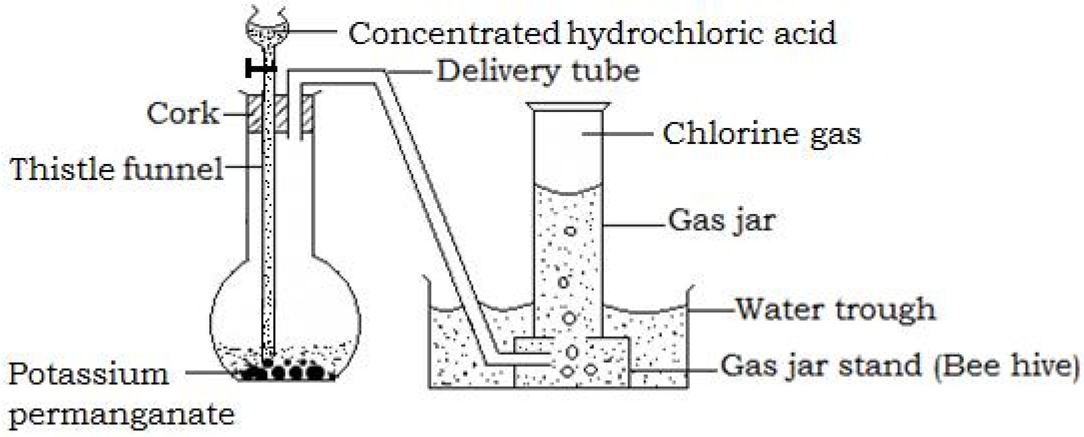
$MnO_2(s) + 4HCl(aq) \longrightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$

Chlorine being a poisonous gas, it is prepared in the fume cupboard.

Preparation of chlorine from hydrochloric acid by oxidation with potassium permanganate

Procedure

- Place solid potassium permanganate in a flask
- Fit the apparatus as shown below



- Drop on the potassium permanganate concentrated hydrochloric acid from the tap funnel. As each drop reaches the Manganese(IV) oxide, a corresponding quantity of chlorine, a greenish yellow gas is evolved.

Equation

$2KMnO_4(aq) + 16HCl(aq) \longrightarrow 2KCl(aq) + 8H_2O(l) + 2MnCl_2(aq) + 5Cl_2(g)$

The gas collection is by down ward deliver of over brine of hot water of in a gas syringe.

This is the most convenient laboratory method because of the following reasons:

- It does not require heating
- Rate of production of chlorine can be regulated
- If the gas is collected over brine, the experiment can be done out of the fume chamber

Chlorine can be prepared from sodium chloride by adding concentrated sulphuric acid to an intimate mixture of sodium chloride and Manganese(IV) oxide. Traces of hydrogen chloride produced are removed by passing the gas over water. The gas is dried and collected by down ward delivery method.

$$2NaCl(s) + MnO_2(s) + 2H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + MnSO_4(aq) + 2H_2O(l) + Cl_2(g)$$

Chlorine can also be prepared from bleaching powder. The bleaching powder is placed in a flask and a dilute acid e.g. nitric acid, hydrochloric acid is added to the powder. Effervescence occurs as a greenish yellow gas (chlorine gas) is evolved. The gas is then dried and collected by any of the methods discussed.

$$CaOCl_2(s) + 2HNO_3(aq) \longrightarrow Ca(NO_3)_2(aq) + H_2O(l) + Cl_2(g)$$

$$CaOCl_2(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + Cl_2(g)$$

Industrial manufacture of chlorine

Chlorine is produced commercially by electrolysis of sodium chloride (brine) solution. Chlorine is evolved at the anode of a specially designed cell, and since the other electrode product (sodium hydroxide, at the cathode) reacts with chlorine, they must be kept apart. This is effected by a circulating mercury diaphragm.

When current is passed through the electrolyte, decomposition takes place as follows.

$$NaCl(aq) \longrightarrow Na^+(aq) + Cl^-(aq)$$

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

At the cathode

Both sodium and hydrogen ions move towards the cathode. Because of the nature of the mercury electrode, it will influence the discharge of sodium ion (Na^+) in preference to hydrogen ions (H^+) since its discharge requires less energy than the discharge of hydrogen.

$$Na^+(aq) + e^- \longrightarrow Na(s)$$

The sodium formed combines with the mercury to form sodium amalgam. The sodium amalgam encounters water in contact with steel grids on which hydrogen has a very low over voltage. Sodium hydroxide is formed and hydrogen gas is liberated.

Equation

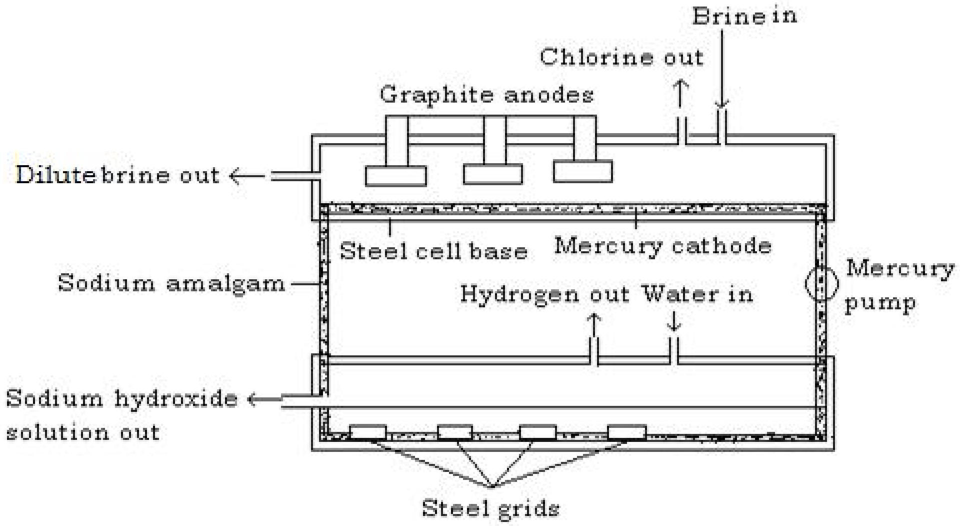
$$Na/Hg(l) + H_2O(l) \longrightarrow NaOH(aq) + H_2(l) + Hg(l)$$

At the anode

Both chloride and hydroxyl ions move to the anode. Due to the high concentration of the chloride ions, it is discharged in preference to hydroxyl ions, therefore forming chlorine gas at the anode.

$$2Cl^-(aq) - 2e^- \longrightarrow Cl_2(g)$$

NB. Carbon anode is used as it is relatively inert and not easily attacked by chlorine. Chlorine is very reactive and attacks most metals.



Properties of chlorine

Physical properties

- It is a greenish yellow gas with a chocking, unpleasant, irritating smell.
- It is slightly soluble in water forming a yellowish chlorine water which is a mixture of hydrochloric acid and hypochlorous acid.

$$H_2O(l) + Cl_2(g) \longrightarrow HOCl(aq) + HCl(aq)$$

- It is denser than air.
- It turns damp blue litmus paper red the bleaches it. Its bleaching action is due to the formation of hypochlorous acid. Dry chlorine does not bleach.
- Dry Chlorine does not bleach and extinguishes a burning splint.

Test for chlorine:

It is a greenish yellow gas which turns moist blue litmus paper red then bleaches it. This is because, the gas is acidic.

Chemical properties

1. Chlorine as a bleaching agent

Pour a little litmus solution into a gas jar of chlorine.

Observation: The litmus immediately turns colourless.

Chlorine bleaches colour from most dyes and will remove colour from writing ink (but not printer's ink, which consists mainly of carbon which chlorine does not attack.)

The bleaching action

Chlorine reacts with water to form hypochlorous acid

$$H_2O(l) + Cl_2(g) \longrightarrow HOCl(aq) + HCl(aq)$$

The hypochlorous acid is a very reactive compound and readily gives up its oxygen to the dye, to form a colourless compound.

$$\underset{\text{(Coloured)}}{\text{Dye}} + HOCl \longrightarrow HCl(aq) + \underset{\text{(Colourless)}}{(\text{Dye} + O)}$$

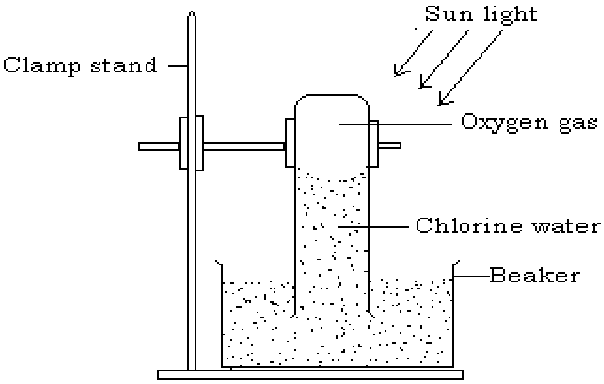
The hypochlorous acid turns the coloured dye to colourless by **oxidation** reaction since it gives up its oxygen to the coloured dye.

This indicates that dry chlorine will not bleach since there will be no hypochlorous acid formed.

Hypochlorous acid is also used to kill bacteria and germs in drinking water, swimming pools and in sewage treatment.

2. Effect of sunlight on chlorine water

Pass chlorine gas into water in a beaker until water becomes yellow green in colour i.e. chlorine water is formed. Fill a long tube with this chlorine water and invert it in a beaker containing some of the water and expose it to bright sunlight.



After some times, a gas collects in the tube and on applying a glowing splint, the gas rekindled it showing that the gas is oxygen.

Equation

$2Cl_2(g) + 2H_2O(aq) \longrightarrow 4HCl(aq) + O_2(g)$

The above reaction occurs instantaneously in two stages as

$H_2O(l) + Cl_2(g) \longrightarrow HOCl(aq) + HCl(aq)$

$2HOCl(aq) \longrightarrow 2HCl(aq) + O_2(g)$

3. Action of chlorine on alkali

a) On cold dilute alkaline solution

Chlorine reacts with dilute aqueous solution of sodium and potassium forming pale yellow solution of the hypochlorite and chloride of the metal.

$Cl_2(g) + 2NaOH(aq) \longrightarrow \underset{\text{(Sodium hypochlorite)}}{NaOCl(aq)} + \underset{\text{(Sodium chloride)}}{NaCl(aq)} + H_2O(l)$

$Cl_2(g) + 2KOH(aq) \longrightarrow \underset{\text{(Potassium hypochlorite)}}{KOCl(aq)} + \underset{\text{(Potassium chloride)}}{KCl(aq)} + H_2O(l)$

Ionically

$Cl_2(g) + 2OH^-(aq) \longrightarrow OCl^-(aq) + Cl^-(aq) + H_2O(l)$

b) On hot concentrated aqueous solution

If chlorine is passed into a hot concentrated solution of potassium, a mixture of potassium chlorate and potassium chloride are formed. A similar reaction occurs when chlorine reacts with concentrated sodium and calcium hydroxide.

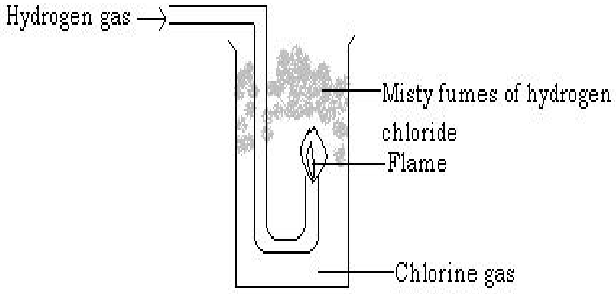
$6KOH(aq) + 3Cl_2(g) \longrightarrow KClO_3(aq) + 5KCl(aq) + 3H_2O(l)$

Or

$3Cl_2(g) + 6OH^-(aq) \longrightarrow ClO_3^-(aq) + 5Cl^-(aq) + 3H_2O(l)$

4. Reaction of chlorine with hydrogen

When a jet of burning hydrogen is lowered into a gas jar of chlorine, it continuous to burn with a white flame producing steamy fumes of hydrogen chloride and the greenish yellow colour of chlorine disappears.



$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$

A mixture of hydrogen and chlorine also explodes when exposed to bright sunlight. This shows the great affinity of chlorine for hydrogen. The reaction is slow in dim sun light and reaction does not take place in the absence of light.

5. Chlorine as an oxidizing agent

An oxidizing agent is one which can

- i) Remove hydrogen from a compound
- ii) Accept electrons donated by metals

Chlorine removes hydrogen from many compounds as well as accepts electrons from metals. The oxidizing property of chlorine is illustrated by the following reactions.

i) Reaction with hydrogen sulphide

When a gas jar of hydrogen sulphide is inverted over a gas jar of chlorine, yellow solids of sulphur and white fumes of hydrogen chloride are formed.

$H_2S(g) + Cl_2(g) \longrightarrow 2HCl(g) + S(s)$

Hydrogen sulphide has been oxidized to sulphur and chlorine itself was reduced to hydrogen sulphide.

The reaction proceeds as follows

$H_2S \longrightarrow 2H^+ + S^{2-}$

$S^{2-} \longrightarrow S + 2e^-$(oxidation)

$Cl_2 + 2e^- \longrightarrow 2Cl^-$ (reduction)

ii) Reaction with ammonia

Ammonia burns in chlorine and is oxidized to nitrogen and chlorine itself is reduced to hydrogen chloride.

$2NH_3(g) + 3Cl_2(g) \longrightarrow N_2(g) + 6HCl(g)$

The hydrogen chloride produced reacts with excess ammonia forming white fumes of ammonium chloride.

$NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s)$

Overall equation for the reaction is

$8NH_3(g) + 3Cl_2(g) \longrightarrow N_2(g) + 6NH_4Cl(s)$

iii) Oxidation of iron(II) to iron(III)

Chlorine accepts electrons like other oxidizing agents and is converted to chlorine ion.

$Cl_2 + 2e^- \longrightarrow 2Cl^-$

The electron that the chlorine accepts is supplied by a reducing agent which in this case is iron(II).

$Fe^{2+} \longrightarrow Fe^{3+} + e^-$

When chlorine gas is bubbled through a pale green solution of iron(II) chloride, it changes to yellow due to the formation of iron(III) chloride. When sodium hydroxide is added to this solution it forms red brown precipitates. This indicates that iron(II) chloride has been oxidized to iron(III) chloride.

$2FeCl_2(aq) + Cl_2(g) \longrightarrow 2FeCl_3$

Or

$2Fe^{2+}(aq) + Cl_2(g) \longrightarrow 2Fe^{3+}(aq) + 2Cl^-(aq)$

iv) Reaction of chlorine with hydrocarbons e.g. turpentine ($C_{10}H_{16}$)

Hydrocarbons consist of only carbon and hydrogen. Chlorine removes hydrogen from hydrocarbon to form hydrogen chloride and black carbon particles are left.

Procedure

Warm a little turpentine in a dish. Dip into it a filter paper and drop the filter paper with turpentine in a gas jar containing chlorine

Observation

A red flash accompanied by a violent reaction occurs. Black cloud of solid particles of carbon is produced and hydrogen chloride gas is also formed.

$C_{10}H_{16}(l) + 8Cl_2(g) \longrightarrow 10C(s) + 16HCl(g)$

The presence of hydrogen chloride can be shown by blowing the fumes from ammonia bottle across the top of the jar. Dense white fumes of ammonium chloride are formed.

$NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s)$

6) Displacement reaction of chlorine

Chlorine being more reactive than the other halogens displaces them from their salts.

i) Reaction of chlorine with potassium bromide

When chlorine is bubbled through a saturated solution of potassium bromide, the clear solution immediately turns red (due to formation of bromine water) and a drop of a red liquid (bromine) is observed at the bottom of the boiling tube.

$2KBr(aq) + Cl_2(g) \longrightarrow 2KCl(aq) + Br_2(l)$

Ionically

$Br^-(aq) + Cl_2(g) \longrightarrow Cl^-(aq) + Br_2(l)$

Bromine dissolves in tetra chloromethane to form a reddish brown solution.

ii) Displacement of iodine

When chlorine is bubble through a solution of potassium iodide, the clear solution turns to the characteristic dark brown "iodine" colour and a black solid (iodine) is deposited as iodine is only slightly soluble in water. On warming the solution, the characteristic violet vapour of iodine is seen.

$2KI(aq) + Cl_2(g) \longrightarrow 2KCl(aq) + I_2(l)$

Ionically

$Br^-(aq) + Cl_2(g) \longrightarrow 2Cl^-(aq) + I_2(l)$

Iodine dissolves in tetra chloromethane to form a violet solution.

7) Reaction of chlorine with non metals

i) Phosphorus

When a piece of dry yellow phosphorus is lowered in a gas jar of chlorine, it burns spontaneously giving off white fumes of chlorides of phosphorus mainly phosphorus tri chloride (PCl_3)

$P_4(s) + 6Cl_2(g) \longrightarrow 4PCl_3(g)$

(Phosphorus(III) chloride)

$P_4(s) + 10Cl_2(g) \longrightarrow 4PCl_5(g)$

(Phosphorus(V) chloride)

ii) Sulphur

When dry chlorine is passed over molten sulphur in a distilling flask, connected to a condenser, a reddish liquid distills over.

$2S(l) + Cl_2(g) \longrightarrow S_2Cl_2(l)$

(disulphur dichloride)

8) Reaction of chlorine with metals

i) Reaction with sodium and magnesium

Burning sodium and magnesium continue to burn is chlorine forming white fumes of sodium and magnesium chlorides respectively.

$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$

$Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)$

ii) Reaction with Dutch metal

Dutch metal is an alloy of copper and zinc. When a piece of Dutch metal is dropped in a gas jar of chlorine and heated, it burns with a green flame to form copper(II) chloride and Zinc chloride.

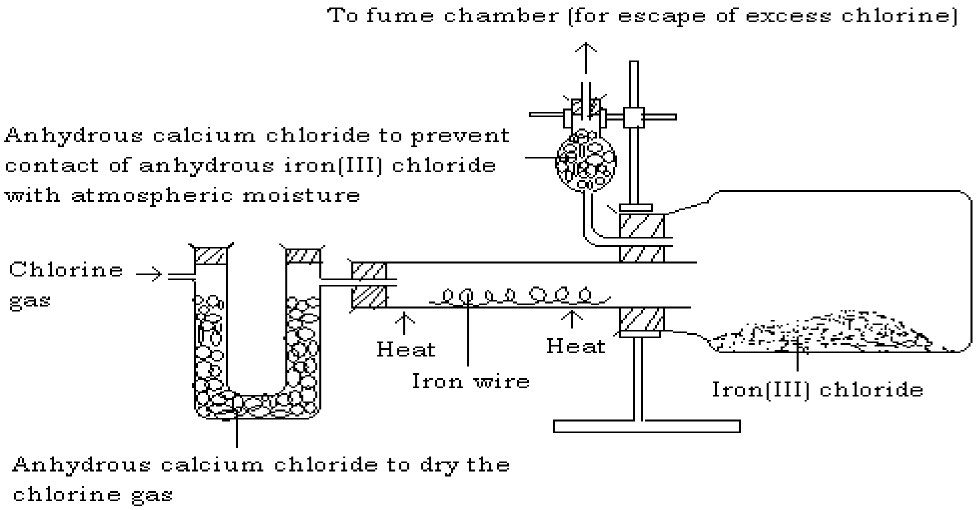
$Cu(s) + Cl_2(g) \longrightarrow CuCl_2(s)$

$Zn(s) + Cl_2(g) \longrightarrow ZnCl_2(s)$

iii) Action of chlorine on iron (preparation of iron(III) chloride by direct synthesis)

Procedure

- Place a coil of iron wire in a hard glass tube in the apparatus below. The iron coil must be free of rust.



- Pass a stream of dry chlorine over it.
- Heat the wire and stop the heating the moment the reaction starts.

Observation

The wire glows and the reaction continues without application of heat indicating that the reaction is **exothermic.** Black crystals of iron(III) chloride sublimes and collects in the small bottle which acts as a condenser.

$2Fe(s) + 3Cl_2(g) \longrightarrow 2FeCl_3(s)$

Formation of iron(III) chloride shows that chlorine is an oxidizing agent. The iron (II) chloride formed is immediately oxidized to iron(III) chloride.

The black crystals of anhydrous iron(III) chloride should be placed in a desiccator as they are very deliquescent.

Note

1. Sodium chloride can be made in a similar way.

   $2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$

2. Iron(II) chloride (white solid) is made in the same way using dry hydrogen chloride instead of chlorine.

$Fe(s) + 2HCl(g) \longrightarrow FeCl_2(s) + H_2(g)$

Uses of chlorine

1. Chlorine is extensively used as a bleaching agent and in the manufacture of bleaching agents.
2. It is used for making domestic antiseptic solutions such as sodium hypochlorite.
3. Chlorine is used in the manufacture of chlorates used for example as weed killers.
4. Manufacture of hydrogen chloride which is used in the manufacture of plastic like PVC (polyvinylchloride).
5. Manufacture of many organic chemicals e.g. tetra chloromethane (CCl_4), 1,1,2-trichloroethene (C_2HCl_3). These compounds are solvents used to remove grease from other substances (degreasing agents), dry cleaning fluids.
6. Chlorine is used to sterilize water for domestic and industrial uses i.e. it kills bacteria and other germs in water and is used in water purification process.

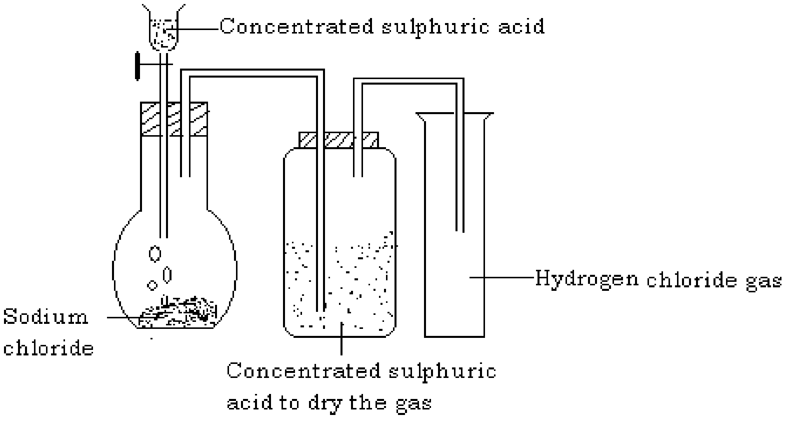
HYDROGEN CHLORIDE

Laboratory preparation

Hydrogen chloride can be prepared in the laboratory by the action of concentrated sulphuric acid on common salt (rock salt).

Procedure

- Place sodium chloride in flask and fit the apparatus as shown below.



- Add concentrated sulphuric acid down the funnel.

Effervescence occurs and misty fumes are observed. The gas is passed through a wash bottle containing concentrated sulphuric acid to dry the gas and it is collected by down ward delivery method as it is denser than air.

$H_2SO_4(l) + NaCl(s) \longrightarrow NaHSO_4(s) + HCl(g)$

Note

1. Sodium chloride is used because it is cheap and readily available.
2. The reaction proceeds in the cold though a further yield is obtained in the industrial process by heating.
3. The sulphate is not obtained under laboratory conditions because its formation requires a higher temperature but is obtained during the industrial process.

Test for hydrogen chloride

1. When the gas is bubbled through a solution of silver nitrate and nitric acid, it forms white precipitates of silver chloride.

$AgNO_3(aq) + HCl(g) \longrightarrow AgCl(s) + HNO_3(aq)$

Or

$Ag^+ (aq) + Cl^-(aq) \longrightarrow AgCl(s)$

2. Hydrogen chloride forms dense white fumes of ammonium chloride with ammonia.

$NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s)$

Properties of hydrogen chloride gas

Physical properties

1. It is a colourless gas with an irritating, choking smell.
2. It forms misty fumes in damp air as it forms tiny droplets of hydrochloric acid.
3. It turns moist blue litmus paper red.
4. It is also soluble in methyl benzene (toluene)
5. Hydrogen chloride gas is denser than air.
6. It is very soluble in water. When it dissolves in air, it forms hydrochloric acid.

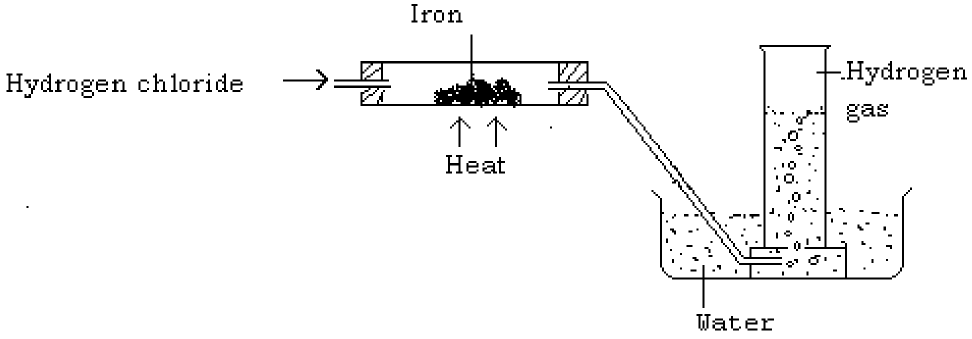
The high solubility of hydrogen chloride gas in water can be demonstrated by the fountain experiment as that of ammonia.

Chemical properties

1. **Deduction of the composition of hydrogen chloride**

a) **Action of dry hydrogen chloride gas on heated iron**

When dry hydrogen gas is passed over heated iron, the gaseous product that is collected over water burns with a pop sound indicating that the gas is hydrogen.



$Fe(s) + 2HCl(g) \longrightarrow FeCl_2(s) + H_2(g)$

b) **Reaction with potassium permanganate ($KMnO_4$)**

When hydrogen chloride gas is slowly passed over potassium permanganate, a greenish yellow gas is formed. The greenish yellow gas is chlorine.

$2KMnO_4(s) + 16HCl(g) \longrightarrow 2KCl(aq) + 2MnCl_2(aq) + 8H_2O(l) + 5Cl_2(g)$

From the above two reactions, conclusions can be made that the components of hydrogen chloride are chlorine and hydrogen.

2. **Reaction with ammonia**

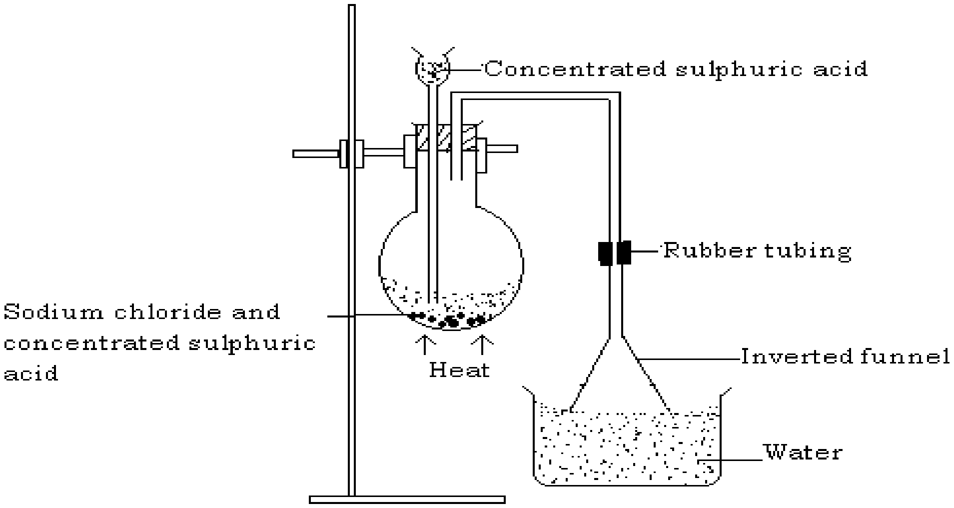
Hydrogen chloride reacts with ammonia forming white fumes of ammonium chloride.

$NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s)$

Preparation of a solution of hydrogen chloride in water (hydrochloric acid)

The solution of the gas in water forms hydrochloric acid. When the gas is bubbled through water until no more of the gas can dissolve, then the product is concentrated hydrochloric acid and contains 36% by mass of hydrogen chloride.

When dissolving, the apparatus should be arranged as below. The method is only suitable if the gas is very soluble.



The rim of the funnel must just be at the surface of water in the beaker this avoids water being sucked into the preparation apparatus; a delivery tube cannot also be used because it would suck up water in to the preparation apparatus.

Properties of hydrochloric acid

It has all the properties typical of an acid like

- Sour taste
- Turns blue litmus paper red
- It is fully ionized in aqueous solution indicating that it is a strong acid

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

$H_2O(l) + HCl(aq) \longrightarrow H_3O^+(aq) + Cl^-(aq)$

- It liberates carbon dioxide with carbonates

$ZnCO_3(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + CO_2(g) + H_2O(l)$

Ionically

$CO_3^{2-}(aq) + 2H^+(aq) \longrightarrow CO_2(g) + H_2O(l)$

- It forms hydrogen with metals above hydrogen in the electrochemical series.

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

- It neutralizes bases to form salts and water

$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$

Ionically

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

$CuO(s) + 2HCl(aq) \longrightarrow CuCl_2(aq) + H_2O(l)$

- Oxidizing agents e.g. Manganese (IV) oxide and potassium permanganate oxidize the concentrated hydrochloric acid to chlorine.

$2KMnO_4(s) + 16HCl(aq) \longrightarrow 2KCl(aq) + 2MnCl_2(aq) + 8H_2O(l) + 5Cl_2(g)$

These properties are due to the fact that hydrogen chloride although a covalent compound ionizes completely when dissolved in water and therefore show typical acidic characteristics.

Uses of hydrochloric acid

- It is used in the removal of rust from iron (descaling) before it is galvanized.
- It is used in cleaning metals before they are electroplated.
- It is used in the manufacture of plastics like polyvinylchloride
- Used in the preparation of soluble chlorides

Preparation of solution of hydrogen chloride in methyl benzene

The preparation is done the same way as above. Hydrogen chloride dissolves but does not ionize in organic solvents like methyl benzene and exists as a covalent compound. Therefore, the solution does not conduct electricity; has no effect on litmus paper; does not react with metals above copper and does not react with carbonates or hydrogen carbonates but reacts with ammonia to form white precipitates of ammonium chloride since the salt is insoluble in organic solvents.

Test for chlorides

Solid chlorides

When concentrated sulphuric acid is added to any chloride, hydrogen chloride gas is evolved.

The gas forms fumes in moist air; turns moist blue litmus paper red; and forms thick white fumes of ammonium chloride with ammonia; and forms white precipitates of silver chloride with acidified silver nitrate solution.

Chloride in solution

To a solution of chloride, add a little nitric acid and then silver nitrate solution. White precipitates of silver chloride are seen.

$$Ag^{+}(aq) + Cl^{-}(aq) \longrightarrow AgCl(s)$$

When a little ammonia solution is added to the above solution, the white precipitates dissolve. The precipitates are insoluble in nitric acid. The only two insoluble chlorides are lead(II) chloride and silver chloride.

Sample questions on Chlorine and its compound

1. *Describe and explain the laboratory preparation and manufacture of chlorine. Using equations, describe the reactions of chlorine with: metals, non metals, water, dilute acids and hydrocarbons. How would you show experimentally that hydrogen chloride is very soluble in water? Outline the uses of chlorine.*
2. *With the aid of a labeled drawing, explain how hydrogen chloride and hydrochloric acid are prepared in the laboratory. Describe experiments to deduce the composition of hydrogen chloride. Describe the reactions of hydrogen chloride with ammonia. Explain the behaviour of hydrogen chloride in (i) water and (ii) methyl benzene.*
3. *Describe one experiment in each case to show (a) how hydrochloric acid acts as a reducing agent (b) how hydrogen chloride can be identified. Give the uses of hydrochloric acid.*

SULPHUR AND ITS COMPOUNDS

Sulphur is in period 3 and group VI of the periodic table. It has atomic number number 16 and electronic configuration 2.8.6.

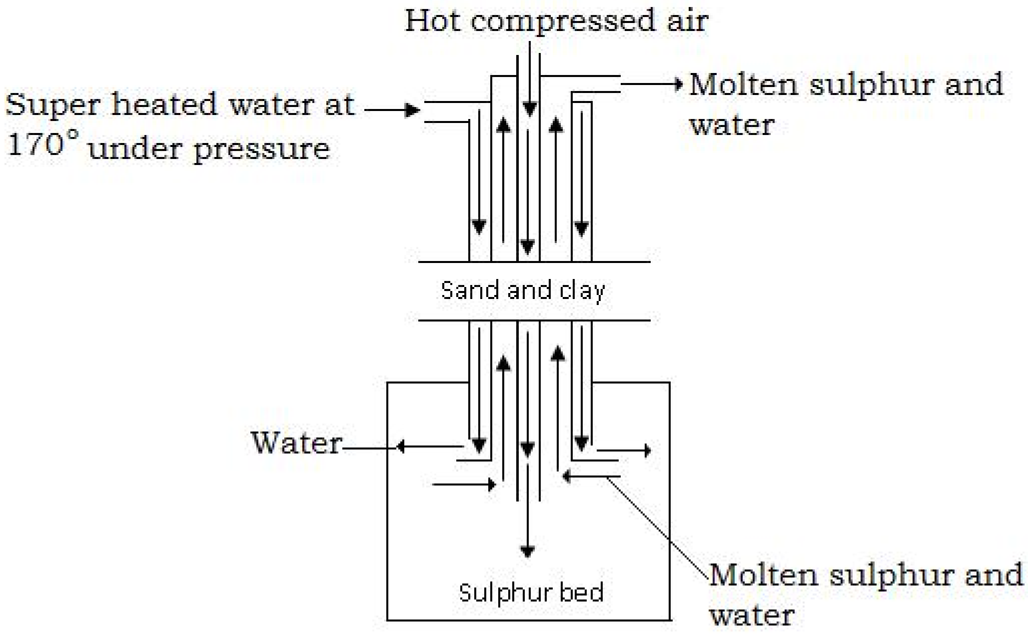
Occurrence

It occurs in both Free State and combined state. In the free state it occurs in underground deposits and it is widely distributed in volcanic regions. In combined states it occurs as sulphates, sulphides (mainly hydrogen sulphide) and sulphite. It also occurs in crude oil.

Extraction of sulpur

Sulphur is found deep below the ground (160-200) metres, this makes mining it impossible. It is extracted by a method invented by a man called Frasch and the method is **Frasch process**. Sulphur is extracted from the underground deposits by this method basing on its low melting point.

The Frasch process



Procedure

1. Drill the Frasch pump consisting of three concentric pipes down the sulphur deposit.
2. Super heated water (steam) at about 170˚C is then forced down the outer most tube to melt the sulphur.
3. A jet of hot compressed air is then pumped down through the inner most tube. This hot compressed air pumps the molten sulphur out through the middle tube to the surface where the sulphur is cooled and solidified. The sulphur obtained is

about 99.5% pure and can be used directly. It is usually sold in two forms either as "flowers of sulphur" a powder or "roll sulphur", cylindrical sticks.

Sulphur from natural gas and petroleum

Natural gas obtained during the distillation of petroleum contains hydrogen sulphide. The hydrogen sulphide can be removed by dissolving it in a suitable solvent. The gas is removed from the solvent and one third (1/3) of the hydrogen sulphide obtained is burnt in oxygen to form sulphurdioxide.

$$2H_2S(g) + 3O_2(g) \longrightarrow 2H_2O(l) + 2SO_2(g)$$

The sulphurdioxide reacts with the two third (2/3) of the hydrogen sulphide left to form sulphur.

$$SO_2(g) + H_2S(g) \longrightarrow S(s) + H_2O(l)$$

The water is evaporated to leave sulphur solids.

Physical properties of sulphur

1. It is a yellow solid at room temperature
2. It is a typical non metal
3. It is insoluble in water but soluble in some organic solvents e.g. methyl benzene and carbon disulphide

Allotropes of sulphur

Sulphur has two major crystalline forms .i.e. allotropes namely

1. Rhombic sulphur (alpha sulphur, α-sulphur)

Rhombic is an octahedral crystal; its bright yellow in color; its melting point is 114˚C; it has a density of 2.8g/cm^3 and it is stable at a temperature below 96˚C.

Formation of rhombic sulphur

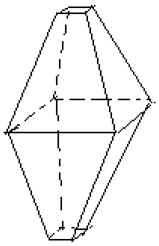
Shake some powdered sulphur with carbon disulphide for some time in a test tube. (Take care to extinguish all flames from the vicinity).

Filter the content of the test tube into a dry beaker through a dry filter paper and funnel over the mouth of the beaker; fasten a filter paper over the mouth of the beaker and pierce a few pin holes in it, and set the beaker aside.

The carbon disulphide will slowly evaporate depositing crystals of rhombic sulphur which because of the slow evaporation will be large enough for their shape to be seen.

NB The formation of the crystal takes place at room temperature.

Structure



2. Monoclinic sulphur (Beta sulphur,β-sulphur)

It is a needle shaped crystal (prismatic); it is very pale yellow in color (almost transparent); it has a melting point of 119˚C; it has a density of 1.98 g/cm^3; it is stable above 96˚C, below 96˚C it reverts to rhombic sulphur.

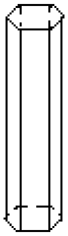
Formation of monoclinic sulphur

Place powdered sulphur in a very large crucible. Heat it and stir gradually while adding some sulphur until the crucible is full of molten sulphur. Use a small flame for heating otherwise the sulphur will burn.

Allow the sulphur to cool. After some time, a solid crust begins to form on the surface.

When the crust has formed make two holes (at wide separation) using a glass rod on it and pour out the molten sulphur from inside which has not crystallized.

Remove the crust and observe needle shaped crystals of monoclinic sulphur formed beneath the crust and through out the inner surface of the crucible.



Transition temperature is a temperature at which rhombic sulphur changes to monoclinic sulphur and vice versa. At a temperature below 96˚C, rhombic sulphur exists and as the temperature goes above 96˚C, the rhombic sulphur changes to monoclinic form. The transition temperature is there fore 96˚C.

To show that both rhombic and monoclinic sulphur are allotropes of sulphur

When the same mass of either rhombic or monoclinic sulphur is burnt in oxygen, the same mass of sulphur dioxide is obtained in each case and nothing else.

Other forms of sulphur

Amorphous sulphur

This is a non crystalline form of sulphur which is insoluble in carbon disulphide(CS_2). It is formed as an almost insoluble powder if a saturated solution of hydrogen sulphide is oxidized by leaving it to stand in open air for some times i.e.

$$2H_2S(aq) + \underset{\text{(From air)}}{O_2(g)} \longrightarrow \underset{\text{(Amorhous sulphur)}}{2S(s)} + 2H_2O(l)$$

Plastic sulphur

If sulphur at its boiling point is poured into cold water, a dark-sticky (elastic) solid called plastic sulphur is formed. Plastic sulphur is elastic because it contains zig-zag (entangled) chains of S_8 molecules.

N.B Plastic sulphur is unstable and slowly hardens to form yellow rhombic sulphur. Plastic sulphur is not a separate allotrope of sulphur since it is not crystalline in nature.

Action of heat on sulphur

Sulpur under goes a series of changes when it is heated. Both rhombic and monoclinic sulphur consist of S_8 molecules with different arrangements. It is this differences in arrangement of sulphur atoms that is responsible for the different observations made when sulphur is heated.

1. If yellow powdered sulphur is heated in the absence of air just above the melting point (about 115˚C), it melts into a clear amber (pale yellow) mobile liquid. This liquid is mobile because the sulphur-8 (S_8) molecule rings can flow over one another with ease.
2. On further heating (to about 160˚C), the yellow liquid becomes darker and very viscous. This is because the S_8 rings are broken and form long chains of sulphur 8 (S_8) atoms. The liquid is viscous because the long chains entangle with one another and thus do not flow readily over each other. Above 160˚C the darker viscous sulphur liquid becomes mobile and reddish brown in color.
3. Near its boiling point (444˚C), the liquid now becomes lighter in color, thin and more mobile (less viscous). This is because the long entangled chains break down forming chains of S_1 and S_2 atoms which can flow more readily.
4. The sulphur eventually boils at 444˚C and forms a light brown sulphur vapor.

Other properties of sulphur

1. It is a reactive element and it combines directly with other elements (metals and non metals)

a) If a piece of burning sulphur is lowered into a gas jar of oxygen, it continues to burn even more brightly with a blue flame forming whites fumes with a choking smell. The white fumes with a choking smell are a mixture of white sulphur trioxide and colorless sulphur dioxide gases.

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

$$2S(s) + 3O_2(g) \longrightarrow 2SO_3(g)$$

b) If mixture of iron dust (iron fillings) and powdered sulphur is heated in a hard glass test tube, the two elements combine vigorously and a spontaneous red glow spreads through the mixture since the reaction is exothermic. A dark grey (black) hard mass of iron(II)sulphide is formed.

$$Fe(s) + S(s) \longrightarrow FeS(s) + heat$$

Almost all sulphides are black.

c) Zinc reacts more vigorously with sulphur if the mixture is heated forming zinc sulphide.

$$Zn(s) + S(s) \longrightarrow ZnS(s)$$

d) Hot copper foil (wire) also glows in sulphur vapor forming copper(I) sulphide.

$$2Cu(s) + S(g) \longrightarrow Cu_2S(s)$$

e) Carbon combines directly with sulphur to form a liquid, carbon disulphide. Very high temperatures are required for this reaction to occur.

$$C(s) + 2S(s) \longrightarrow CS_2(l)$$

f) Reaction with acids

Dilute acids have no effects on sulphur. However,

a) Hot concentrated sulphuric acid oxidizes sulphur to sulphur dioxide and the sulphuric acid its self is also reduced to sulphur dioxide.

$$S(s) + 2H_2SO_4(aq) \longrightarrow 3SO_2(g) + 2H_2O(l)$$

b) Hot concentrated nitric acid oxidises sulphur to sulphuric acid and brown fumes of nitrogen dioxide is given off.

$$6HNO_3(aq) + S(s) \longrightarrow 6NO_2(s) + H_2SO_4(aq) + 2H_2O(l)$$

Uses of sulphur

1. It is used in the making of matches, gun powder and fire works.
2. Used in the vulcanization (hardening) of rubber
3. Used in the manufacture of sulphuric acid in the contact process.
4. Used as a fungicide and in medicine, ointments and drugs used for the treatment of skin diseases. e.g. Sulphonamide.
5. Used in the production of calcium hydrogen sulphite, $Ca(HSO_3)_2$ that acts as a bleaching agent in the wood pulp in manufacture of paper.
6. Sulphur is used in the manufacture of various compounds like carbon disulphide (CS_2) and sodium thiosulphide $Na_2S_2O_3$ used in photography.
7. Fruit trees are sprayed with sulphur products like carbon disulphide (CS_2) to kill insects and fungi which cause diseases.

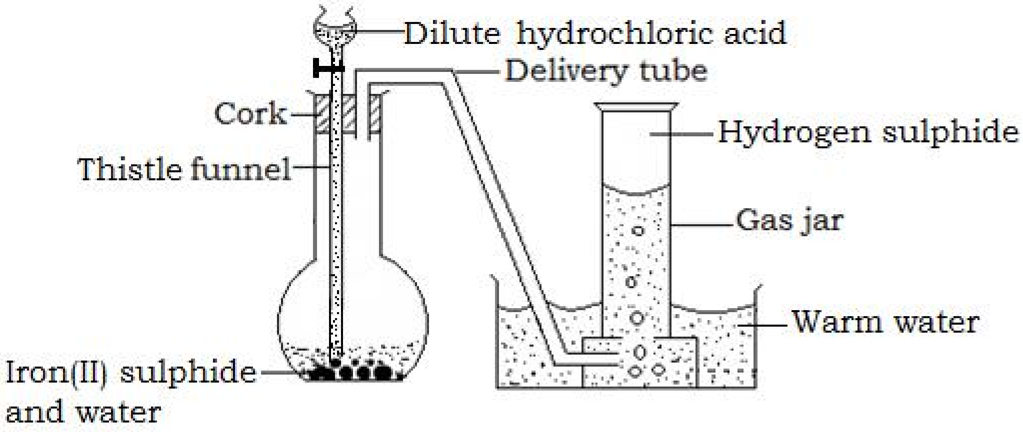
COMPOUNDS OF SULPHUR

HYDROGEN SULPHIDE (H_2S)

Laboratory preparation

It can be prepared in a laboratory by the action of dilute hydrochloric acid or dilute sulphuric acid on iron(II) suphide. The preparation must be done in a fume cupboard as hydrogen sulphide gas is very poisonous.

Set up of apparatus



As the acid reaches the iron (II)sulphide, effervescence begins and the hydrogen sulphide is collected over warm water since in is soluble in cold water. If it is required dry, the gas is passed over anhydrous calcium chloride and then collected by downward delivery method.

Equation

$FeS(s) + 2HCl(aq) \longrightarrow H_2S(g) + FeCl_2(aq)$

Dilute sulphuric acid may also be used

$FeS(s) + H_2SO_4(aq) \longrightarrow H_2S(g) + FeSO_4(aq)$

NB

1. Concentrated sulphuric acid is never used to dry the hydrogen sulphide as it reacts with the gas.
   $(3H_2S(g) + H_2SO_4(aq) \longrightarrow 4S(s) + 4H_2O(l))$
2. The hydrogen sulphide prepared in this case contains some impurities especially hydrogen chloride gas and hydrogen.

Physical properties

i) It has a strong repulsive characteristic of a rotten egg smell
ii) It is a colorless gas
iii) It is very poisonous but not as dangerous as carbon monoxide
iv) It is slightly denser than air that is why it is collected by downward delivery
v) It can be liquefied under high pressure

vi) It dissolves in cold water forming a fairly weak acidic solution

$H_2S(s) + (aq) \longrightarrow H_2S(aq)$

The weak acid formed dissociates forming ions

$H_2S(aq) \longrightarrow H^+(aq) + HS^-(aq)$

Test for hydrogen sulphide

a) Hydrogen sulphide can easily be detected by its strong repulsive smell of rotten eggs.

b) When hydrogen sulphide is passed through a solution of Lead(II)nitrate or Lead(II)ethanoate, a black precipitate of lead(II)sulphide is observed.

$$Pb^{2+}(aq) + S^{2-}(aq) \longrightarrow PbS(s)$$

i.e.

With Lead(II)nitrate

$$Pb(NO_3)_2(aq) + H_2S(g) \longrightarrow PbS(s) + 2HNO_3(aq)$$

With Lead(II)ethanoate

$$(CH_3COO)_2Pb(aq) + H_2S(g) \longrightarrow PbS(s) + 2CH_3COOH(aq)$$

Chemical properties of hydrogen sulphide

1. Hydrogen sulphide burns in excess air (oxygen) with a pale blue flame forming suphur dioxide and water vapor.

$2H_2S(g) + 3O_2(g) \longrightarrow 2SO_2(g) + 2H_2O(g)$

However, in limited supply of air, a yellow deposit of sulphur is formed.

$2H_2S(g) + O_2(g) \longrightarrow 2S(s) + 2H_2O(g)$

2. Hydrogen sulphide is a powerful reducing agent and it gives up its hydrogen to many compounds which are therefore reduced and hydrogen sulphide itself is oxidized to sulphur. For example

a) When sulphur dioxide is bubbled through a solution of hydrogen sulphide, a yellow deposit of sulphur appears and water is formed.

$$2H_2S(aq) + SO_2(g) \longrightarrow 3S(s) + 2H_2O(l)$$

b) Hydrogen sulphide reacts with halogens to form hydrogen halides and yellow deposits of sulphur e.g. when hydrogen sulphide is mixed with chlorine gas, a pale yellow deposit of sulphur and white fumes of hydrogen chloride gas are produced.

$$H_2S(g) + Cl_2(g) \longrightarrow S(s) + 2HCl(g)$$

When hydrogen sulphide is bubbled through bromine water, the reddish brown bromine water is decolorized and pale yellow precipitate of sulphur appears.

$$H_2S(g) + Br_2(l) \longrightarrow S(s) + 2HBr(aq)$$

c) When hydrogen sulphide is bubbled through a solution of iron(III)chloride, the solution changes from pale yellow to pale green. This is because hydrogen sulphide reduces iron(III)chloride (the pale yellow solution) to iron(II)chloride (a pale green solution). The hydrogen sulphide itself is oxidized to sulphur which appears as a yellow deposit. Hydrogen chloride gas is also formed which dissolves to form hydrochloric acid.

$$2FeCl_3(aq) + H_2S(g) \longrightarrow 2FeCl_2(aq) + S(s) + 2HCl(aq)$$

Or

$$2Fe^{3+}(aq) + S^{2-}(aq) \longrightarrow 2Fe^{2+}(aq) + S(s)$$

d) When hydrogen sulphide is bubbled through a solution of acidified potassium manganate(VII), the solution changes color from purple to colorless and a yellow deposit of sulphur is also formed. This is because, hydrogen sulphide reduces manganate(VII) ions (purple in color) to manganese(II) ions(colorless) and the hydrogen sulphide itself is oxidized to sulphur.

$$2KMnO_4(aq) + 6H_2SO_4(aq) + 5H_2S(g) \longrightarrow K_2SO_4(aq) + 2MnSO_4(aq) + 5S(s) + 8H_2O(l)$$

e) When hydrogen sulphide is bubbled through a solution of acidified potassium dichromate(VI) the solution changes from orange to green as a result of reduction of the the dichromate(VI) ions to chromium(III) ions.

$$2K_2Cr_2O_7(aq) + 8H_2SO_4(aq) + 3H_2S(g) \longrightarrow 2Cr_2(SO_4)_3(aq) + 2K_2SO_4(aq) + 5S(s) + 7H_2O(l)$$

f) Hydrogen sulphide reduces concentrated sulphuric acid according to the following reaction

$$3H_2S(g) + H_2SO_4(aq) \longrightarrow 3S(s) + 4H_2O(l)$$

g) When hydrogen sulphide is bubbled through concentrated nitric acid, brown fumes of nitrogen dioxide together with a pale yellow precipitate of sulphur are observed.

$$H_2S(g) + 2HNO_3(aq) \longrightarrow 2NO_2(g) + S(s) + 2H_2O(l)$$

3. Hydrogen sulphide gas precipitates insoluble sulphides e.g. it precipitates black copper(II)sulphide from blue copper(II)sulphate solution.

$$CuSO_4(aq) + H_2S(g) \longrightarrow CuS(s) + H_2SO_4(aq)$$

OXIDES OF SULPHUR

There are two principle oxides of sulphur namely sulphur dioxide and sulphur trioxide.

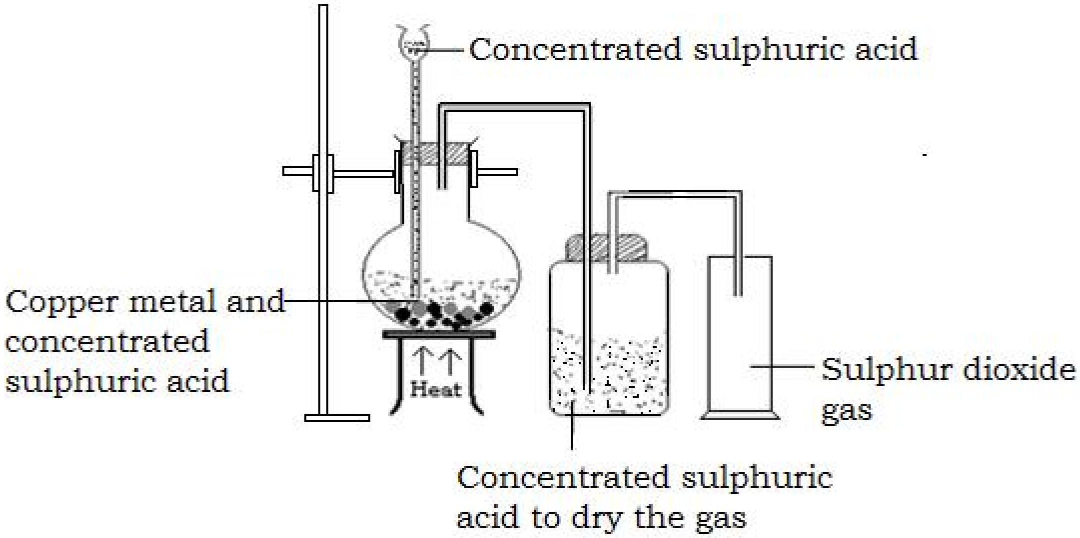
SULPHUR DIOXIDE, SULPHUR (IV) OXIDE, SO_2

Laboratory preparation

Sulphur dioxide in the laboratory can be prepared in two ways

a) Action of concentrated sulphuric acid on copper metal
b) By action of dilute sulphuric acid or hydrochloric acid on any sulphite salt e.g sodium sulphite(Na_2SO_3).

Preparation by the action of concentrated H_2SO_4 on copper metal



Place copper metal in the flask and arrange the apparatus as shown above.

Gently heat the mixture until when it is hot. Effervescence occurs as sulphur dioxide is evolved.

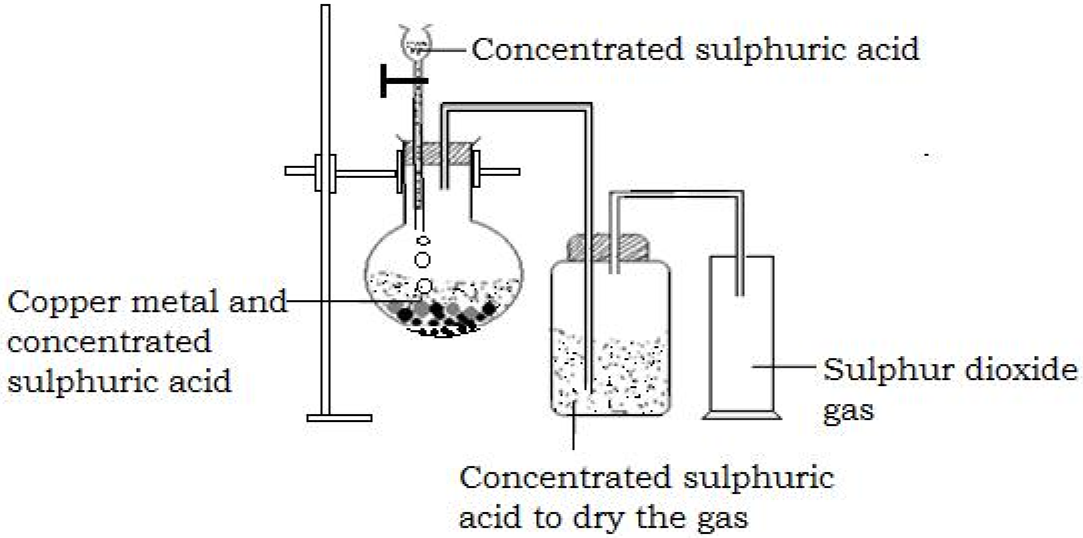
Equation

$Cu(s) + H_2SO_4(aq) \longrightarrow CuSO_4(aq) + SO_2(g) + H_2O(l)$

The gas is passed through a wash bottle containing concentrated sulphuric acid to dry the gas and it is the collected by downward delivery since it is denser than air. The gas is not collected over water as it is very soluble in water.

Laboratory preparation of sulphur dioxide from sodium sulphite and dilute sulphuric or hydrochloric acid

The procedure and the arrangement of apparatus remains as shown above but no heating is required for the reaction. However, if the reaction slows down, the flask may be heated gently.



Equation

$Na_2SO_3(s) + 2HCl(aq) \longrightarrow SO_2(g) + 2NaCl(aq) + H_2O(l)$

Sodium hydrogen sulphite can also be used under the same conditions.

$NaHSO_3(s) + HCl(aq) \longrightarrow SO_2(g) + 2NaCl(aq) + H_2O(l)$

Physical properties of sulphur dioxide

1. It is a poisonous gas
2. It is a colorless gas with a characteristic pungent smell
3. It is denser than air
4. It can easily be liquefied under pressure
5. It is an acidic gas i.e. it turns moist blue litmus paper red
6. It is very soluble in water forming sulphurous acid

Chemical properties of sulphurdioxide

1. Reaction with alkalis

Sulphur dioxide is neutralized by alkalis

i) When the alkali is in excess sulphites are formed.

$2NaOH(aq) + SO_2(g) \longrightarrow Na_2SO_3(aq) + H_2O(l)$

ii) When the sulphur dioxide is in excess, hydrogen sulphites are formed and water is not a product in this case.

$NaOH(aq) + SO_2(g) \longrightarrow NaHSO_3(aq)$

2. As a reducing agent

In the presence of water, sulphur dioxide behaves as a reducing agent

a) When sulphur dioxide is bubbled through acidified potassium dichromate(VI) solution, the solution changes from orange to green. This is because, sulphur dioxide reduces chromium (VI) ions to chromium (III) ions.

$K_2Cr_2O_7(aq) + H_2SO_4(aq) + 3SO_2(g) \longrightarrow Cr_2(SO_4)_3(aq) + K_2SO_4(aq) + H_2O(l)$

NB This is a characteristic test for sulphurdioxide

b) Sulphur dioxide also reduces acidified potassium Manganate(VII) to manganese (II) sulphate. The color changes from purple to colorless and the sulphur dioxide is itself oxidized to sulphuric acid.

$2Mn\ {}_4^-(aq) + 2H_2O(aq) + 5SO_2(g) \longrightarrow 5SO_4^{2-}(aq) + 2Mn^{2+}(aq) + 4H^+(aq)$

This is also used as a test for sulphur dioxide.

c) When sulphur dioxide is bubbled through a solution of iron(III)suphate, the color changes from yellow to pale green. This is because the sulphur dioxide reduces iron (III) sulphate to iron (II) sulphate and the sulphur dioxide is oxidized to sulphuric acid.

$Fe_2(SO_4)_3(aq) + 5SO_2(g) + 2H_2O(l) \longrightarrow 2FeSO_4(aq) + 2H_2S0_4(aq)$

d) Sulphur dioxide reduces concentrated nitric acid to form brown fumes of nitrogen dioxide and itself is oxidized to sulphuric acid.

$SO_2(g) + 2HNO_3(aq) \longrightarrow 2NO_2(g) + H_2SO_4(aq)$

e) Colored solutions of halogen are made colorless when sulphur dioxide is bubbled through them. This is because sulphur dioxide reduces the halogens to hydrogen halides. For example

When sulphur dioxide is bubbled through chlorine water (yellowish green liquid), a mixture of hydrochloric acid and sulphuric acid are obtained which appear colorless.

$$H_2O(l) + SO_2(g) + Cl_2(g) \longrightarrow 2HCl(aq) + H_2SO_4(aq)$$

Bromine water (a reddish brown liquid) is turned colorless as bromine is reduced to hydrobromic acid by sulphur dioxide.

$$H_2O(l) + SO_2(g) + Br_2(g) \longrightarrow 2HBr(aq) + H_2SO_4(aq)$$

3. As an oxidizing agent

Sulphur dioxide acts as an oxidizing agent when it reacts with reducing agents more powerful than itself. Consider the reactions below

a) When sulphurdioxide is bubbled through a solution of hydrogen sulphide, a yellow precipitate is observed. This is because sulphur dioxide oxidizes hydrogen sulphide to yellow sulphur and sulphur dioxide is itself reduced to sulphur.

$$SO_2(g) + 2H_2S(aq) \longrightarrow 3S(s) + 2H_2O(l)$$

b) When a piece of magnesium ribbon is lowered into a gas jar of sulphur dioxide, it continuous to burn with a bright flame to form white solids (magnesium oxide) and a yellow solid(sulphur). This is because sulphur dioxide oxidizes magnesium to magnesium oxide and itself is reduced to sulphur.

$$2Mg(s) + SO_2(g) \longrightarrow 2MgO(s) + S(s)$$

4. As a bleaching agent

Sulphur dioxide bleaches wet flowers like roses and hibiscus and any other wet material by **reduction**. It does this by removing oxygen from the colored material.

$\boxed{\text{Dye} + H_2}$

Complex which
appears colourless

$Dye + SO_2(g) + 2H_2O(l) \longrightarrow \qquad + H_2SO_4(aq)$

The original color of the bleached material may be restored after prolong exposure to air due to **aerial oxidation**. This explains why old news papers appear yellow after some times.

5. Reaction with oxygen

Under normal conditions, sulphur dioxide does not react with oxygen. However when a mixture of dry and pure sulphur dioxide and oxygen is passed over heated platinum catalyst (or vanadium(V)oxide), sulphur trioxide gas is formed.

$2SO_2(g)\ O_2(g) \longrightarrow 2SO_3(g)$

6. As an acid anhydride

Sulphur dioxide is an acid anhydride, reacting with water to form sulphurous acid.

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

Sulphurous acid is a dibasic acid and reacts with sodium hydroxide solution to form two salts of sodium, the acid salt;sodium hydrogen sulphite ($NaHSO_3$)and the normal salt;sodium sulphite (Na_2SO_3)

$NaOH(aq) + H_2SO_3(aq) \longrightarrow NaHSO_3(aq) + H_2O(l)$

$2NaOH(aq) + H_2SO_3(aq) \longrightarrow Na_2SO_3(aq) + 2H_2O(l)$

Uses of sulphur dioxide

1. It is used in the manufacture of sulphuric acid by contact process.
2. It is a poisonous gas and it is used in fumigation to kill germs in clothes and houses.
3. It is used for preservation of food staff and fruits during transportation and storage, as it prevents fermentation.

4. It is used as a bleaching agent e.g it is used to make calcium hydrogen sulphite ($Ca(HSO_3)_2$) that makes wood pulp white in paper manufacture, used to bleach silk, straw e.t.c.

Sulphites (SO_3^{2-})

These are salts derived from sulphurous acid.

Test for sulphite

Procedure

To a solution of substance containing sulphite, add 3 drops of barium nitrate solution followed by excess nitric acid. (or add barium chloride solution followed by excess hydrochloric acid)

Observation

White precipitates immediately appear which dissolve with effervescence on adding dilute nitric acid.

$$SO_3^{2-}(aq) + Ba^{2+}(aq) \longrightarrow BaSO_3(s)$$

On addition of nitric acid

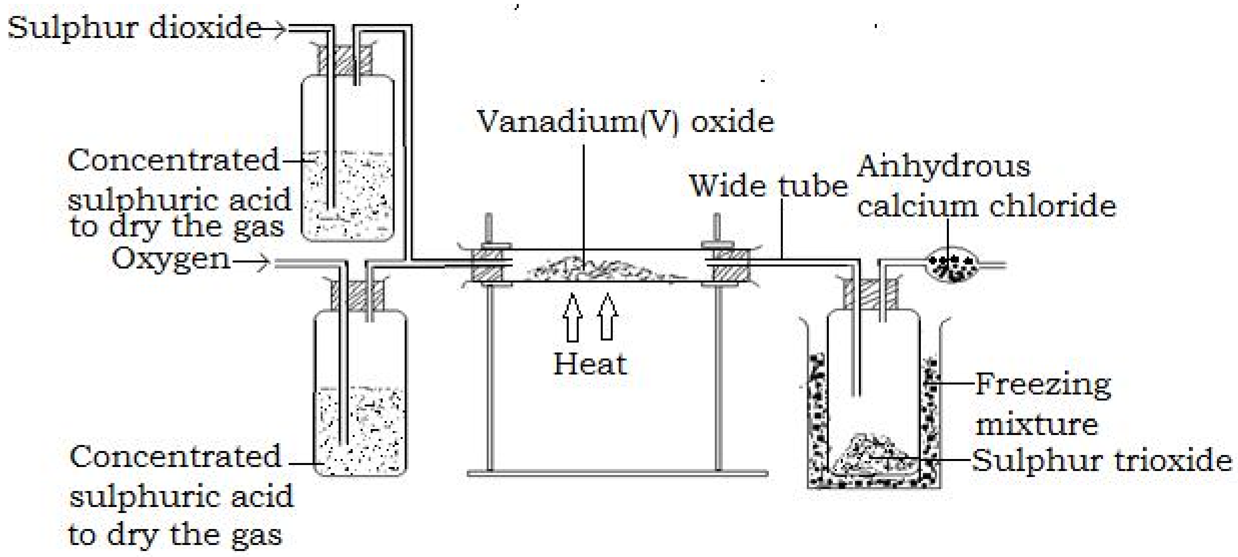
$$BaSO_3(s) + HNO_3(aq) \longrightarrow Ba(NO_3)_2(aq) + SO_2(g) + H_2O(l)$$

SULPHUR TRIOXIDE (SO_3)

Laboratory preparation

It can be prepared in the laboratory by passing a dry mixture of oxygen and sulphur dioxide over a heated platinum catalyst or (Platinized asbestos) at a temperature of 450-500˚C.

Set up of apparatus



Equation for the reaction

$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$

The sulphur trioxide is seen as dense white fumes and may be solidified in a freezing mixture of ice and a little sodium chloride. The sulphur trioxide container is protected from atmospheric moisture by calcium chloride tube.

Properties of sulphur trioxide

1. It has very high affinity for water and combines with it violently forming sulphuric acid.

$$SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq)$$

This reaction is highly exothermic i.e. it gives out a lot of heat and it is because of this that sulphur trioxide is kept in air tight containers.

2. Sulphur trioxide combines with concentrated sulphuric acid to form fuming sulphuric acid called oleum.

$$SO_3(g) + H_2SO_4(aq) \longrightarrow H_2S_2O_7(l)$$

SULPHURIC ACID (H_2SO_4)

Large scale (Industrial) manufactureof sulphuric acid by contact process

In the manufacture of sulphuric acid by contact process, sulphur dioxide and oxygen are the starting materials.

The sulphur dioxide is oxidized to sulphur trioxide which is then absorbed by concentrated sulphuric acid forming oleum (fuming sulphuric acid) to which water is added to form the sulphuric acid.

This process can be divided into the following essential stages;

a) Preparation of sulpur dioxide

Sulpur dioxide can be obtained from the following source;

i) Burning sulpur in air. This is cheap and produces sulphur dioxide in large quantities

$$S_{(s)} + O_{2(g)} \longrightarrow SO_{2(g)}$$

ii) Roasting sulphide ores in air

$$4FeS_{2(s)} + 11O_{2(g)} \longrightarrow 8SO_{2(g)} + 2Fe_2O_{3(s)}$$

(Iron pyrite)

$$2ZnS_{(s)} + 3O_{2(g)} \longrightarrow 2SO_{2(g)} + 2ZnO_{(s)}$$

(Zinc blende)

Other sources of sulphur dioxide include; burning of hydrogen sulphide from crude oil in air; flue gas desulphurization in power stations e.t.c.

Oxygen is obtained from fractional distillation of liquid air.

b) Purification of the gases

The sulphur dioxide and the oxygen are purified and dried (i.e. cleared off any dust particles and other impurities which can poison the catalyst especially if it is platinum.

c) Preparation of sulphur trioxide

The purified gases are passed over a finely divided vanadium (V) oxide, (V_2O_5) catalyst at a temperature of 450-500°C and a pressure of 2-3 atmospheres, sulphur trioxide is formed. Vanadium (V) oxide is commonly used because it is cheaper and not easily poisoned by impurities.

$$2SO_{2\ (g)} + O_{2\ (g)} \underset{\substack{\text{Temp } 450\text{-}500\ ^0C \\ \text{Pressure, 2-3 atm}}}{\overset{\text{Catalyst } (V_2O_5)}{\rightleftharpoons}} 2SO_{3\ (g)}$$

The catalyst Vanadium (V) oxide is so effective that 95% conversion of sulphur dioxide to sulphur trioxide is achieved at 450-500°C and 2 atmospheres. The reaction is exothermic and there fore produces heat enough to maintain the temperature of the catalyst.

d) Conversion of sulphur trioxide to sulphuric acid

Sulphur trioxide,SO_3 must not be allowed to come in contact with water as the reaction is intensely exothermic that it vaporizes the sulphuric acid formed (i.e. produces a lot of mist consisting of dry droplets of H_2SO_4).

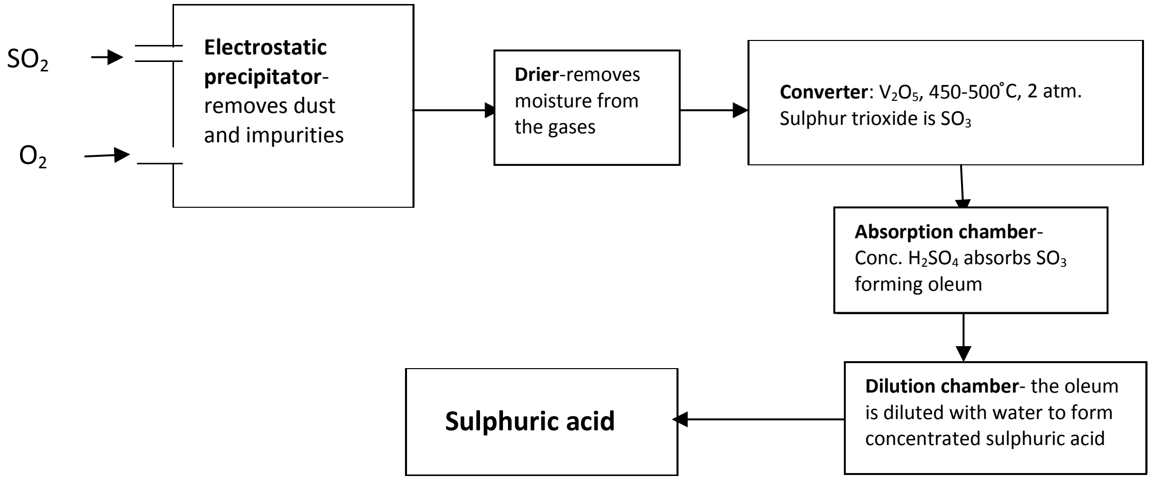
To prevent this happening, the sulphur trioxide, SO_3 is absorbed in concentrated sulphuric acid, H_2SO_4 to form an oily liquid called an Oleum

$$H_2SO_4(aq) + SO_3(aq) \longrightarrow H_2S_2O_7(l)$$

The oleum produced is carefully diluted to give 95-98% pure concentrated sulphuric acid.

$$H_2S_2O_7(l) + H_2O(l) \longrightarrow 2H_2SO_4(aq)$$

Summary of the contact process



In the above process, the following conditions favor high yield of sulphur trioxide:

- Presence of a catalyst. The catalyst must be finely divided to increase the surface area for the reaction.
- Low temperature(450-500°C) as the reaction is exothermic (releases heat),
- Slightly high pressure above the atmospheric pressure as the reaction is accompanied by a decrease in volume.
- High concentration of oxygen or sulphur dioxide.

Properties of sulphuric acid

a) Physical properties

1. Concentrated Sulphuric acid is a dense, colorless and oily liquid.
2. Concentrated sulphuric acid has very high affinity for water and a lot of heat is produced when the acid is diluted.
3. Concentrated sulphuric acid is hygroscopic. I.e. gradually absorbs moisture from the air and therefore when left exposed to air, in a beaker, the total volume gradually increases due to absorption of water. This is why it is used as a drying agent for many of the gases.

b) Chemical properties

1. Sulphuric acid as an acid

It is the dilute sulphuric acid that reacts as a typical acid.

i. Dilute sulphuric acid ionizes to form hydrogen ions

$H_2SO_4(aq) \longrightarrow H^+(aq) + SO_4^{2-}(aq)$

ii. Sulphuric acid liberates hydrogen gas from reactive metals e.g. Na, Mg, and Ca.

$Ca(s) + H_2SO_4(aq) \longrightarrow Ca(OH)_2(aq) + H_2(g)$

iii. It reacts with bases to form salt and water only. Since it is a dibasic acid, it produces two types of salts, the normal salt (sulphate) when alkali is in excess and the acidic salt (hydrogen sulphate) when the acid is in excess.

$H_2SO_4(aq) + NaOH(aq) \longrightarrow NaHSO_4(aq) + 2H_2O(l)$

$H_2SO_4(aq) + 2NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$

iv. It reacts with carbonates and hydrogen carbonates to liberate carbondioxide gas

$Na_2CO_3(aq) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + H_2O(l) + CO_2(g)$

$Ca(HCO_3)_2(aq) + H_2SO_4(aq) \longrightarrow CaSO_4(aq) + H_2O(l)\ CO_2(g)$

NB With calcium carbonate and lead (II) carbonate, there is little effervescence and the reaction stops soon because the sulphates formed are insoluble and form a coating around the carbonate preventing any further attack by sulphuric acid on the carbonate.

2. As an oxidizing agent

Hot concentrated sulphuric acid is a powerful oxidizing agent and in all reactions, it is itself reduced to sulphur dioxide. Metals are oxidized to sulphates and non metals to their oxides. Example include:

a) Hot concentrated sulphuric acid oxidizes copper to copper (II) sulphate and the sulphuric acid itself is reduced to sulphur dioxide.

$2H_2SO_4(l) + Cu(s) \longrightarrow CuSO_4(aq) + SO_2(g) + 2H_2O(l)$

b) When charcoal is heated with concentrated sulphuric acid, the charcoal is oxidized to carbon dioxide and the sulphuric acid is reduced to sulphur dioxide.

$2H_2SO_4(l) + C(s) \longrightarrow CO_2(aq) + 2SO_2(g) + 2H_2O(l)$

c) Hot concentrated sulphuric acid oxidizes sulphur to sulphur dioxide and the sulphuric acid itself is reduced to sulphur dioxide also.

$$2H_2SO_4(l) + S(s) \longrightarrow 3SO_2(g) + 2H_2O(l)$$

d) When a sample of hydrogen sulphide gas is bubbled through concentrated sulphuric acid, the hydrogen sulphide is oxidized to sulphur and the sulphuric acid is reduced to sulphur dioxide.

$$H_2S + H_2SO_4(l) \longrightarrow S(s) + SO_2(g) + 2H_2O(l)$$

3. As a dehydrating agent

Concentrated sulphuric acid is a very strong dehydrating agent i.e. it removes the elements of water from many compounds. Examples

a) When concentrated sulphuric acid is added to blue copper (II) sulphate crystals, the copper(II)sulphate crystals gradually become white as their water of crystallization is lost.

$$CuSO_4.5H_2O(s) \xrightarrow{\text{Concentrated } H_2SO_4} CuSO_4(s) + 5H_2O(g)$$

b) When cold concentrated sulphuric acid is added to sugar crystals in an evaporating dish, the sugar crystals turn progressively from white, to yellow then to brown and finally to black. A spongy black mass of charcoal rises almost filling up the dish , water vapour is given off and the dish becomes very hot as the reaction is exothermic (generates energy).

$$C_{12}H_{22}O_{11}(s) \xrightarrow{\text{Concentrated } H_2SO_4} 12C(s) + 11H_2O(g) + \text{Heat}$$

In this reaction ,the concentrated acid removes the elements of water from sugar leaving a black mass of carbon.

A similar action is the explanation of a very marked corrosive action of the acid on flesh and cloth made of cotton. Cotton is largely cellulose whose simplest formula is $(C_6H_{10}O_5)n$.

c) Concentrated sulphuric acid dehydrates oxalic acid on heating to form a mixture of carbon monoxide, carbon dioxide and water.

$$H_2C_2O_4(s) \xrightarrow{\text{Concentrated } H_2SO_4} CO(g) + CO_2(g) + H_2O(l)$$

d) Excess concentrated sulphuric acid dehydrates ethanol at a temperature of 170˚C forming ethene.

$$C_2H_5OH(l) \xrightarrow{\text{Concentrated } H_2SO_4\text{,170˚C}} C_2H_4(g) + H_2O(l)$$

When concentrated sulphuric acid removes elements of water from a compound with the formation of a new compound, it is described as a **dehydrating agent**.

4. As a drying agent

When concentrated sulphuric acid removes water from a mixture, it acts as a **drying agent**.

Concentrated sulphuric acid reacts exothermically with water. When a solution is made, it is essential to pour the acid into water, stirring to disperse the heat evolved. It is dangerous to add water to concentrated sulphuric acid as small pockets of water are likely to boil.

Gases are dried by bubbling them through concentrated sulphuric acid. For basic gases like ammonia, another drying agent is used.

Uses of sulphuric acid

1. Used in the manufacture of fertilizers like ammonium sulphate.
2. Making of paints and pigments
3. Manufacture of detergents and soap
4. Production of other chemicals such as metallic sulphates, hydrochloric acid, hydrofluoric acid and plastics.
5. Extraction of metals and metal manufacturing including pickling to clean metallic surfaces.
6. Extraction of alkenes in petroleum refinery.
7. With nitric acid, it is used to make dyes and explosives.

Sulphates

These salts are derived from sulphuric acid.

All sulphates are soluble in water except barium sulphate, lead(II)sulphate and calcium sulphate is slightly soluble in water.

Action of heat on sulphates

Most of the sulphates are resistant to heat, but if they are hydrated, they lose their water of crystallization and become powdery upon slight heating. E.g.

$$Na_2SO_4.10H_2O(s) \longrightarrow Na_2SO_4(s) + 10H_2O(l)$$

However, iron (II) sulphate, copper (II) sulphate, ammonium sulphate and sulphates of other metals lower than copper in the reactivity series are decomposed upon strong heating. For example

i. When a green hydrated solid of iron (II) sulphate is heated gently, it loses its water of crystallization which condenses on the cooler part of the test tube forming dirty yellow anhydrous solids of iron (II) sulphate.

$$FeSO_4.7H_2O(s) \longrightarrow FeSO_4(s) + 7H_2O(l)$$

On further heating, the anhydrous dirty yellow solids decompose giving off sulphur dioxide (which turn orange potassium dichromate green) in addition to white fumes of sulphur trioxide and leave a brown solid of iron (III) oxide.

$$2FeSO_4(s) \longrightarrow Fe_2O_3(s) + SO_3(g) + SO_2(g)$$

Overall equation: $2FeSO_4.7H_2O(s) \longrightarrow Fe_2O_3(s) + 14H_2O\ (l) + SO_3(g) + SO_2(g)$

ii. When a blue copper(II)sulphate crystal is heated, it losses its water of crystallization forming white anhydrous powder of copper(II)sulphate.

$$CuSO_4.5H_2O(s) \longrightarrow CuSO_4(s) + 5H_2O(l)$$

On strong heating, the white anhydrous white solid decomposes into white fumes of sulphur trioxide, a colorless gas (sulphur dioxide) and a black solid residue (copper(II)oxide).

$$2CuSO_4(s) \longrightarrow 2CuO(s) + SO_3(g) + SO_2(g)$$

Chemical test for sulphates

To the solution of the suspected sulphate in water, add barium chloride and dilute hydrochloric acid (barium nitrate solution and dilute nitric acid can also be used.) a white precipitate which is insoluble in excess acid indicates the presence of a sulphate.

$$BaCl_2\,(aq) + Na_2SO_4\,(aq) \longrightarrow BaSO_4(s) + 2NaCl\,(aq)$$

White precipitate
(Insoluble barium
sulphate)

Ionically,

Ionically,

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$$

Sulphur and its compounds

1. *Describe fully the changes which take place when powdered sulphur is gradually heated to its boiling point in the absence of oxygen. Mention briefly the proof that rhombic and monoclinic sulphur are allotropes.*
2. *Describe the large scale extraction of sulphur. Describe the preparation and collection of gas jars of hydrogen sulphide starting from powdered sulphur. Describe an experiment to show the action of this gas as (a) a combustible substance (b) a reducing agent (c) an acid forming salt.*
3. *What is the effect of heat on ferrous sulphate crystals? Make a drawing to show how you can convert sulphur dioxide to sulphur trioxide. What is the action of sulphur dioxide on (a) nitric acid (b) hydrogen sulphide (c) chlorine.*
4. *Explain how sulphuric acid is manufactured on a large scale. Outline four large scale uses of the acid. Describe one experiment in each case to show how sulphuric acid acts as (a) an oxidizing agent (b) a dehydrating agent (c) a bleaching agent*

5. Give atleast five chemical properties of concentrated sulphuric acid. Why is this acid said to be dibasic? How can sulphates and sulphites be identified in the laboratory.
6. Describe the preparation and collection of hydrogen sulphide. Give four chemical properties of the gas. What changes occur when the gas is bubbled through solutions of (a) copper sulphate (b) lead nitrate (c) chlorine. Write equations for the reactions.
7. Make a labeled drawing of apparatus you would use to prepare sulphur dioxide. How does this gas react with (a) chlorine water (b) hydrogen sulphide (c) sodium hydroxide solution? How can sulphur dioxide be converted to sulphuric acid?
8. Starting with roll sulphur, how can rhombic sulphur crystals, monoclinic sulphur crystals and plastic sulphur are obtained.

ENERGY CHANGES IN CHEMICAL REACATIONS (THERMO CHEMISTRY)

All matter posses energy in one form or another. Energy can neither be created nor destroyed but only transformed from one form to another. Some substances have energy due to movement of the particles and this is referred to as **kinetic energy.** The energy that a substance possesses due to its position or due to the arrangement of its component parts is the **potential energy.** Due to the fact that energy can be present in many forms makes it difficult to determine the magnitude of the total energy possessed by a substance, it is easier to determine the energy changes which take place when the substance interacts with other substances.

Energy change

Energy change in a reaction is the energy given out or taken in by a substance during a reaction. Energy changes can be recognized in many different forms such as:

- Light e.g. when magnesium burns in air
- Sound e.g. explosion of hydrogen and oxygen
- Electricity e.g. in cell reaction (electrochemical cells)
- Heat e.g. burning of fuel like charcoal and wood.

In some cases a reaction may involve a mixture of several energy forms, for example the reactions of potassium with water.

In thermo chemistry, we mainly deal with heat change (enthalpy change) as it is the most common form of energy change.

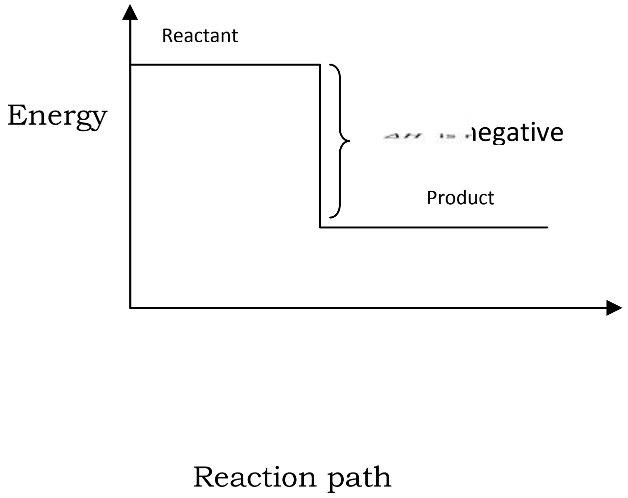
Heat (enthalpy) change

Many reactions are accompanied by marked heat changes. There are basically two types of heat changes or enthalpy changes.

1. An exothermic reaction

Is a reaction in which heat is liberated to the surrounding. It is indicated by a rise in temperature. In such reactions, the products have less energy content than the reactants.

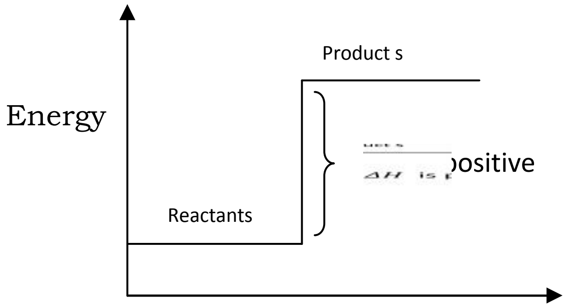
Energy profile for exothermic reactions



2. An endothermic reaction

Is a reaction in which heat is absorbed from the surrounding. In this reaction, the products are richer in energy than the reactants.

Energy profile for endothermic reactions



Reaction path

Units for heat changes: KJ/mol ($KJmol^{-1}$)

Heat of reaction and ΔH notation

This is the heat change which takes place when substances react to form one mole of the product.

The heat content of a reacting system is denoted by H and we normally consider the changes in heat content-delta H, (ΔH) since it is not easy to precisely determine H.

$\Delta H = H_{(product)} - H_{(reactants)}$

If heat is lost by the reacting system (in exothermic reaction), the ΔH is negative. For example;

$C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -394KJ/mol$

This means that when 1 mole of carbon in solid state is completely burnt in gaseous oxygen, it forms I mole of gaseous carbon dioxide with evolution of 394 KJ of energy. (394 KJ are evolved when 12g of carbon combines with 32g of oxygen)

If heat is taken in by the reacting system (in endothermic reactions), the ΔH is positive. For example;

$C(s) + 2S(s) \longrightarrow CS_2(l) \quad \Delta H = +106\ KJ/mol$

This means that when 1 mole of carbon reacts with 2 moles of sulphur, it forms I mole of carbon disulphide with absorption of 106 KJ of energy. (106 KJ are absorbed when 12g of carbon combines with 64g of sulphur).

Sources of heat changes

The source of heat change in a chemical reaction is the potential energy which substances posses. Bonds (electrovalent, covalent e.t.c.) between atoms or ions represent a form of stored potential energy and a source of heat changes in chemical reactions. It is always necessary to provide energy to break bonds in chemical compounds and in the same way, energy is evolved when a new bond is created from the constituent atoms.

Bond breaking requires supply of heat from the surrounding (endothermic process) while bond formation liberates heat to the surrounding (exothermic process).

Types of heat (enthalpy) changes

Enthalpy of combustion

This is the heat change when one mole of a substance is completely burnt in oxygen. All combustion reactions are exothermic hence heat is always evolved and ΔH value is always negative. For example;

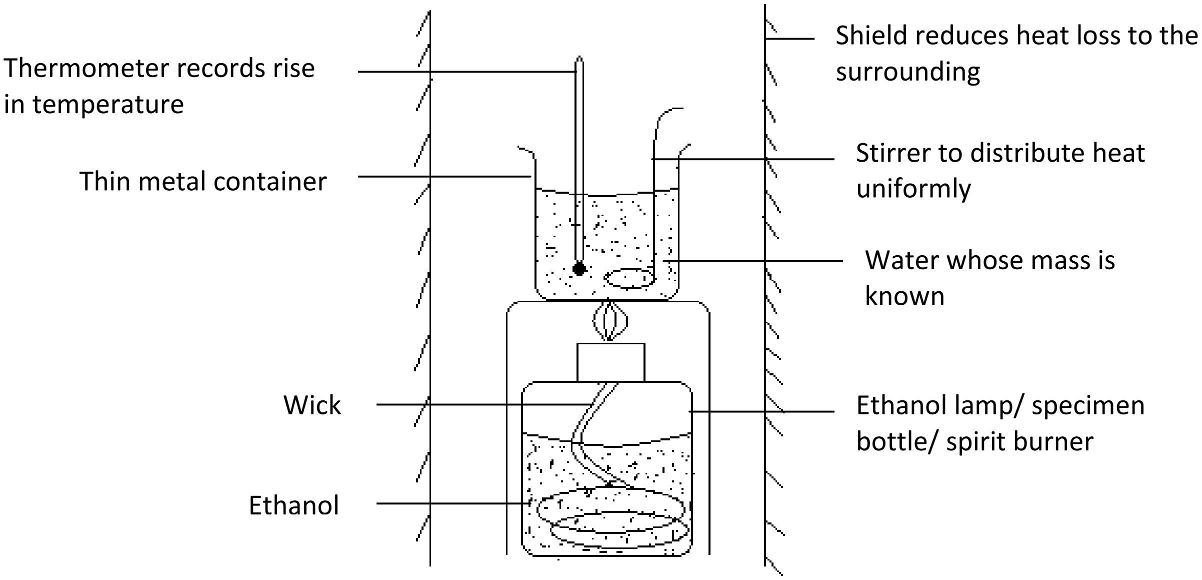
$C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -394 KJ/mol$

$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) \quad \Delta H = -1560 KJ/mol$

Expt. To determine the heat of combustion of ethanol

An ethanol lamp with a wick inserted in it is used.

Set up of apparatus



Procedure

- A known mass of ethanol is put in the ethanol lamp and the mass of ethanol lamp and ethanol is weighed, M_1 g.
- A known mass of water is M_w is put is a thin metal container.
- The initial temperature of the water, T_1 is recorded.
- Light the wick and stir the water constantly until when the temperature of the water rises by about 30˚C. Extinguish the flame and record the temperature T_2 of the water.
- Reweigh the ethanol lamp and its content and record the mass M_2 g.

In this experiment,

1. A wire gauge is not used as it would absorb some of the heat.
2. The shield reduces heat loss to the surrounding.
3. It is assumed that all the heat produced by the burning ethanol is absorbed by the water. This is actually not true as some of the heat is usually lost to the surrounding and the container holding the water. The values obtained are usually lower than the actual values.

Treatment

Initial mass of ethanol lamp and ethanol before burning =M_1 g

Mass of ethanol lamp and ethanol after burning =M_2 g

Mass of ethanol burnt =(M_1-M_2)g

Mass of water = M_w g

Temperature rise of water = (T_2-T_1)

Specific Heat Capacity of water (Cw) = 4.2J/g/K

Heat gained by water= mass of water x Cw x temperature change

$$= MwCw(T_2-T_1)$$

Moles of ethanol burnt= $M_1 - M_2/_{Mr}$,where Mr is the molar mass of ethanol

$M_1 - M_2/_{Mr}$ moles of ethanol evolves $MwCw(T_2-T_1)$

1 mole of ethanol evolves $\frac{MwCw(T2-T1)}{M_1-M_2}$ x Mr

Thus, the heat of combustion of ethanol is $\frac{MwCw(T2-T1)}{M_1-M_2}$ x Mr in J/mol

Ethanol is an example of fuel. The following are qualities of a good fuel.

1. It must produce a lot of heat.
2. It must be cheap and must burn easily.
3. It should not burn with much smoke (pollutant gas)
4. Is should be easily transported with little or no fire risks.
5. It should be easily safely stored.

Example

1. Use the following specimen results to calculate the heat oif combustion of ethanol.

Initial mass of ethanol lamp and ethanol before burning =29.974 g
Mass of ethanol lamp and ethanol after burning =29.592 g
Initial temperature of water = 17.7˚C
Final temperature of water = 41.2˚C
Volume of water in can= 100 cm^3

(Given, density of water= 1g/cm^3; specific heat capacity of water, Cw= 4.2J/g/K; C=12, H=1,O=16)

Solution

Heat transferred to water by burnt ethanol =mass of water x Cw x temperature change
= $MwCw(T_2-T_1)$
but Mw= density of water x volume of water, (1 x 100) =100g
Therefore,

Heat liberated =100 x 4.2 x (41.2-17.7) J
=9870J

Mass of ethanol burnt=(29.974 - 29.592)g = 0.382g
Molar mass of C_2H_5OH =12x2+1x5+16x1+1 = 46g

0.382g of ethanol liberates 9870J of heat
46 g of ethanol liberates ($\frac{46 \text{ x}9870}{0.382}$)J of heat

=1188534.031 J

=1188.5 KJ

The molar heat of combustion is -1189 KJ/Mol

2. Ethanol burns in excess air according to the equation

$2CH_3OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(g)$ $\Delta H = -730 KJ/mol$

Given that 3.2g of methanol were burnt, determine

i) The amount of heat liberated

ii) The mass of carbon dioxide produced

Solution

i) From the equation,

2 moles of CH_3OH liberate 730 KJ

2(12+1x3+16+1)g of CH_3OH liberate 730 KJ

64g of CH_3OH liberate 730 KJ

1g of CH_3OH liberate $({}^{730}/_{64})$ KJ

3.2g of CH_3OH liberate $({}^{730x3.2}/_{64})$ KJ

=32.5KJ

ii) From the equation,

2 moles of CH_3OH liberate 2 moles of CO_2

2(12+1x3+16+1)g of CH_3OH liberate 2(12+16x2)g of CO_2

64g of CH_3OH liberate 88g of CO_2

1g of CH_3OH liberate $({}^{88}/_{64})$g of CO_2

3.2g of CH_3OH liberate $({}^{3.2x88}/_{64})$g of CO_2

=4.4g of CO_2

3. The enthalpy for the complete combustion of carbon is -393KJ/Mol

a) Write an equation for the complete combustion of carbon

b) 80 kg of charcoal costs 4000/=. Calculate the cost of charcoal required to produce 16375 KJ.

Solution

a) $C(s) + O_2(g) \longrightarrow CO_2(g)$ ΔH= -393KJ/mol

b) From the equation

393 KJ of heat is liberated by 12g of Carbon

1 KJ of heat is liberated by $(^{12}/_{393})$g of Carbon

16375 KJ of heat is liberated by $(^{12x16375}/_{393})$g of Carbon

=500g or 0.5 Kg

80 kg of charcoal (carbon) costs 4000/=

1 kg of charcoal (carbon) costs $(^{4000}/_{80})$/=

0.5 kg of charcoal (carbon) costs $(^{4000x0.5}/_{80})$/=

=25/=

Heat of solution

Heat of solution is the heat change which takes place when one mole of a substance is dissolved in a specified amount of solvent. The enthalpy change can either be negative if the reaction is exothermic of positive if it is endothermic.

Expt. To determine the heat of solution of sodium hydroxide

Procedure

- Place a known volume of water, V cm^3 in a plastic beaker.
- Stir well with a thermometer and record the temperature of the water, T_1
- Carefully add a known mass, M of sodium hydroxide pellets in the water.
- Stir well until there is no change in temperature; record the final temperature T_2 of the solution.

Treatment of results

Volume of water used = V cm^3

Mass of sodium hydroxide = M g

Initial temperature = T_1

Final temperature = T_2

(Given, density of water=1g/$cm^{3;}$ ^{specific} heat capacity of solution (Cs)=4.2J/g/˚C)

Mass of water= density x volume, =1xV =Vg

Mass of solution = (M+V)g

Temperature change =(T_2-T_1)

Heat liberated by solution =mass of solution x Cs x ∆temperature

=(M+V) x Cs x (T_2-T_1) J

M g of sodium hydroxide liberates (M+V) x Cs x (T_2-T_1) J

1g of sodium hydroxide liberates ($^{(M+V) \text{ x Cs x } (T2-T1)}/_M$) J

Molar mass of NaOH =23+16+1 =40

40 g of sodium hydroxide liberates ($^{(M+V)\text{x Cs x } (T2-T1) \text{ x } 40}/_M$) J

Then, the heat of solution = ($^{(M+V)\text{x Cs x } (T2-T1) \text{ x } 40}/_M$) J/mol

Or ($^{(M+V)\text{x Cs x } (T2-T1) \text{ x } 40}/_{1000M}$) KJ/mol

Example

1. When 10g of ammonium chloride was dissolved in 100 cm^3 of water, the temperature changed from 21˚C to 19˚C. Determine the molar heat of solution of ammonium chloride (SHC of solution=4.2J/g/˚C; density of water= 1 g/cm^3; N=14, H=1, Cl=35.5)

Solution

Heat change of solution = mass of solution x SHC of solution x Temperature change

Mass of solution = mass of water + mass of ammonium chloride

=(100 x 1) + 10, = 110g

Temperature change = 21-19 =2˚C

$\Delta H_{Solution}$ = (110 x 4.2 x2) J

= 924 J

Molar mass of NH_4Cl= 14+1x4+35.5 = 53.5g

10 g of NH_4Cl liberates 924J

1 g of NH_4Cl liberates ($^{924}/_{10}$)J

53.5 g of NH_4Cl liberates ($^{53.5x924}/_{10}$)J

=4943.4 J

Or 4.9434 KJ

The molar heat of solution of ammonium chloride is 4.9434 KJ/mol.

2. When 16g of anhydrous copper (II) sulphate were dissolved in 150 cm^3 of water, the temperature changes from 24˚C to 29˚C. Determine the heat of solution of the anhydrous copper (II) sulphate and with a reason, state whether the reaction is exothermic or endothermic.

Solution

Heat change of solution = mass of solution x SHC of solution x Temperature change

Mass of solution = mass of water + mass of copper(II)sulphate

=(150 x 1) + 16, = 166g

Temperature change = 29-24 =5˚C

$\Delta H_{Solution}$ = (166 x 4.2 x5) J

= 3486 J

Molar mass of $CuSO_4$= 64+32+16x4= 160g

16 g of $CuSO_4$ liberates 3486 J

1 g of $CuSO_4$ liberates ($^{3486}/_{16}$)J

160 g of $CuSO_4$ liberates ($^{3486x160}/_{16}$)J

=34860 J

Or 34.860KJ

The molar heat of solution of copper(II)sulphate is 34.860KJ/mol.

The reaction is endothermic as it was accompanied by heat absorption and rise in the temperature of the product.

Enthalpy of neutralization

This is the heat change that takes place when an acid reacts with a base to produce one mole of water. The heat change is as a result of the reaction of hydrogen ions of the acid and hydroxyl ions of the base.

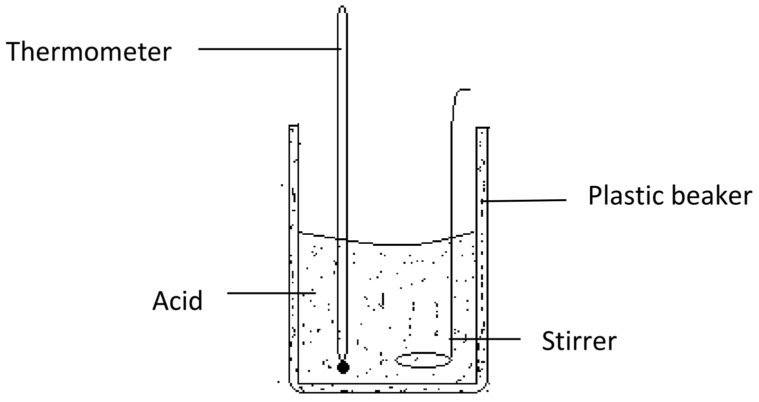
$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l) \quad \Delta H_{Neutralization}$$

For strong acid and strong alkalis, the heat of neutralization is almost constant. This is because, they are fully ionized in their aqueous solutions and the solution consists of only ions, therefore when the solutions are mixed, there is no bond breaking as the ions are already existing freely. E.g. a solution of hydrochloric acid consists of only H^+ and Cl^- ions and sodium hydroxide solution consists of only Na^+ and OH^- ions, when the two solutions are mixed, there is no bond breaking as the ions are already separated.

Weak acids and weak bases are not fully ionized and their heats of neutralization are not constant. This is because some of the energy is first used in the dissociation of the molecules of the weak acid/base before neutralization can take place. Hence the value is usually less than that of strong acids and bases.

Expt. To determine heat of neutralization of an acid and a base

Set up



Procedure

- Put a known volume of acid V_A of known concentration M_A in a plastic beaker
- Stir and record the temperature T_A of the acid.
- Put in another plastic container a known volume V_B of the alkali of known concentration M_B.
- Stir and record the temperature T_B of the alkali
- Quickly mix the two solutions, stir well using a thermometer and record the maximum temperature T_2 of the mixture.

Treatment of results

Initial temperature of acid = T_A

Initial temperature of base =T_B

Initial average temperature = $T_A - T_B/_2$ = T_1

Maximum temperature of the mixture = T_2

Temperature change = T_2-T_1

Mass of solution= (V_A+V_B) x density of solution. Assuming density=1g/cm^3

=(V_A+V_B) g

Heat change = mass of solution x SHC of solution x Temperature change

=(V_A+V_B)xSHCx(T_2-T_1) J

Moles of acid neutralized = $M_AV_A/_{1000}$

$M_AV_A/_{1000}$ moles of acid evolves (V_A+V_B)xSHCx(T_2-T_1) J

1 mole of acid evolves $(VA + VB)xSHCx(T2 - T1)\ \ 1000/_{M_AV_A}$J

The heat of neutralization of the acid is $(VA + VB)xSHCx(T2 - T1)\ \ 1000/_{M_AV_A}$ J/mol

Example

1. $50cm^3$ of 2M HCl and 50 cm^3 of 2M NaOH both at 22˚C where mixed in a plastic beaker. The mixture was stirred and the maximum temperature contained was 35˚C.(SHC of solution=4.2J/g/˚C; Density of solution =1 g/cm^3)
   a) Write the ionic equation for the recation
   b) Calculate the heat change for the reaction
   c) Calculate the heat change when 1 mole of the acid is neutralized by 1 mole of the base i.e. molar enthalpy of neutralization.
   d) In another experiment, 50 cm^3 of 2M NH_4OH was used instead of the NaOH. State whether the heat of reaction was greater than, equal to or less than the value you calculated.

Solution

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

b) Heat change = mass of solution x SHC of solution x Temperature change

= (50+50)x1 x 4.2 x (35-22) J

=(100x4.2x13) J

=5460J

c) Number of moles in $50cm^3$ of 2M HCl

$1000cm^3$ of solution contains 2 moles of HCl

$1cm^3$ of solution contains $({}^{2}/{}_{1000})$ moles of HCl

$50cm^3$ of solution contains $({}^{2x50}/{}_{1000})$ moles of HCl

= 0.01 moles of HCl

0.01 moles of HCl produces 5.46KJ

1mole of HCl produces $({}^{5.46}/{}_{0.01})$ KJ

=546 KJ

d) Ammonium hydroxide is a weak base and not completely ionized, its heat of neutralization is less than the calculated value.

Exercise

1. When $50cm^3$ of 0.5M HCl acid solution was added to $50cm^3$ of 0.5M KOH solution in a calorimeter, there was a temperature rise from 27.5˚C to 30.8˚C. Calculate the

enthalpy of neutralization of the reaction. (density of water=$1g/cm^3$ and its heat capacity is 4.2 J/g/˚C)

2. Use the following results to calculate the heat of neutralization of an acid.
   (Assume density of solution =$1g/cm^3$) (Answer=54.6KJ/mol)
   Volume of 2M NaOH = $20cm^3$
   Volume of 2M HCl = $20cm^3$
   Initial temperature of HCl= 14.0˚C
   Initial temperature of NaOH= 14.4˚C
   Final temperature of the mixture = 27.2˚C

Enthalpy of displacement

Elements that are more reactive displace those that are less reactive than them during chemical reactions. For example zinc displaces copper from its compound say copper(II)sulphate. During such reactions, heat is normally liberated.

$Zn(s) + CuSO_4(aq) \longrightarrow Cu(s) + ZnSO_4(aq)$

Expt. To determine the heat of displacement of reaction between copper (II) sulphate solution and zinc

Procedure

- Put a known volume of copper(II)sulphate solution V_1 of known concentration M_1 in a plastic beaker.
- Stir and record the temperature T_1 of the solution.
- Add a known mass of zinc powder to the copper (II) sulphate solution, stir the mixture and record the maximum temperature of the mixture, T_2.

Treatment of results

Volume of copper (II) sulphate solution = V_1

Concentration of copper (II) sulphate solution = M_1

Initial temperature =T_1

Final temperature = T_2

Temperature change = T_2-T_1

Mass of solution= V_1 x density of solution. Assuming density=1g/cm^3

= V_1 g

Heat change = mass of solution x SHC of solution x Temperature change

= V_1x SHCx (T_2-T_1) J

Moles of copper (II) sulphate that reacted = $M_1V_1/_{1000}$

From the equation of reaction

$Zn(s) + CuSO_4(aq) \longrightarrow Cu(s) + ZnSO_4(aq)$

The moles of zinc that reacted= moles of copper (II) sulphate = $M_1V_1/_{1000}$

$M_1V_1/_{1000}$ moles of zinc evolves V_1xSHCx(T_2-T_1) J

1 mole of zinc evolves $\mathrm{V1xSHCx(T2 - T1)\ 1000}/_{M_1V_1}$ J

The heat of displacement is $\mathrm{V1xSHCx(T2 - T1)\ 1000}/_{M_1V_1}$ J/mol

Example

1. When 6.5g of zinc powder where added to 250cm3 of 0.5 M copper(II)sulphate solution in a plastic cup, 5.46 KJ of heat was liberated.

a) Why was a plastic cup used and not a metallic cup?

b) Write the ionic equation for the reaction between zinc powder and copper (II) sulphate solution.

c) Calculate

i) The number of moles of zinc in 6.5g of zinc powder

ii) The number of moles of zinc that reacted with copper (II) sulphate solution

iii) The heat energy produced when 1 mole of zinc reacts with 1 mole of copper (II) sulphate (enthalpy of displacement)

(Given Cu=65)

Solution

a) To minimize heat loss and gain to and from the surrounding

b) $Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(aq)$

c) i) 65g of Cu is contained in 1 mole

1g of Cu is contained in $(^{1}/_{65})$ moles

6.5g of Cu is contained in $(^{1x6.5}/_{65})$ moles

=0.125 moles of $CuSO_4$

ii) 1000cm³ of solution contains 0.5 moles of $CuSO_4$

1cm³ of solution contains $(^{0.5}/_{1000})$ moles of $CuSO_4$

250cm³ of solution contains $(^{0.5x250}/_{1000})$ moles of $CuSO_4$

=0.125 moles of $CuSO_4$

From the equation of reaction, $Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(aq)$

1mole of $CuSO_4$ reacts with 1mole of Zn

0.125 moles of $CuSO_4$ reacts with $(^{0.125}/_{1})$ mole of Zn

= 0.125 moles of Zn

iii) 0.125 moles of Zn evolves 5.46 KJ

1 mole of Zn evolves $(^{5.46}/_{0.125})$ KJ

=43.6 KJ/mol

Heat of formation

This is the heat change when one mole of a substance is formed from 1 mole its elements in their standard states. For example

$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(l) \qquad \Delta H_{Formation} = -286$ KJ/mol

286JK of heat is evolved when one mole of water is formed from I mole of gaseous hydrogen and ½ mole of gaseous oxygen.

Exercise

1. From the equation $4Al(s) + 30_2(g) \longrightarrow 2Al_20_3(s)$ $\Delta H_{Formation}$ = -3360 KJ/mol
   a) Calculate the enthalpy of formation of aluminium oxide (answer -1680 KJ/mol)
   b) Determine the amount of heat produced when 10.2g of aluminium oxide is formed from its elements in their standard states.
   c) What volume of oxygen will be required for the formation of aluminium oxide at room temperature? (Al=27. O=16, 1 mole of a gas at stp occupies 24000cm^3)

Heat changes during physical processes

Melting

The particles in a solid are packed close to each other and do not move from one place to another. When a solid is heated, the particles gain more kinetic energy and vibrate more vigorously. Eventually they are able to leave their positions in the structure and move past one another; they flow as a liquid. At this point, the melting point of the solid is reached. The heat supplied to melt the solid causes negligible rise in temperature and therefore referred to as '**latent**' heat (literally meaning hidden heat). This is because, the heat supplied at first simply weakens the strong intermolecular forces of attraction holding the particles together. Melting is an endothermic process as it involves heat absorption.

The **enthalpy of melting** or **enthalpy of fusion** of a substance is the enthalpy change when one mole of a substance changes from solid to liquid state (i.e. when it melts)

The **enthalpy of freezing** of a substance is the enthalpy change per mole when the substance changes from liquid to solid state (i.e. when is freezes)

These changes do not take place under standard conditions. Enthalpy changes of freezing and melting are quoted at 1 atmosphere and at the temperature at which the change of state takes place. The enthalpy of freezing is the negative of the enthalpy of melting.

$\Delta H_{freezing} = -\Delta H_{melting}$

Vaporization

The particles in a liquid are in motion and the forces of attraction keep them in the liquid phase. When a liquid is heated, the average kinetic energy of the molecules increases. Some molecules with energy above average, are able to break away from the attraction of other molecules and enter the vapor phase. Vaporization is endothermic.

The **enthalpy of vaporization** is the enthalpy change that takes place when one mole of substance changes state from liquid to vapor.

$\Delta H_{vaporisation}$ = Enthalpy of vapor - Enthalpy of liquid

Condensation

This is the reverse of vaporization and it is exothermic.

$\Delta H_{condensation} = -\Delta H_{vaporisation}$

Sample questions on Electrochemistry

1. *Explain what is meant by exothermic and endothermic reactions. Use energy profile diagrams to illustrate.*
2. *Define the following terms: heat of combustion, heat of solution, het of neutralization, enthalpy of displacement, and heat of formation. Describe an experiment to determine the enthalpy of combustion of ethanol.*
3. *Explain why the enthalpy of neutralization of sodium hydroxide by hydrochloric acid is the same as the heat of neutralization of potassium hydroxide by nitric acid i.e.. $-57.3KJmol^{-1}$*
4. *An important gaseous fuel is butane C_4H_{10}. For the combustion of butane, $\Delta H = -2880KJmol^{-1}$. (i) write down the equation for butane burning in excess air, including the heat change. (ii) calculate the quantity of heat which would be evolved when $16dm^3$ of butane measured at rm*

temperature and pressure are burnt. (the molar volume of a gas at may be taken as 24dm3 at r.t.p.) (**Ans. = 1920KJ**)

5. When hydrogen chloride gas dissolves in water, there is an appreciable rise in temperature. What type of reaction is suggested and how do you account for the temperature change? Hydrogen chloride dissolves in tri chloromethane without any apparent change in temperature. What does this suggest.
6. Write an ionic equation for the neutralization of potassium hydroxide with sulphuric acid. When $100cm^3$ of a 0.25M sulphuric acid was added to $100cm^3$ of a 0.5M potassium hydroxide solution, the temperature of the solution rose from 25.6°C to 28.9°C. Calculate the enthalpy of neutralization of potassium hydroxide with sulphuric acid. (Density of water = $1g/cm^3$; the specific heat capacity of water = $4.2Jg^{-1}$ $^oC^{1}$) (**Ans.=** $\mathbf{-55.44KJmol^{-1}}$).Ammonia solution was neutralized with sulphuric acid. Comment on the numerical value of the enthalpy of neutralization compared to that in (b); and give a reason for your answer.
7. In an experiment to determine the enthalpy of combustion of ethanol, $100cm^3$ of water was measured into a thin metal can. The can was heated at the bottom by a known amount of burning ethanol. Using the specimen results below, calculate the molar enthalpy of combustion of ethanol

   Results:

   Initial temperature = 19.82°C

   Final temperature of water = 40.30°C

   Mass of ethanol + metal can before burning = 26.88g

   Mass of ethanol + metal can after burning = 26.56g

   (Assume; the specific heat capacity of water = 4.2J/g/°C, density of water = 1gcm-3, C=12, O=16, H=1) (Ans.= $\mathbf{-12346\ KJmol^{-1}}$)

8. In an experiment to determine the enthalpy of neutralization of hydrochloric acid and sodium hydroxide solution. $50cm^3$ of 1M hydrochloric acid solution was measured into a plastic cup and its initial temperature was noted. $50cm^3$ of 1M sodium hydroxide solution with known initial temperature was mixed with acid.

   Specimen results:

Initial temperature of acid = 20.80°C

Initial temperature of NaOH (aq) = 21.73°C

Final temperature of mixed = 27.90°C

(Assume: Density of the all solutions = $1gcm^{-3}$, specific heat capacity of the solution = 4.2J/g/°C)

Use results above to calculate molar enthalpy of neutralization of the acid by the base (Ans.= $-57.5KJmol^{-1}$*)*

RATES OF CHEMICAL REACTIONS (CHEMICAL KINETICS)

The study of chemical reactions and factors that affect the rates of chemical reactions is known as **chemical kinetics.**

The rate of a chemical reaction is the speed at which products are formed or reactants are used up in a chemical reaction.

Rate of reaction = $amount\ of\ products\ formed\ (moles\ per\ litre) / time\ (seconds)$

Most reactions occur in solutions and the amount is usually measured in moles/litre and time is measured in seconds, therefore, the unit of rate of reaction is **mol/l/s.**

Rates of reactions indicate how fast reactions are occurring, some reactions occur very rapidly e.g. explosions and precipitation while others occur very slowly e.g. rusting and fermentation. Some reactions proceed at a moderate rate e.g. reaction between hydrochloric acid and zinc metals.

Rate curves

The change in concentration (amount) of reactants of products with time is often plotted as the rate curve (figure a).

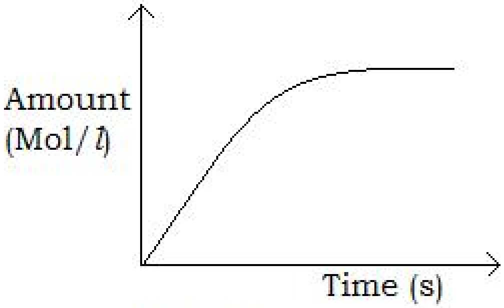


Figure a

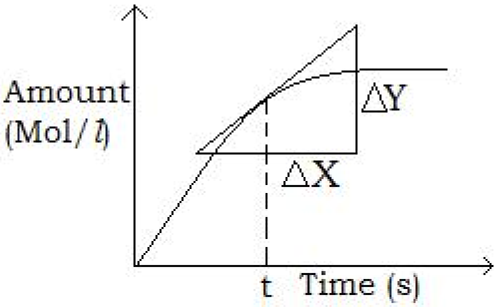


Figure b

The rate of reaction at any time, t can be found from the rate curve by drawing a tangent at that time as shown in figure (b). The gradient of the tangent is obtained as the rate of reaction at that time.

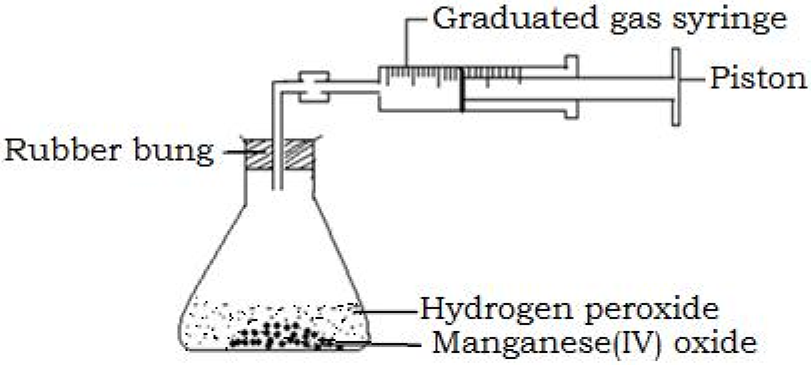
Gradient (rate of reaction) = $\text{change in } y \text{ axis}/_{\text{change in } x \text{ axis}}$ i.e Gradient = $^{\Delta y}/_{\Delta x}$

Determination of rate of reaction by

a) Change in gas volume

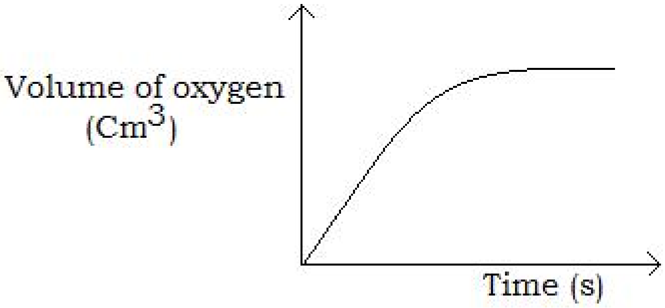
In reactions in which gases are formed, the volume of the gas can be recorded at various times. The rate curve is drawn and used to determine the rate of reaction. Examples are the decomposition of hydrogen peroxide and reaction of an acid and a metal.

i) Decomposition of hydrogen peroxide



Procedure

- A known volume of hydrogen peroxide is put in a flask as shown above and manganese (IV) oxide added to it.
- A rubber bung connected to a gas syringe is immediately inserted to close the flask. The stop clock is started the same time the flask is closed.
- The volume of oxygen gas collected in the syringe is read and recorded after a fixed time interval until when the reaction stops.
- The volume of gas evolved is then plotted against time. A rate curve as below is obtained.



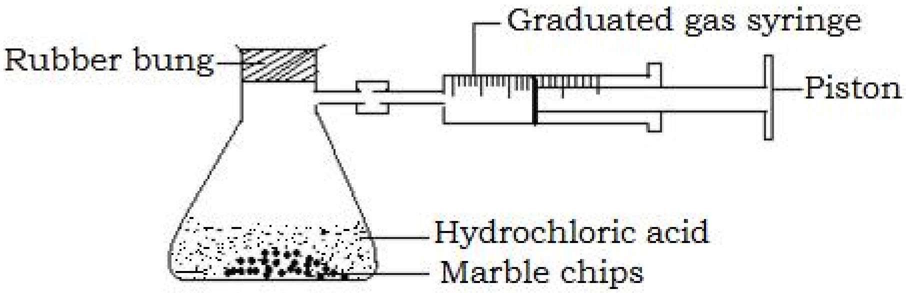
Equation for the decomposition

$2H_2O_2\ (aq) \xrightarrow{\hspace{4em}} 2H_2O\ (l) + O_2\ (g)$

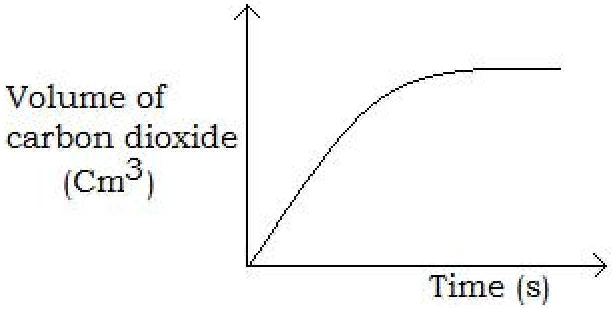
ii) Measuring volume of carbon dioxide produced when calcium carbonate reacts with hydrochloric acid

Procedure

- Place a known volume of dilute hydrochloric acid in a flask as shown below and add a known mass of calcium carbonate.
-



- A rubber bung connected to a gas syringe is immediately inserted to close the flask. The stop clock is started the same time the flask is closed.
- The volume of carbon dioxide gas collected in the syringe is read and recorded after a fixed time interval until when the reaction stops.
- The volume of carbon dioxide gas evolved is then plotted against time. A rate curve as below is obtained.



Equation

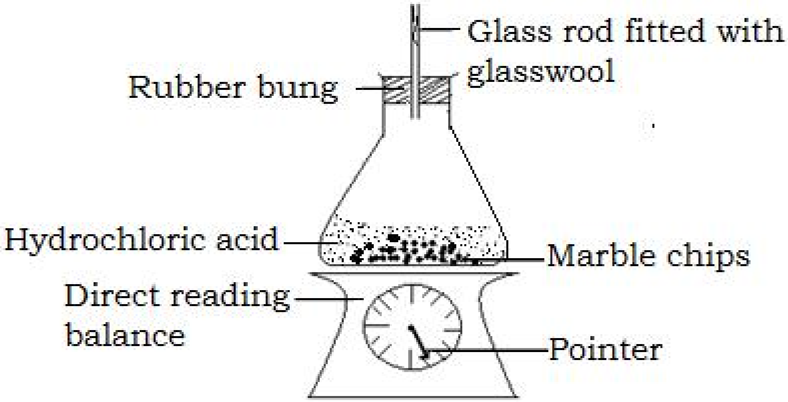
$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

NB. In the above case, it is necessary to add a little of the carbonate before the known mass of the carbonate is added. This is necessary because some of the carbon dioxide formed dissolves in the reaction mixture (solution), it is necessary to first saturate the solution with carbon dioxide so that all the carbon dioxide produced from the known mass is wholly measured.

b) Change in mass

Measuring the mass of the reaction mixture as carbon dioxide is evolved from reaction of calcium carbonate and hydrochloric acid

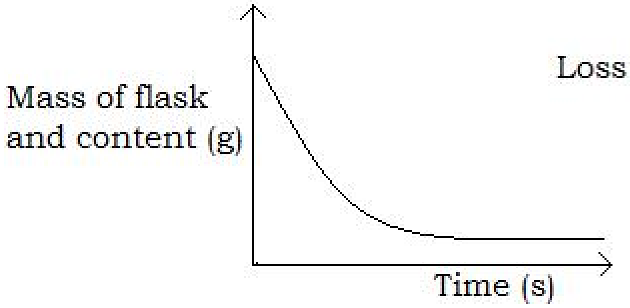
Set up

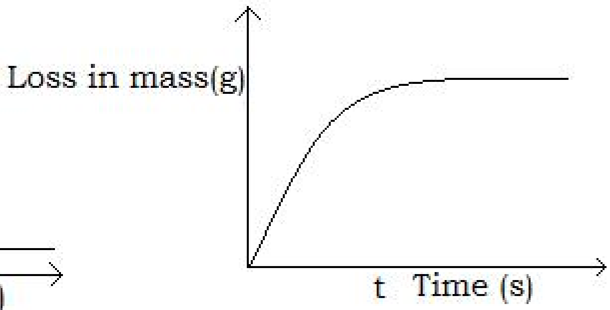


Procedure

- A flask containing a known volume of hydrochloric acid is weighed using a direct reading balance.

- A known mass of marble chips is added carefully and a rubber bung carrying glass tubing with cotton wool is immediately inserted to close the flask. A stop clock is started at the same time.
- The mass of the flask and its content is recorded at a regular interval of time.
- The results are then plotted on a graph.





The gradient is steep at the start of the reaction and gradually reduces until when it finally becomes zero (where it levels off) - at the end of the reaction. This implies that the rate of reaction is highest at the start and gradually decreases until when it becomes zero at the end of the reaction. This is because, at the start of the reaction, the concentration (amount) of reactants is highest but keeps on decreasing until when it becomes zero at the end of the reaction when all of it (reactants) have been consumed or used up.

The collision theory and rate of reaction

The collision theory states that before two or more substances can react, they must first collide. During a chemical reaction, molecules of reactants tend to approach one another, collide then chemical reactions take on. Therefore, the rate of chemical reaction depends on how close together the molecules of reactants are and how fast they are moving. Consequently the frequency of collision of molecules of reactants and rate of reaction is affected.

Factors affecting rates of reactions

A number of factors influence the rate of chemical reactions and these include: temperature, surface area (size of particles), concentration of reactants, light, catalyst, and pressure (especially for gaseous reactions).

1. Temperature

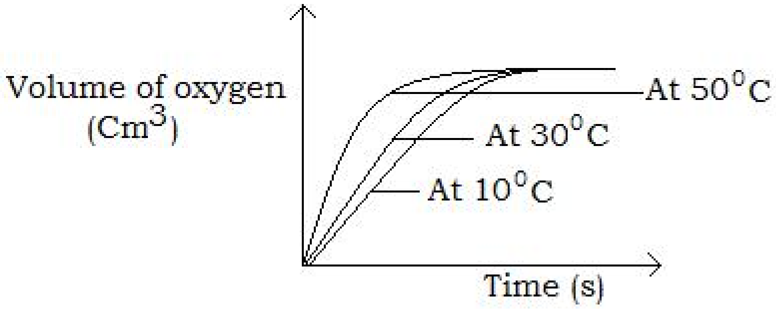
The rate of chemical reactions increase with increase in temperature. For every 10K rise in temperature, the rate of reaction is approximately doubled.

Explanation

When temperature increases, the molecules/ions of reactants gain more kinetic energy and tend to move faster. Their frequency of collision consequently increases which results into increased rate of reaction. Therefore, the higher the temperature, the higher the rate of reactions.

Example

Hydrogen peroxide decomposes at room temperature in the presence of manganese (IV) oxide to produce oxygen. The rate at which oxygen gas was produced at different temperatures was determined by measuring the volume of oxygen gas evolved at different temperatures and the graph below was obtained.



The graphs all end at the same level as the same amount of reactants were used.

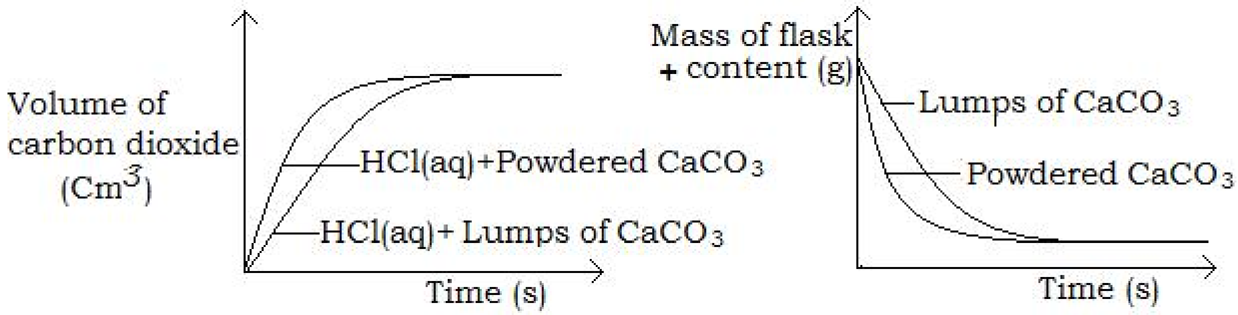
2. Surface area (size of the particles)

An increase in the surface area of the particles of solid reactants increases the rate of reaction if other factors are kept constant. This is because; increase in surface area increases the frequency of collision. Consequently, solids react more faster when in powdery form than when in large lumps.

Example

Calcium carbonate reacts more rapidly with hydrochloric acid when in powdery form than when in form of marble chips. This is because the powder form has a large surface area exposed and more readily collide with the acid.

Graphical illustration

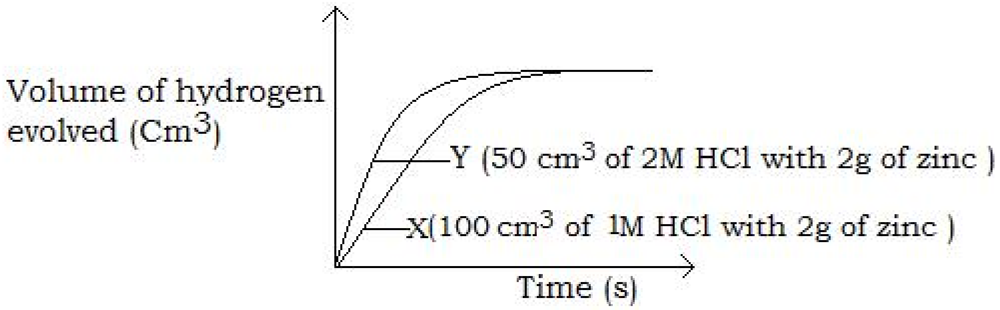


3. Concentration of reactants

Concentration refers to how close together the solute particles are in a given solution. Increasing the concentration of reactants increase the rate of a chemical reaction. The higher the concentration, the closer are the solute particles the higher the frequency of collision of the solute particles and this results into an increase in the rate of reaction.

Examples

i) Reaction of zinc and dilute hydrochloric acid



When 100cm^3 of 1M HCl was reacted with 2g of Zinc and the volume of hydrogen gas evolved measured at regular time interval is plotted against time, curve X is obtained. Curve Y is obtained when 50 cm^3 of 2M HCl reacted with 2 g of zinc.

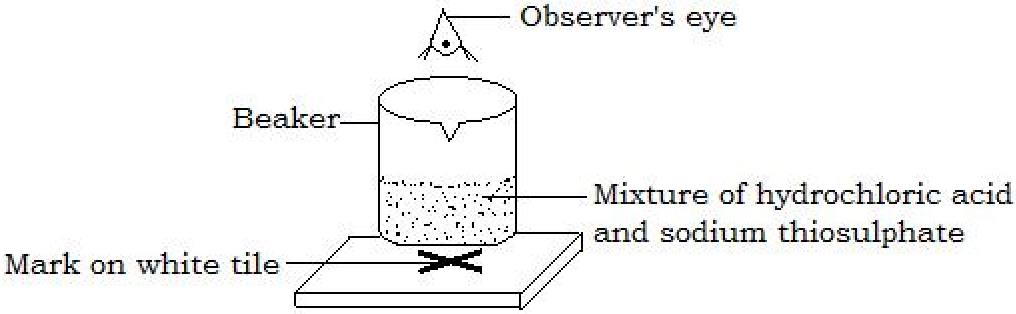
The gradient of curve Y is steeper than the gradient of curve X, therefore, the rate of reaction in Y is higher than in X. This is because in Y, a more concentrated acid was used than in X and the number of hydrogen ions per unit volume is higher in Y making them to collide more frequently and react more often than in X.

The gradient of Y is approximately twice that of X, because the acid used in Y is twice more concentrated as that used in X.

The curve for Y levels off first because the hydrochloric acid used in Y being more concentrated makes the reaction to reach completion much earlier. The maximum volume of gas (hydrogen) produced in X and Y is the same since the number of moles of hydrogen ions present in both acids are the same.

ii) Reaction of hydrochloric acid and sodium thiosulphate

Reaction of dilute hydrochloric acid and sodium thiosulphate produces precipitates of **sulphur.** The intensity of the precipitate can be studied by placing the beaker with the two reactants over a white tile with a mark on it as shown below.



The precipitates will eventually cover the mark. The time taken for the mark to disappear from view is recorded. The time for disappearance of the mark for different concentrations of the acid and thiosulphate are recorded.

Equation

$Na_2S_2O_3(aq) + 2HCl(aq) \longrightarrow S(s) + SO_2(g) + 2NaCl(aq) + H_2O(l)$

Ionic

$S_2O_3^{2-}(aq) + 2H^+(aq) \longrightarrow S(s) + SO_2(g) + H_2O(l)$

If the concentration of the dilute hydrochloric acid or thiosulphate is increased, the time for disappearance of the mark decreases hence the rate is faster when the solutions are made more concentrated.

Trial question

Procedure

- Mark a cross with blue or black ink on a piece of paper.
- Place 50cm^3 of sodium thiosulphate solution in a beaker.
- Add 5.0cm^3 of 2M hydrochloric acid provided to the sodium thiosulphate solution and start the stop clock immediately.
- Shake the beaker gently for the solution to mix well and then place it on the paper over the cross.
- Watch the cross through the solution. i.e. from above the beaker.
- Stop the clock as soon as the cross disappears.
- Record the time taken for the cross to disappear.
- Pour away the mixture and rinse the beaker thoroughly.
- Place 40cm^3 of thiosulphate solution and then add 10cm^3 of water into the solution.
- Added 5 cm^3 of the acid as before and then follow the procedure given above. Record the time taken for the cross to disappear.
- Repeat the above procedure for the rest of the mixture as shown in the table below.

Experiment	Volume of $Na_2S_2O_3$/cm^3	Volume of water/ cm^3	Volume of HCl/ cm^3	Time /s	1able below /time (s^{-1})
1	50	0.0	5.0		
2	40	10.0	5.0		
3	30	20.0	5.0		
4	20	30.0	5.0		
5	10	40.0	5.0		

a) Write an ionic equation for the reaction.
b) Why does the cross disappear in the experiment?
c) Plot a graph of volume of thiosulphate against time (describe and explain the shape of the graph)

d) Plot a graph of $^1/_{time}$ against volume of thiosulphate solution (calculate the gradient/ slope of your graph and state its unit)

4. Pressure

This affects gaseous reactions since gases unlike solids are compressible. Increasing pressure on gases brings the reactant particles close to each other increasing the frequency of collision and hence the rate of reaction. Pressure can be increased by decreasing the volume of the container. For example in the Haber process, a large yield of ammonia is obtained from high pressure as per the following reaction.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

5. Light

Light is another form of energy that speeds up rates of chemical reactions. Reactions whose rates are affected by light are said to be **photosensitive.** Examples of such reactions include:

i) Reaction of hydrogen and chlorine which occurs explosively in the presence of light.

$$H_2\ (g) + Cl_2\ (g) \xrightarrow{\text{U.V light}} 2HCl\ (g)$$

ii) Photosynthesis which is a chemical reaction dependent on light. The light is trapped by chlorophyll-a green pigment in plants.

$$6CO_2\ (g) + 6H_20(l) \longrightarrow 6O_2\ (g) + C_6H_{12}O_6\ (aq)$$

iii) Formation of silver from silver salts that takes place when a photographic film is exposed to light.

$$2AgBr(s) \xrightarrow{\text{U.V light}} 2Ag(s) + Br_2\ (g)$$

When testing for chlorides using silver nitrate, the white precipitates obtained (silver chloride) slowly becomes grey on exposure to light. This is due to decomposition to form metallic silver.

$$2AgCl(s) \xrightarrow{\text{U.V light}} 2Ag(s) + Cl_2\ (g)$$

6. Catalyst

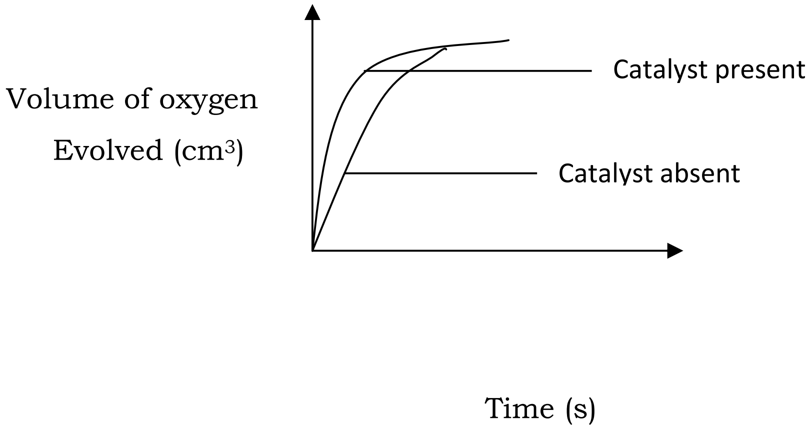
A catalyst is a substance that alters the rate of a chemical reaction but remains chemically unchanged at the end of the reaction. A catalyst may change physically during the course of a reaction but not chemically.

NB. To prove that catalysts do not change chemically, weigh the mass of the catalyst before and after a reaction- the mass does not change.

A catalyst (positive catalyst) lowers the activation energy required to initiate a chemical reaction and consequently increases the rate of the reaction.

Illustration

The rate of decomposition of hydrogen peroxide is low without a catalyst and increases when a catalyst (manganese (IV) oxide) is added.



At the beginning, the curves are steep but get less steep as time goes by until the end of the reaction where they level off. The rate of reaction is highest at the beginning but decreases steadily throughout the reaction until when the reaction stops. This is because the concentration of hydrogen peroxide decreases as the reaction goes on.

Curve for the reaction that had a catalyst has a steeper gradient and levels off before the curve for reaction without a catalyst. This is because; a catalyst ensures a more effective collision per second and therefore increasing the rate of reaction.

Some catalysts lower the rate of chemical reactions and are known as negative catalysts.

Chemical reactions and equilibrium/'balance'

Chemical reactions can further be sub divided into classes of reactions each of which has its own characteristics and a few of which are: combination, displacement, decomposition and double decomposition.

i) Combination

This takes place when two or more substances combine to form a single substance. E.g when iron is heated with sulphur to form iron (II)sulphide.

$$Fe(s) + S(s) \longrightarrow FeS(s)$$

When lead (IV) oxide is lowered into a gas jar of sulphur dioxide, they combine and form lead (II)sulphate.

$$PbO_2(s) + SO_2(s) \longrightarrow PbSO_4(s)$$

ii) Decomposition

This occurs when a compound spits up into simpler substances. This change usually takes place without the presence of a second substance and very often the action of heat is sufficient to cause the reaction to take place. E.g. When calcium carbonate is heated in a crucible to bright red, it decomposes to form calcium oxide and carbon dioxide.

$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$

iii) **Displacement**

This occurs when one element (or group) takes the place of another element (or group) in a compound. E.g. if zinc is placed in copper (II) sulphate solution, copper is displaced by zinc and zinc sulphate is left.

$Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$

When chlorine is bubbled through a solution of potassium bromide, the chlorine displaces bromine and red bubbles of bromine are formed. A solution of potassium bromide is left.

$2KBr(aq) + Cl_2(g) \longrightarrow 2KCl(aq) + Br_2(g)$

iv) **Double decomposition**

In this reaction, two compounds take part, both are decomposed and two new substances are formed by exchange of radicals. In most cases both reactants are soluble in water and only one of the products is soluble. The insoluble product is precipitated out in form of solids (it is this product that is normally wanted).

When hydrogen sulphide gas is bubbled through a solution of copper (II)sulphate, copper (II)sulphide and sulphuric acid are formed.

$CuSO_4\ (aq) + H_2S\ (g) \longrightarrow CuS(s) + H_2SO_4\ (aq)$

Less frequently is the wanted product of double decomposition more volatile and it is driven of as a gas or by heating as a vapour of a more volatile liquid.

$NaCl(aq) + H_2SO_4(aq) \longrightarrow NaHSO_4(s) + HCl(g)$

Reversible reactions

Most reactions proceed in only the forward direction until when one of the reactants or all the reactants are used up then the reaction stops. Such a reaction is known as an **irreversible reaction.**

An irreversible reaction is represented by a full arrow ‘ → ’ .In such a reaction, the reactants and the products can easily be identified. E.g. the reaction of hydrogen and oxygen to form water.

$H_2(g) + ½O_2(g) \longrightarrow H_2O(g)$

However, there are some reactions in which the direction of chemical change can be reversed by changing the conditions under which the reaction is taking place. The reaction can proceed in both forward and backward directions depending on the conditions to which the reaction is subjected i.e. the products can react to form back the substances that were initially reacting. Such a reaction is known as a **reversible reaction.** E.g when hydrated copper (II) sulphate is heated, the blue colour of the crystals changes to white due to formation of anhydrous copper (II) sulphate.

$CuSO_4.5H_20(s) \longrightarrow CuSO_4(s) + 5H_2O\ (g)$

(blue crystals) (White powder)

However, the white anhydrous copper(II) sulphate can be turned to blue hydrated form by adding water to it drop by drop.

$CuSO_4(s) + 5H_2O(l) \longrightarrow CuSO_4.5H_20(s)$

(white powder) (blue crystals)

The two equations above can be written as one:

$CuSO_4.5H_20(s) \rightleftharpoons CuSO_4(s) + 5H_2O(g)$

The sign $\rightleftharpoons$ indicates that the reaction proceeds in both forward and backward directions.

Chemical balance (chemical equilibrium)

In reversible reactions, both forward and backward reactions can take place at the same time and the reaction might come to some kind of a ‘balance’ in which the products and reactants are present at the same time.

Suppose n molecules of substance A reacts with m molecules of substance B to form x molecules of C and y molecules of D as shown below,

$$nA + mB \rightleftharpoons xC + yD$$

As soon as a little of C and D are formed, a reverse reaction will begin. At first, the forward reaction will predominate, but, as C and D accumulate, the reverse reaction will build up until when a ‘balance, (equilibrium position) is reached with both the forward and backward reactions proceeding at the same rate. The composition of the mixture will appear to be constant though it is the net result of the two opposing reactions.

Factors affecting chemical equilibrium (balance)

The factors affecting equilibrium position were investigated by Louis Henri Le Chatelier and he summarized it in a statement known as Le Chatelier’s principle. The principle states that

“*if an external condition is applied to a reaction in equilibrium, the equilibrium will shift to a direction which tends to cancel the effect of the applied condition*”

These factors that affect equilibrium position are temperature, pressure (for gases), concentration and catalyst.

i) Temperature

The effect of temperature depends on whether the reaction is exothermic or endothermic.

If a reaction is exothermic (gives out heat) and heat is applied to the system, the reaction shifts in a direction that tends to lower the temperature i.e. the reverse reaction occurs. Consider the reaction below in the Haber process of manufacturing ammonia.

$$N_2(s) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H = -92\ kJ/mol$$

Since the reaction is exothermic, when temperature is raised, the reverse reaction which consumes heat occurs i.e. more of nitrogen and hydrogen will be produced from ammonia. The backward reaction is thus favoured.

If the reaction is endothermic (absorbs heat) and the temperature of the system is raised, the reaction proceeds in such a way that the heat applied is used up. The reaction proceeds in a forward direction.

ii) Pressure

A change in pressure affects mainly reactions which involve gases. The effect of pressure depends on whether there is an increase or decrease in number of molecules. Reactions that do not proceed with change in the number of molecules/moles are not affected by pressure.

Consider the reaction in Haber process below

$$N_2(s) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

The reaction proceeds with a decrease in number of molecules and thus decrease in pressure. Increasing pressure makes the reaction to proceed in the forward direction and produce more ammonia. When pressure is decreased, backward reaction is favored since the reaction proceeds with increase in pressure.

iii) Concentration

If the concentration of one of the substances present in equilibrium reaction is changed without change in any other conditions then, by Le Chatelier's principle, the position of equilibrium will move to decrease the concentration of the added substance.

Consider the reaction below occurring in a vessel,

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$$

If extra oxygen was pumped into the vessel, the position of equilibrium will move from left to right i.e. the forward reaction is favoured producing more nitrogen monoxide. In a similar way, if a substance taking part in a chemical equilibrium is removed by some means, then, the equilibrium position will change to produce more of that substance.

iv) Catalyst

This only increases the rate at which equilibrium is attained. It has no effect on the amount of products formed but just makes the reaction to proceed faster.

Important industrial applications of chemical equilibrium

1. Synthesis of ammonia by Haber process

Ammonia is manufactured by Haber process according to the following equation:

$N_2(s)+3H_2(g) \rightleftharpoons 2NH_3$ (g)the reaction is exothermic (ΔH=-92 kJ/mol)

1 vol 3 vol 2 vol (at constant pressure and temperature)

In this process, ammonia is prepared from nitrogen and hydrogen. The nitrogen used is got from distillation of liquid air and nitrogen from water gas. The yield of ammonia will depend on conditions that favour its production and these include;

High pressure: This is because the reaction proceeds with a decrease in number of moles and hence decreases in volume. High pressure would increase the yield of ammonia (Le Chatelier's principle), therefore a pressure of 200-500 atm is used.

Low temperature: Since the reaction proceeds with evolution of heat, the forward reaction is favoured by low temperature, so, the formation of ammonia will increase with decrease in temperature and decrease with increase in temperature. In practice a temperature of about 450˚C is used.

However, at very low temperatures, the rate of the reaction is reduced, so it is necessary to introduce a catalyst which would give a high yield in spite of a relatively low temperature. The catalyst that is used is **finely divided iron.**

High concentration of nitrogen and hydrogen: Increasing the concentration of either nitrogen of hydrogen leads to production of more ammonia as the added nitrogen of hydrogen will be used up to form ammonia.

2. Formation of sulphuric acid by contact process

This involves the following steps

i) Formation of sulphur dioxide. The sulphur dioxide can be formed by

a) Burning sulphur in air

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

b) Burning metal sulphides in excess air e.g.

$$2PbS(s) + 3O_2(g) \longrightarrow 2PbO(s) + 2SO_2(g)$$

$$ZnS(s) + 3O_2(g) \longrightarrow 2ZnO(s) + 2SO_2(g)$$

ii) Purification and formation of sulphur trioxide

The sulphur dioxide and oxygen are purified, mixed together and reacted to form sulphur trioxide.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \qquad \Delta H = -197\ kJ/mol$$

Since the reaction is exothermic, the yield of sulphur trioxide is favoured by low temperature. In practice, a temperature of 450-500°C is used. The rate of the reaction is increased by adding a catalyst (**vanadium (V) oxide**).

The reaction proceeds with a decrease in number of moles and thus decrease in volume, therefore high pressure is required for more yield of sulphur trioxide. In practice, the pressure used is between 1-10 atm since the cost of maintaining high pressure is high.

iii) The sulphur trioxide is dissolved in concentrated sulphuric acid to form fuming sulphuric acid (oleum)

$SO_3(g) + H_2SO_4(aq) \longrightarrow H_2S_2O_7(aq)$

The sulphur trioxide is not directly dissolved in water because it would react exothermically.

iv) The fuming sulphuric acid is diluted with water to form very concentrated sulphuric acid (about 98% concentrated)

$H_2S_2O_7(aq) + H_2O(l) \longrightarrow 2H_2SO_4(aq)$

3. Manufacture of nitric acid

Ammonia is burnt in excess air over red hot catalyst (platinum (90%)/rhodium (10%) gauze catalyst) to form nitrogen monoxide according to the reaction,

$4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(l)$

The nitrogen monoxide is cooled and reacts with excess air to form nitrogen dioxide.

$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$

The nitrogen dioxide is mixed with excess oxygen and bubbled through water (hot) to form nitric acid. Excess oxygen is used to prevent formation of nitrous acid.

$4NO_2(g) + 2H_2O(l) + O_2(g) \longrightarrow 4HNO_3(aq)$

Sample questions on Rates of reaction

1. *Explain (i) what is meant by the term "rate of a chemical reaction" (ii) the effect of concentration of reactant on the rate of a chemical reaction. The table below shows the times taken for reaction of a certain substance Z to go to completion when solutions containing various concentrations of Z were used.*

Concentration of Z (moldm $^{-3}$)	*0.1*	*0.3*	*0.4*	*0.6*	*0.8*
Time, t for completion of reaction (s)	*120*	*40*	*30*	*20*	*15*
Reciprocal of time $^1/_t$ (s^{-1})					

Calculate the value of $^1/_t$ for each time, t above and enter your answer in the space provided in the table above. Plot a graph of $^1/_t$, vertical axis against concentration of Z. Deduce from your graph how the rate of the reaction varies with concentration of Z. Draw a sketch graph to show how volume of carbon dioxide would vary with time if excess dilute hydrochloric acid was added to a certain mass Wg of marble chips and label it x. Draw on the same axes the sketch graph you would expect if equimolar volume of the hydrochloric acid was added to Wg of finely ground marble chips; and label it Y.

State one factor which can affect the rate of a chemical reaction other than concentration. Mention the effect of the factor you have stated in (d) (i) on the rate of reaction.

2. *Describe fully one reaction to illustrate each of the following. (a) a reaction whose rate is increased by raising temperature (b) a reaction whose rate is increased by raising pressure (c) a reaction whose rate is affected by presence or absence of light.*

3. *The rate of decomposition of solution of hydrogen peroxide is increased by the presence of a catalyst. Write down an equation for the decomposition and name the catalyst. Describe with the aid of a diagram an experiment to measure the rate of decomposition of the hydrogen peroxide?*
4. *Name four factors that affect the rate of a reaction. Describe how the named factors affect the rate of reaction.*
5. *Dilute hydrochloric acid reacts with Zinc according to the equation;*

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

25cm^3 of dilute hydrochloric acid of various concentrations were placed in six different beakers. 0.30g of the same Zinc powder were separately added to each beaker. The time taken for effervescence to stop was noted.

Table of results

Concentration of HCl (moldm-3)	0.25	0.50	0.75	1.00	1.25	1.50
Time taken for effervescence to stop (s)	3.20	22.0	14.5	9.5	6.0	3.0

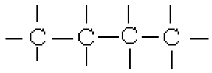
*Plot a graph of concentration of acid against time. Use the graph to determine the rate of reaction (i) After 10 seconds (**Ans.= -0.053moldm^{-3}s^{-1}**) (ii) After 25 seconds(**Ans.= -0.376Moldm^{-3}s^{-1}**). Compare the reaction rates in b(i) and b(ii). Explain your answer*

ORGANIC CHEMISTRY

Organic chemistry is the chemistry of carbon and its compounds. These organic compounds contain carbon as the basic frame work and other elements like hydrogen, nitrogen and chlorine are attached to it.

Carbon has a unique behavior in a chemical sense because:

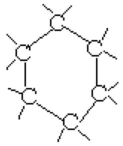
1. It can form a very long chain of carbon atoms which can be up to 2000 atoms.

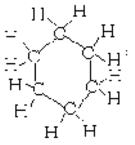
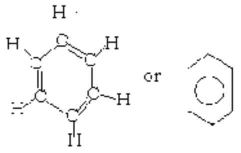
   $-C-C-C-C-$

   These compounds consisting of chains of carbon atoms are called **aliphatic compounds.** These compounds can be **saturated** (if all the carbon atoms are joined to each other by a single covalent bond e.g. ethane, CH_3-CH_3) or **unsaturated** (if it contains multiple covalent bonds i.e. either double or triple e.g. ethene, $CH_2=CH_2$ and ethyne, $H-C\equiv C-H$,

2. It can form a ring of carbon atom. The compounds that form rings of carbon atoms are **alicyclic compounds.**



These compounds can also be **saturated**

e.g. cyclohexane  or **unsaturated** e.g. benzene 

Because of these two unique behaviors, carbon can form very many and complex compounds which has made it necessary for its study under a separate branch called organic chemistry. However, for historical and conventional reasons some simpler compounds such as carbon dioxide (CO_2) and sodium carbonate (Na_2CO_3) are usually studied under non carbon compound in inorganic chemistry.

Classification of organic compounds

Organic compounds can be classified into several groups. The simplest of the organic compounds are hydrocarbons. Other groups include: alcohols, esters, carboxylic acids, amines, ketones, alcohols and ethers.
These groups are differentiated from each other by **functional groups**.

Functional groups are groups of atoms that are common to a given homologous series and are responsible for chemical reactions. Examples of functional groups include:

-OH for alcohols e.g. ethanol, CH_3CH_2OH; methanol, CH_3OH
-COOH for carboxylic acids e.g. Ethanoic acid, CH_3COOH; methanoic acid HCOOH
-NH_2 for amines e.g. amino ethane, $CH_3CH_2NH_2$, amino propane, $CH_3CH_2CH_2NH_2$

Homologous series

This is a series of organic compounds related to each other by the same functional group. Characteristics of homologous series include:

i) All members conform to a general molecular formula e.g.
C_nH_{2n+2} for alkanes. If n=2, C_2H_6(ethane); if n=4, C_4H_{10} (butane)

C_nH_{2n} for alkenes. If n=2, C_2H_4 (ethene); if n=3, C_3H_6 (propene)

ii) Members of the same homologous series have the same chemical properties (though varying in vigour/speed)

iii) The physical properties of the members change gradually with increase in molecular mass. E.g. boiling point, melting point and density increase with increase in molecular mass; there is a gradual change in state down the group (methane is a gas, pentane is a liquid and decane is a solid); solubility decreases down the group as molecular mass increases.

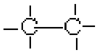
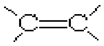
iv) Members in each homologous series differ from the next by $-CH_2$ group (methylene group).

v) Members have the same general method of preparation

Hydrocarbons

These are compounds consisting of only hydrogen and carbon atoms. They have a general formula of C_xH_y where x and y can be any numerical whole numbers.

Hydrocarbons are classified into three main groups as alkanes, alkenes and alkynes. These three are differentiated by the following functional groups.

Alkanes,  $-\overset{|}{\underset{|}{C}}-\overset{|}{\underset{|}{C}}-$ Alkenes,  Alkynes, $-C\equiv C-$

Alkanes

These are saturated hydrocarbons with the general formula of C_nH_{2n+2}. Where n is the number of carbon atoms. Alkane members are referred to as the paraffin i.e. they have little affinity to react.

Sources of alkanes

The main sources of alkanes include:

i) Natural gas. This contains mainly methane with small amounts of other gases like propane and butane. Methane is formed by anaerobic decomposition of organic matter and it is found in swamps, stagnant ponds and marshes.

ii) Petroleum. This contains a wide range of alkanes ranging from molecular gases to high molecular waxy solids (C_2-C_{40}). Petroleum is formed by anaerobic decomposition of sea plants and animals. The components of petroleum are separated by fractional distillation, a process known as refining.

Nomenclature of alkanes

According to IUPAC (International Union of Pure and Applied Chemistry), all members of alkanes have their names ending with the suffix –ane.

Value of n	**Formula**	**Name**
1	CH_4	Methane
2	C_2H_6	Ethane
3	C_3H_8	Propane
4	C_4H_{10}	Butane
5	C_5H_{12}	Pentane

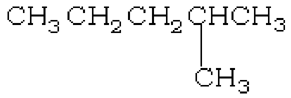
Straight chain alkanes have a prefix n before the normal name e.g.

$CH_3CH_2CH_2CH_3$ n-butane

In branched chains, the branch may be a hydrocarbon or other atoms like chlorine, and bromine.

The hydrocarbon side chains have one hydrogen less the parent alkanes and are generally referred to as **alkyl groups**. The alkyl groups derive their names from respective parent alkanes e.g.-CH_3 (methyl); -CH_2CH_3 (ethyl); -$CH_2CH_2CH_3$ (propyl); -$CH_2CH_2CH_2CH_3$ (butyl).

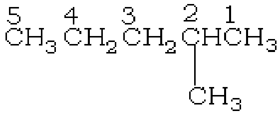
For branched alkanes e.g.



$$CH_3CH_2CH_2\underset{\displaystyle CH_3}{\underset{|}{C}}HCH_3$$

i) Name the longest unbranched carbon chain .i.e. the longest carbon chain consists of 5-carbon atoms, it is therefore a derivative of pentane

ii) Give number showing the position of the branch. Count from the side that will give the branch the lowest possible number. E.g. the branch is on the 2nd carbon atom (the branch replaces hydrogen on the 2nd carbon)



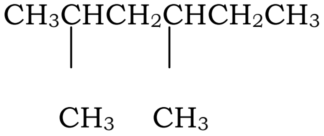
$\overset{5}{CH_3}\overset{4}{CH_2}\overset{3}{CH_2}\underset{CH_3}{\overset{2}{CH}}\overset{1}{CH_3}$

iii) Name the branch (substituent group) i.e. methyl group

So, write the name of the alkane starting with the carbon position on which the branch is located (2); put a dash (-); write the name of the branch/substituent group (methyl) followed by the name of the longest straight carbon chain.

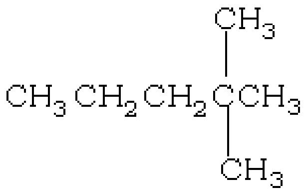
The above compound is therefore 2-methylpentane.

vi) If the branches of side chains are more than one and are similar, di, tri, etc are used.



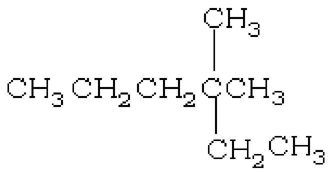
$CH_3\underset{CH_3}{CH}CH_2\underset{CH_3}{CH}CH_2CH_3$

2,4-dimethylhexane



$CH_3CH_2CH_2\overset{CH_3}{\underset{CH_3}{C}}CH_3$

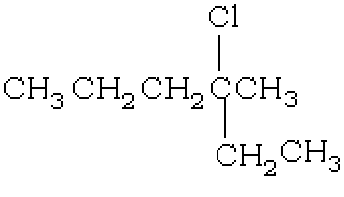
2,2-dimethylpentane



$CH_3CH_2CH_2\overset{CH_3}{\underset{CH_2CH_3}{C}}CH_3$

3,3-dimethyl hexane

vii) If the side chains are different, naming follows alphabetical order



$CH_3CH_2CH_2\overset{Cl}{\underset{CH_2CH_3}{C}}CH_3$

Note: the longest unbranched chain has 6-C atoms, it's there fore a derivative of hexane.

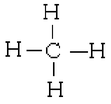
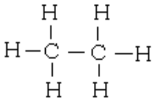
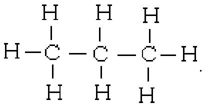
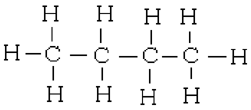
3-chloro,3-methylhexane

Molecular and structural formulae

Molecular formula shows the number of each kind of atoms present in one molecule of a compound. It does not show the arrangement of atoms in the molecule.

Structural formula (graphical formula) shows the arrangement of atoms in one molecule of a compound.

Alkanes like other hydrocarbons and other organic compounds have covalent bonds between the atoms. In alkanes, the carbon atoms use all the four outer most electrons to form covalent bonds by sharing with other carbon atoms and hydrogen atoms. Because all the electrons are used up in the formation of covalent bonds, they are called **saturated hydrocarbons**.

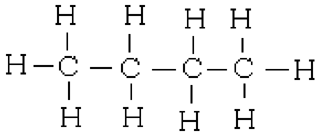
Compound	**Molecular formula**	**Structural formula**
Methane	CH_4	 $H-\overset{H}{\underset{H}{C}}-H$
Ethane	C_2H_6	 $H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$
Propane	C_3H_8	 $H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$
Butane	C_4H_{10}	 $H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$

Isomerism

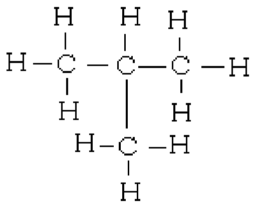
This is the existence of a compound with the same molecular formula but different structural formula.

Isomers are compounds with the same molecular formula but different structural formula.

All hydrocarbons with four or more carbon atoms per molecule posses isomers. E.g. butane (C_4H_{10})

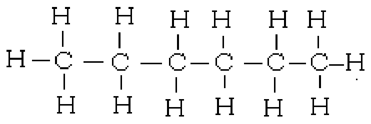


n-butane

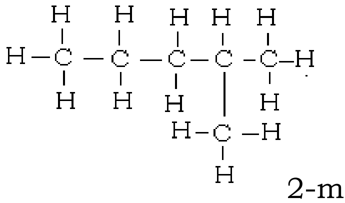


2-methyl propane

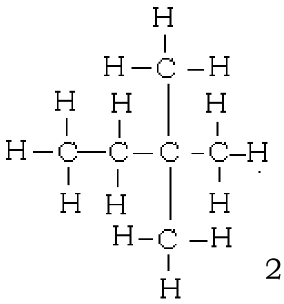
Hexane (C_6H_{14})



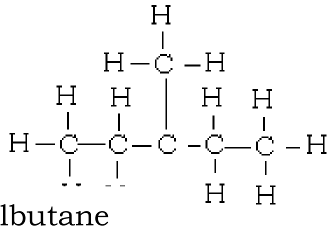
n-hexane



2-methylpentane



2,2-dimethylbutane



3-methylpentane

Properties of alkanes

Alkanes are not so reactive and under go combustion and chlorination reactions only.

1. Combustion

Alkanes under go complete combustion in plenty of oxygen to form carbon dioxide and water vapour. For example, methane explodes in air/ oxygen on application of flame

$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$ + Heat.

Incomplete combustion of alkanes in limited supply of oxygen produces carbon monoxide, and sometimes carbon and water vapour.

$2CH_4(g) + 3O_2(g) \longrightarrow 2CO(g) + 4H_2O(g)$ + Heat.

Or

$CH_4(g) + O_2(g) \longrightarrow C(s) + 2H_2O(g)$ + Heat.

The combustion of alkanes produces considerable amount of heat. This explains why they are used as fuel for domestic and industrial uses.

The in complete combustion of carbon occurs in cylinders of petrol engines that results in to release of poisonous carbon monoxide and some times even carbon. It is therefore dangerous to run a car engine in a garage where there is no free air circulation.

2. Chlorination

Alkanes under go substitution reaction with halogens. A substitution reaction is reaction in which an atom or a group of atoms in a compound is/are replaced by other atoms.

For the case of alkanes, this is only possible with halogens e.g. when sunlight shines on a mixture of methane and chlorine, the chlorine replaces hydrogen in a chain reaction i.e. substitution reaction occurs as follows:

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

(Chloromethane)

$CH_3Cl(g) + Cl_2(g) \longrightarrow CH_2Cl_2(g) + HCl(g)$

(Dichloromethane)

$CH_2Cl_2(g) + Cl_2(g) \longrightarrow CHCl_3(g) + HCl(g)$

(Trichloromethane)

$CHCl_3(g) + Cl_2(g) \longrightarrow CCl_4(g) + HCl(g)$

(Tetrachloromethane)

This reaction occurs rapidly in bright sunlight and when chlorine is in excess.

Petroleum (Crude oil)

Petroleum is formed by anaerobic decomposition of sea plants and animals. It is oil consisting of different alkanes normally ranging from C_5H_{12} to $C_{43}H_{88}$. The oil deposits are usually found with sand and brine.

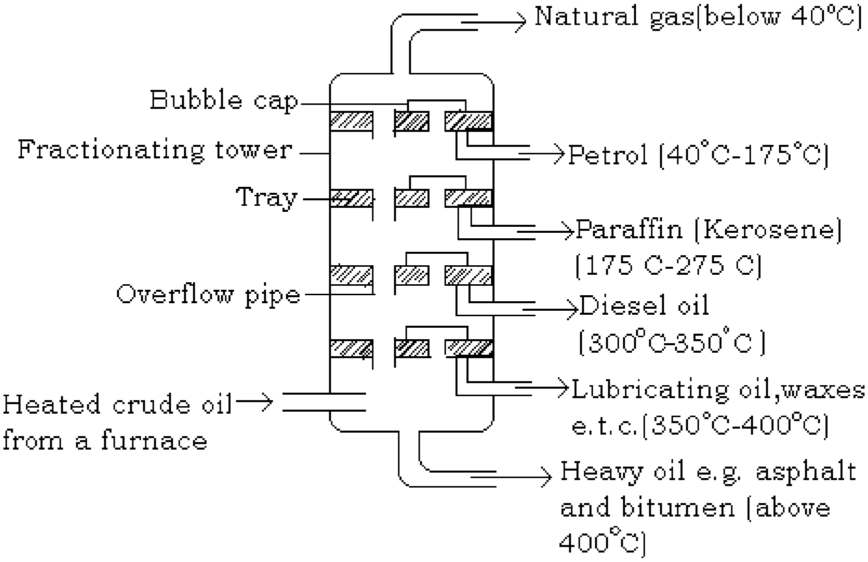
Refining fuel

The different alkanes that make up petrol can be separated by fractional distillation. This is based on the boiling points of the different components.

After the removal of impurities mainly sulphur compounds, it is heated until when most of it vaporizes. The vapour is passed into the bottom of a tall fractionating tower. The fractionating tower is divided into several compartments each cooler than the one below it.

During fractional distillation, the fraction of petroleum that is most volatile settles at the top and the non volatile heavy oil runs out and the bottom of the column.

Fractionating tower



Crude oil is separated into the following fractions

Fraction	**Distilling temperature**	**Number of carbon atoms**	**uses**
Natural gas e.g methane,ethane,propane	Below 40°C	C_1-C_4	Fuel for lighting and heating purposes
Petrol + naphthalene	40°C to 175°C	C_5-C_{10}	Fuel; solvent for grease, paints and stain;vanish,dry cleaning
Paraffin	175°C to 275°C	C_{10}-C_{14}	Fuel
Diesel + gas oil	275°C to 350°C	C_{14}-C_{18}	Fuel
Lubricating oil+paraffin Wax+vaseline	Above 350°C	C_{18}-C_{40} and above	Lubrication, making candle, making Vaseline.

The gas oil fraction is cracked to yield more petrol.

Cracking of gas oil

Fractional distillation of crude oil above only yields 20% of the petrol. More petrol is produced by the cracking process.

Cracking is the breaking down of large complex hydrocarbons into smaller molecules (of short carbon chain) by use of heat or catalyst. Heavy alkanes are cracked to produce useful alkenes and fuel of high quality (relatively smaller alkanes). E.g.

$C_{10}H_{22} \longrightarrow C_7H_{16} + C_3H_6$

Cracking can be classified in to:

i) **Thermal cracking:** this involves heating of large hydrocarbons at high pressures to break them into smaller molecules**.**

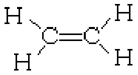
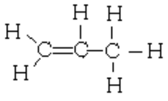
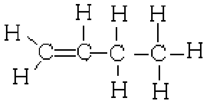
ii) **Catalytic cracking:** this involves the use of a catalyst to break down large and complex hydrocarbons in to simpler ones. Catalysts commonly used are silicon(IV) oxide and aluminium oxide. Catalytic cracking takes place at a relatively low temperature and pressure.

ALKENES

Alkenes are unsaturated hydrocarbons with a general formula of C_nH_{2n}. where n=2 or more. They are characterized by possession of a double bond between carbon atoms.

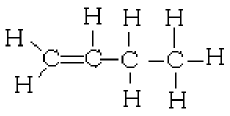
Nomenclature and structure

Alkenes are named as alkanes except that their names end with suffix –ene. Consider the table below.

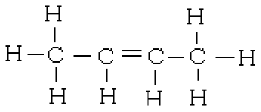
N	Molecular formula	Name	Structural formula
2	C_2H_4	Ethene	 or $CH_2=CH_2$
3	C_3H_6	Propene	 or $CH_2=CHCH_3$
4	C_4H_8	Butene	 or $CH_2=CHCH_2CH_3$

Isomerism

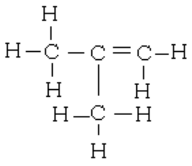
Isomerism in alkenes begins when n=4 i.e from butene. Isomers of butane are:



But-1-ene



But-2-ene



2-methylpropene

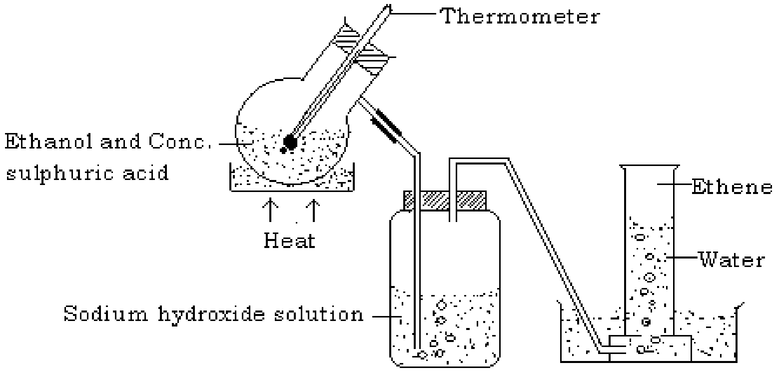
Ethene

This is the simplest alkene with molecular formula, C_2H_4.

Laboratory preparation

Ethene is prepared by dehydration of ethanol using excess concentrated sulphuric acid.

Set up



Procedure

- To $50cm^3$ of ethanol, add $100cm^3$ of concentrated sulphuric acid slowly while shaking under a tap
- The apparatus is set as above and the mixture heated with care to 180°C. Ethene is evolved and is collected over water.

NB. The wash bottle of alkali solution removes sulphur dioxide produced in small quantity as ethanol reduces sulphuric acid slightly. The alkali also removes fumes of the acid.

$$CH_3CH_2OH(l) \xrightarrow[170^\circ C\text{-}180^\circ C]{\text{Excess Conc. } H_2SO_4} C_2H_4(g) + H_2O(g)$$

Some times aluminium sulphate is added to the reaction to reduce frothing.

Ethane can also be prepared by catalytic dehydration of ethanol. Here, ethanol vapor is passed over a heated catalyst to produce ethane.

Properties of ethene

Physical properties

- Is a colorless gas with a faint sweet smell
- It is insoluble in water but soluble in organic solvents eg benzene and methylbenzene
- It is slightly less dense than air

Chemical properties

Alkenes are generally more reactive than corresponding alkanes. They undergo the following reactions

a) Combustion

Ethane burns in excess oxygen with a smoky flame since it contains a relatively high percentage of carbon forming carbon dioxide and water vapor

$$C_2H_{4(g)} + 3O_{2(g)} \longrightarrow 2CO_{2(g)} + 2H_2O_{(g)}$$

b) Addition reaction of ethene

Ethene and other unsaturated compounds undergo addition reactions an addition reaction is one in which a molecule adds to an unsaturated compound by breaking the double bond or triple bond

Examples;-

i) When ethene gas is bubbled through bromine water, bromine water changes from red brown to colorless i.e. bromine water is decolorized or the red brown color of bromine is discharged

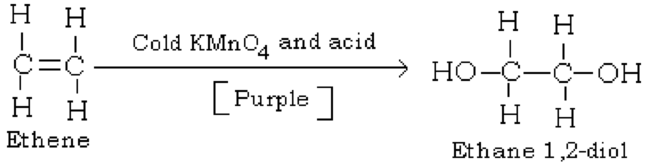
$$C_2H_{4(g)} + Br_2(l) \longrightarrow C_2H_4Br_{2(l)}$$

Or

$CH_2{=}CH_2(g) + Br_2(l) \longrightarrow CH_2BrCH_2Br(l)$

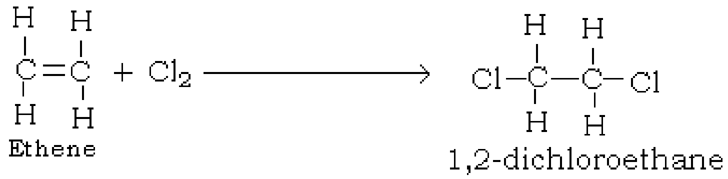
This is the common test for unsaturation.

ii) Acidified potassium manganate (VII) solution (purple) is decolorized if ethene is bubbled through the solution



The above is alternative test for unsaturation.

When chlorine gas is mixed with ethane and the mixture left in sunlight, the two combine forming a colorless oily liquid called 1,2-dichloroethane



c) Hydrogenation (addition of hydrogen)

When hydrogen and ethane mixture is passed over a finely divided nickel catalyst which is heated to about 200^{0}C ethane is formed

$CH_2=CH_2(g) + H_2(g) \longrightarrow CH_3CH_3(g)$

d) Reaction with sulphuric acid

Ethene undergoes an addition reaction with fuming concentrated sulphuric acid to form an oily liquid called ethyl hydrogen sulphate

$CH_2=CH_2(g) + H_2SO_4(l) \longrightarrow CH_3CH_2HSO_4(l)$

When the liquid formed above is boiled with water, ethanol is produced

$CH_3CH_2HSO_4\ (l) + H_2O(l) \longrightarrow CH_3CH_2OH(l) + H_2SO_4(l)$

e) Polymerization

Ethene under a very high pressure becomes a liquid. When this liquid is strongly heated to about 200^{0}C in the presences of a little oxygen catalyst, a white waxy solid (polyethene) is obtained.

$$nCH_2{=}CH_2 \longrightarrow \left[CH_2CH_2\right]_n$$

Uses of ethene

- it is used in the manufacture of ethanol
- it is used in the ripening of fruits
- it is used in the manufacture of plastics (synthetic polymers e.g. polythene)
- it is also used in preparing other solvents
-

Polymerization

Is the combination of many molecules of the same compound with relatively small molecular masses to form one complex molecule with very large molecular mass.

The complex molecule with a large molecular mass formed by the combination of many molecules of relatively small molecular masses is called the **polymer.** The small molecules from which a polymer is built are called **monomers**

Types of polymerization

These are mainly two i.e. addition and condensation

Addition polymerization

This is a combination of many small but unsaturated molecules to form a large molecule without any other product. In this case, the polymer posses the same empirical formula as the monomer. E.g. in the formation of polyethene

$$\underset{\text{(Monomer-ethene)}}{nCH_2{=}CH_2} \longrightarrow \underset{\text{(Polymer-polyethene)}}{\left[CH_2CH_2\right]_n}$$

Condensation polymerization

In the condensation polymerization, two different molecules combine to form one large molecule with consequent loss of simple molecules like water, hydrogen chloride etc.

so the empirical formula of the monomer and the polymer are not the same e.g. formation of starch from glucose and formation nylon 6,6

Types of polymers

Polymers can broadly be divided into two groups namely **natural polymers** and **synthetic polymers**

Natural polymers

Polymer	**Monomer**	**Use**
Starch	Glucose	Source of energy
Proteins	Amino acids	Repair of worn out tissues
Cellulose	Glucose	Cell walls
Glycogen	Glucose	Source of energy
Lipids (fats and oil)	Fatty acids and glycerol	Source of energy,
Natural rubber	Isoprene	Making foot wears

Synthetic (artificial) polymers

Polymer	**Monomer**	**Use**
Polyethene	ethene	-Making containers, eg plastics bowls and dust bins -used as wrapping materials
Polyvinyl chloride (PVC)	Chloro ethene	-Making rain coats -Electrical insulation -making pipes and films
Polystyrene	styrene	-making packing materials - making house hold items such as combs, plastic cups and a common lining in refrigerators

Classes of addition polymers

There are two main classes of addition polymers i.e. plastics and rubber.

Plastics

A plastic is a substance which when soft can be formed into different shapes

Plastics are minor products formed by cracking of crude oil eg poly ethene. Polyvinyl chloride. Melamine

All synthetic polymers are plastics in nature

Advantages of plastics

-They are good thermal and electrical insulators

- They can easily be shaped and molded (they are ductile)

- They resistant to acids and alkalis and they do not rust

-plastics can be colored when they are being manufactured and they do not need repainting

- They are light and therefore portable

-They are cheap

Disadvantages

- Produce poisonous fumes when they are burnt
- They are non biodegradable i.e they do not decay naturally
- Where serious fire hazards occur, molten plastics can inflict very severe burn

Types of plastics

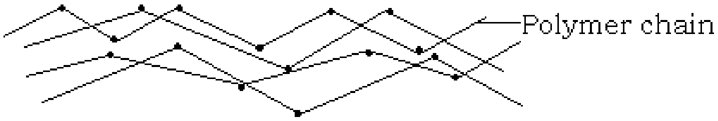
Plastics can be put into two types depending on their behavior upon heating i.e thermo- softening plastics and thermo-setting plastics

a) Thermo- softening plastics (Thermo-plastics)

These are plastics that soften or melt when heated and can be therefore be moulded into any shape while they are still soft. The plastics only harden when they cool.

Structure of thermo-plastics

The long polymer chains in thermoplastics lie along side each other. They may be entwined on each other but the polymer chains are not linked (not bonded to each other). When heated, the chains slide over each other making them soft and runny.



Examples of thermoplastics

1. Polythene

Polythene is a polymer of ethene. There are two types of polythene i.e. low density polyethene and high density polythene.

i) Low density polythene

This is made by polymerizing ethene at a high pressure of 1000-2000 atmospheres and temperature of 200˚C. Oxygen is used as a catalyst. It has a lower softening temperature of 105˚C-120˚C. The low density is due to poor packing of the branched polymer chains.

The low density polythene is soft, light and flexible

Uses:

For making polythene bags; insulation of electric cables because they can withstand bad weather conditions; making of squeeze bottles such as wash bottles; making plastic bags.

Disadvantage:

At boiling water temperature, they become soft so much that they become flappy and lose shape.

ii) High density polythene

It is made by polymerizing ethene at low pressure (5-25 atmospheres) and low temperature (20-50˚C) in the presence of a Ziegler catalyst. It has a higher softening temperature of about 140˚C. The high density is due top the close packing of the unbranched polymer chains. Very few of these polymers may be branched.

They are much harder and stiff and do not lose shape at boiling water temperature.

Uses

For making crates e.g. of beer and sodas, bowls, toys, buckets, food boxes, e.t.c.

2. Polyvinyl chloride (PVC)/Polychloroethene

PVC is made by polymerization of vinyl chloride (chloroethene).

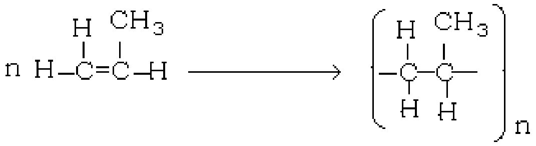
$$nCH_2{=}CHCl \longrightarrow \left(CH_2CHCl \right)_n$$

(Monomer-chloroethene) (Polymer-polychloroethene)

PVC are more rigid than polyethene and are used for making water pipes, light switches and sockets, insulation for electric cables,, carpets, plastic rain coats e.t.c.

3. Polypropene

This is made by polymerizing propene at a high pressure in the presence of a Ziegler catalyst.



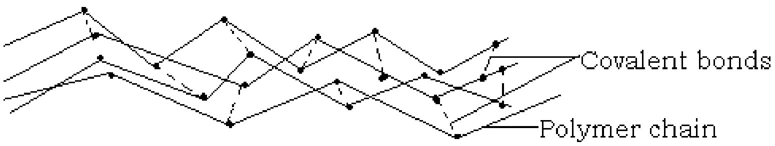
$$n\, H{-}\overset{H}{\overset{|}{C}}{=}\overset{CH_3}{\overset{|}{C}}{-}H \longrightarrow \left[-\underset{H}{\underset{|}{\overset{H}{\overset{|}{C}}}}-\underset{H}{\underset{|}{\overset{CH_3}{\overset{|}{C}}}}- \right]_n$$

It is used for making rope and for packaging.

b) Thermosetting plastics

These are plastics which do not soften or melt on heating and therefore cannot be remoulded into different shapes once they are set. They simply decompose upon heating. Thermosetting plastics have polymer chains which are bonded/ linked to each other. This is called cross linking.

Structure



Examples of thermosetting plastics include: Bakelite (used for making electric plugs, sauce pan handler, switches); melamine (used for making cups and children dishes).

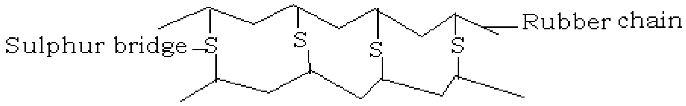
Natural rubber

Natural rubber is obtained from a rubber tree as a milky liquid called latex. Latex can be coagulated by addition of a little ethanoic acid to form a solid of high molecular weight.

The monomer of rubber is isoprene (2-methylbuta-1,3-diene)

Vulcanization of rubber

Rubber in its natural state is not strong or elastic enough and it is made more strong and useful by vulcanization which involves heating the rubber with sulphur. The sulphur combines with rubber forming cross linkages between natural rubber chains.



Vulcanized rubber is stronger, more elastic and more durable.

Uses of vulcanized rubber

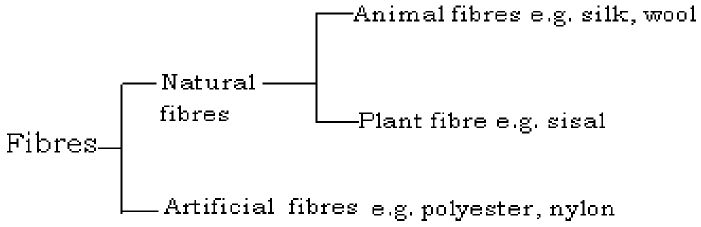
- It is used in the manufacture of tyres
- Used in the manufacture of foot wears

Condensation polymers

Fibres

These are polymers which can be drawn into threads. This is because, the forces of attraction between the linear molecules are weak but those between individual atoms are strong.

Classification of fibres



Advantages of synthetic/artificial polymers

- Relatively low production cost compared to the cost of extracting natural polymers.
- They are usually stronger and more resistant to corrosive substances like acids compared to natural polymers.
- They can easily be modified depending on the purpose for which the polymer is required unlike natural polymers which are hard to modify. As well their quality can easily be improved in terms of appearance, strength e.t.c.

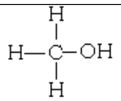
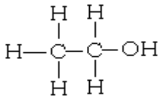
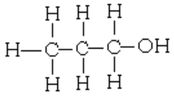
Disadvantages of synthetic polymers

- Many are non biodegradable causing pollution to the environment.
- When burnt, they produce toxic gases like hydrogen cyanide (from polypropenenitrile) thus endangering lives of the people working in the factories.

Alcohols/Alkanols

These are organic compounds with hydroxyl (-OH) group attached to the hydro carbon. Alcohols have a general formula of $C_nH_{2n+1}OH$.

Members of the series

N	Molecular formula	Structural formula	Name
1			Methanol
2			Ethanol
3			Propanol

Alcohols are named by placing –ol in the place of –e in the corresponding alkane members.

Ethanol

Physical properties

- It is a colourless liquid with a strong characteristic smell

- It is a volatile liquid and boils at 78˚C
- It is very soluble in water

Chemical properties

1. Combustion

Ethanol burns completely in air with a blue non luminous flame producing carbon dioxide and water vapour.

$C_2H_5OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$

2. Dehydration

When a little concentrated sulphuric acid is added to ethanol, an oily liquid called ethyl hydrogensulphate is produced and the reaction is exothermic.

$C_2H_5OH(l) + H_2SO_4(l) \longrightarrow C_2H_5HSO_4(l) + H_2O(l) + \text{Heat}$

When concentrated sulphuric acid is heated with ethanol, it produces ethene.

$CH_3CH_2OH(l) \xrightarrow[170^\circ C - 180^\circ C]{\text{Conc. } H_2SO_4} C_2H_4(g) + H_2O(g)$

Preparation of ethanol

Ethanol is manufactured/ prepared by the process of fermentation of carbohydrates such as starch and sugars.

Fermentation

This is a process in which carbohydrates like starch and sugars are converted to alcohol by enzymes. The enzymatic break down of glucose yields simple compounds like ethanol and carbon dioxide. Some heat is as well generated. Fermentation takes place in the absence of oxygen (anaerobic process).

Preparation from starch

Starch is heated with malt at a temperature of 60˚C. Malt contains an enzyme diastase which hydrolyses starch to maltose.

$$2C_6H_{10}O_5(aq) + H_2O(l) \longrightarrow C_{12}H_{22}O_{11}(aq)$$

(Starch) (Maltose)

Yeast is added at room temperature to the mixture and left to ferment for 2-3 days. Yeast contains two enzymes, maltase and zymase. Maltase catalyses the hydrolysis of maltose to glucose as below.

$$C_{12}H_{22}O_{11}(aq) + H_2O(l) \longrightarrow 2C_6H_{12}O_6$$

(Maltose) (Glucose)

Zymase catalyses the breakdown of glucose into ethanol, carbon dioxide, producing heat in the process.

$$C_6H_{12}O_6(aq) \longrightarrow 2C_2H_5OH(l) + 2CO_2(g) + \text{Heat}$$

(Glucose) (Ethanol)

The crude ethanol produced can then be concentrated or purified by fractional distillation.

Preparation of ethanol from millet

- Millet flour is mixed with little water to form paste. The mixture is then put under ground for about 8 days.
- It is then removed, roasted and dried under the sun.
- The dried material is then mixed with germinated millet flour (yeast).
- Water is added and the mixture allowed to ferment for about 3 days in a warm place. This forms a local drink known as “Malwa”.

Preparation of ethanol from ripe bananas

- Ripe bananas are squeezed to obtain the juice.
- The juice is filtered to remove the solid particle.
- The juice is mixed with roasted sorghum flour and the mixture allowed to ferment for 1-3 days in a warm place. A crude form of ethanol locally known as “Tonto” is obtained.

Beer is made by the fermentation of the starch in barley; wine by the fermentation of sugars in grapes. Spirits are obtained by distillation of dilute solutions produced by fermentation and there fore have an increased alcoholic content.

Uses of ethanol

- It is used as an alcoholic beverage e.g. beers, wines and spirits
- It is used as a solvent for paints, varnishes e.t.c
- It is used as a fuel
- It is used as a preservative and for sterilization
- It is used as a thermometric liquid especially in minimum and maximum thermometers.

SOAPS AND DETERGENTS

Soap

Soap is a sodium or potassium salt of a long chain carboxylic acid known as sodium or potassium stearate.

Manufacture of soap

The process of making soap using an alkali and fat/oil (ester) is known as **saponification**.

Boil vegetable oil (from coconut, ground nuts, cotton e.t.c) or animal fat (from cattle or sheep) with concentrated sodium hydroxide solution until a uniform solution is obtained. Allow the solution to cool. Concentrated solution of sodium chloride (brine) is added to precipitate the soap which floats on the surface. The process of precipitating the soap is known as "**salting out**". The soap is then removed and treated further to produce pure soap.

Perfumes may, dyes and disinfectants may be added to make toilet soap e.g. Geisha

General equation

$RCOOR^1 + NaOH \longrightarrow RCOONa + R^1OH$

Fat/oil + Alkali $\longrightarrow$ Soap + Glycerol

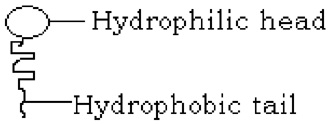
Note

1. Potassium hydroxide can be used instead of sodium hydroxide. Potassium soaps are normally milder and there fore used mainly as toilet soaps.
2. Oils are liquid esters at room temperature whereas fats are solids at room temperature.

Cleaning action of soap

Soaps and detergents act in a similar way to facilitate the cleaning process. They act by lowering the surface tension of water and thus enable the water to spread and wet more effectively i.e.to break up and disperse grease particles.

Dirt is fixed on objects by oil films. Soap has two parts i.e. the long hydro carbon tail that is soluble in oil but insoluble in water (hydrophobic tail) and a carboxylic acid head that is soluble in water (hydrophilic head) but insoluble in oil.



During washing, the hydrophobic tail dissolves in the oil film and the hydrophilic head remains in water, this crates tension making the grease particles to split up into tiny globules which are carried away by water. The dirt particles get suspended in water, a process known as **emulsificationSoapless (synthetic) detergents**

A detergent is any substance that facilitates the cleaning process. This means that soap is also a detergent although the name is used for other substitutes of soap like Omo, Nomi, Ariel, Toss e.t.c.

The synthetic detergents function in the same way as soap but they are more soluble than soap and there fore clean more effectively. Even when hard water is used, they do not form scum but soap does.

The soapless detergents are made from concentrated sulphuric acid and hydrocarbons obtained from petrol refining.

Laboratory preparation of a soapless detergent from castor oil

Procedure

- Add 1cm^3 of castor oil into a test tube, then carefully add 2cm^3 of concentrated sulphuric acid while stirring with a glass rod
- Gently warm the mixture and add about 10cm^3 of 4M sodium hydroxide and stir. The mixture gets hot, viscous and dark.
- Add 5cm^3 of distilled water and stir. Then decant to separate the liquid from the solids. The solid is the soapless detergent which is then washed with distilled water.

Advantages of soapless detergents

- They are more soluble in water than soap and there fore clean more effectively.
- They do not form scum with hard water there fore can be used with both hard and soft water. Soap forms scum with hard water.

Disadvantages of soapless detergents

- It is more expensive than soap
- Some soapless detergents are non biodegradable and there fore accumulate in the environment. Soap is biodegradable.
- The phosphates from soapless detergents when washed in to water bodies' causes eutrophication. This leads to pollution of water bodies.

Sample questions on Organic chemistry

1. *Ethene can undergo polymerization. Explain what is meant by the term "polymerization of ethene". Name the product of polymerization of ethene and write equation for the reaction leading to the formation of the product that you have named. State one use of the product you have named.*

   *On polymerization, ethene formed a compound T, molecular mass = 16,660. Determine the number of moles of ethene molecules that combined to form T. (C = 12, H = 1).(**Ans.=595 moecules**)State the term which is used to describe a single unit of the ethene molecule in T.*

Distinguish between the terms "Synthetic polymer" and "natural polymer", and use silk and nylon to match with the type of polymer that you have distinguished. State one use each of silk and nylon. State (i) one characteristic property of thermosetting plastics and thermoplastics. (ii) one example each of thermosetting plastics and thermoplastics.

2. Define the following terms as appied to organic chemistry: isomerism, polymerization, a hydrocarbon, homologous series, functional group, cracking, hydrogenation of ethene and vulcanization of rubber. Some organic compounds are said to be saturated and some unsaturated. Mention two examples of unsaturated and saturated organic compounds.
3. Give the names of each of the following compounds: CH_3OH; $CH_3CH_2CH_3$; CH_3CH_2COOH; $CH_2{=}CH_2$. Which of these organic compounds are gases at room temperature? Which of them is/are saturated and give a reason? Mention the monomer in the following polymers: polyethene, polyvinyl chloride and polypropene.
4. Discuss the reactions of a named alkane with oxygen and chlorine. Lower alkanes can be obtained by refining petroleum, describe the process of fuel refining.
5. Describe briefly the preparation of ethene from ethanol. How is ethene converted back to ethanol? When ethene is bubbled through bromine water and acidified potassium potassium permanganate, the bromine water and permanganate solutions turn colourless. Explain with the aid of equations. With examples, differentiate between addition and condensation polymerization. Outline four uses of ethene.
6. What are plastics? Describe in detail is meant by thermosetting and thermo softening plastics, use examples to illustrate. Explain the disadvantages of using plastics. Explain briefly why rubber has to be vulcanized and give the uses of vulcanized rubber.
7. Explain the term fermentation and describe how ethanol is obtained from a named substance by fermentation. Use equations to illustrate. Which other product is given out during the process? Outline at least three uses of ethanol.

8. *Explain briefly what is meant by (i) soap and (ii) detergent. Describe how soap and detergents are made. Explain the cleaning action of soap. Outline the advantages and disadvantages of soap less detergents.*
9. *Under what conditions does ethanol react with concentrated sulphuric acid to produce the colourless gas ethene* (C_2H_4)? *Write an equation for the reaction.*

   What is the reaction of ethene upon bromine water and upon hydrogen chloride. Give the names and formula of the products formed.
10. *Select one member member of each of the following homologous series: (i) alkenes and (ii) alkanols. In each case give its name , molecular formula and structural formula, two chemical reactions and two uses in every day life.*

APPLIED CHEMISTRY

METAL AND METAL EXTRACTION

Metals are electropositive i.e. they ionize by lose of electrons to form positively charged ions. They are therefore reducing agents.

Occurrence of metals

Metals occur in the earth crust as ores.

An ore is a rock which contains metal compounds from which metals can be extracted e.g. rock salt (NaCl) for sodium; spathic iron ore ($FeCO_3$) and iron pyrite (FeS_2) for iron.

Extraction of metals

For a metal to be extracted, its concentration in a given ore must be high. There fore, it is necessary to concentrate (purify) the ore before extracting the metal.

The method of metal extraction normally depends on the position of the metal in the electrochemical/reactivity series.

K Na Ca	These are very reactive metals and therefore form very stable compounds. They are not found as free elements and can only be extracted from their compounds by electrolysis.

These are moderately reactive metals. They usually occur as ores inform of oxides, sulphides, carbonates e.t.c. they are extracted by reduction of the ores using either carbon or carbon monoxide.

These are not very reactive. They are there for found as free elements. They

Mg
Al
Zn
Fe
Pb
Cu
Hg
Ag
Au

Extraction of metals as a reduction process

Metals in their ores (oxides or other salts) exist in the ionized condition. During extraction, metallic ions in the ores gain the necessary number of electrons to form the corresponding metal atoms. Since the process involves taking up electrons, extraction of metals is essentially a reduction process. For example:

$$Na^+(aq) + e^- \longrightarrow Na(s)$$

(Ion from ore e.g. sodium chloride) (Sodium metal atom)

(Ion from ore e.g. siderite) $\longrightarrow$ (Iron metal atom)

SODIUM

Principle/chief ores of sodium are:

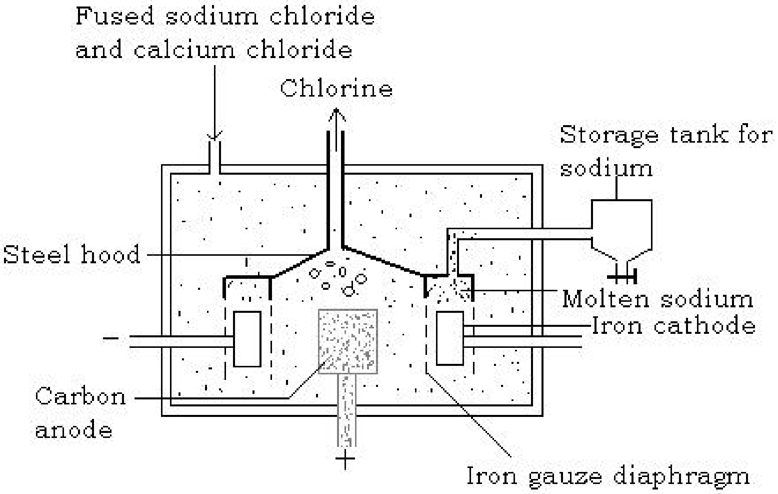
- Rock salt (sodium chloride), NaCl. This is common in salt lakes like lake Katwe in Uganda.
- Soda ash (sodium carbonate), Na_2CO_3. This is common at lake Magadi in Kenya.

- Sodium nitrate (Chile salt petre), $NaNO_3$. Common in chile.

Extraction process

Sodium is extracted by electrolysis of fused sodium chloride in the Down's cell using Downs's process. As the melting point of sidium chloride is high (about 800˚C), calcium chloride is added to lower the melting point to about 600˚C.

The Down's cell



The Down's cell has an outer iron shell lined with firebrick. Iron gauze cylinders separate the graphite anode from the ring shaped iron cathode; this prevents the mixing of the products at the different electrodes.

A high current (about 30000 A) is used to keep the electrolyte in a molten state.
Chlorine escapes via the hood. Sodium collects in the inverted trough placed over the cathode, rises up the pipe and is tapped off through the iron vessel. Sodium metal is collected upwardly in the Down' cell because of its low density that makes it to float over the reacting mixture.

Reaction at the cathode

At the cathode, sodium ions gain electrons and therefore are discharged to form sodium metal which is deposited in the molten state and collected over dry nitrogen.

$$Na^{+}(aq) + e^{-} \longrightarrow Na(s)$$

Reaction at the anode

At the anode chlorine ions are discharged as they lose electrons to form chlorine atoms. The chlorine atoms pair up forming chlorine molecules (gas) which is collected as a valuable product.

$$2Cl^{-}(aq) - 2e^{-} \longrightarrow Cl_2(g)$$

Uses of sodium

- It is used in the manufacture of sodium cyanide (NaCN) which is used in the extraction of gold.
- Manufacture of sodamide ($NaNH_2$)
- Manufacture of sodium peroxide (Na_2O_2)
- Its alloy is used in the manufacture of anti-knock additives for petrol e.g. tetraethyllead.
- Sodium is used for the withdrawal of heat from some nuclear reactors. It is therefore used as a coolant.
- Sodium is used as a reducing agent in the laboratory.
- Sodium vapour lamps are used for street lighting.

Extraction of intermediate metals

The metals are extracted by reduction and there have four basic stages in the process.

1. Concentration of the ore
2. Roasting of the ore to convert the carbonates and sulphides to oxides and remove water vapour. It is easier to reduce the oxide than the carbonate.
3. Reduction of the ore
4. Purification/refining of the metal extracted

Concentration of the ore

This process involves removing impurities. Impurities can be separated by physical means e.g. picking by hand, washing, using a magnet or solvent extraction. In the case of copper, the ore is first crushed and then mixed with water. Air is then blown through the mixture and the clean ore separates and collects at the surface as froth. This is referred to as froth floatation.

Roasting

The concentrated (pure) ore is roasted in air at high temperatures to produce oxides which are easier to reduce than carbonates or sulphides. Sulphur is removed as sulphur dioxide and carbon dioxide driven off from the carbonates.

Reduction of the ore

The ore is usually heated in a furnace in the presence of a suitable reducing agent, usually coke (carbon) or carbon monoxide which converts the ore into the required metal.

Purification/refining of the metal

The metal obtained is often impure. Purification is normally done by electrolysis in the case of copper and zinc. In other cases, the impure metal is heated in a hearth open to air where the impurities oxidize and rise to the surface as a scum and can be removed.

IRON

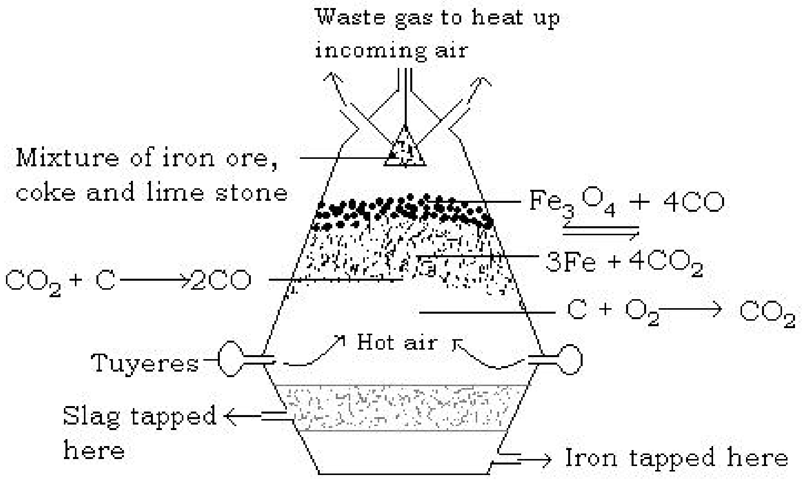
The principle/chief ores of iron are: Haematite, Fe_2O_3; Magnetite, Fe_3O_4; Iron pyrite, FeS_2; Siderite or Spathic iron ore ($FeCO_3$); Limonite ($Fe_2O_3.xH_2O$).

Extraction process

The iron ore is crushed and roasted in air to remove water and other non metallic impurities especially sulphur and phosphorus which are oxidized away as gaseous oxides. The roasted ore is now mainly Fe_3O_4. When an ore is roasted in air and Fe_3O_4 is the main product, it is known as sintering

The roasted ore is mixed with coke (carbon) and limestone (calcium carbonate) and introduced into the blast furnace where the reduction of the ore takes place.

The blast furnace



Very hot air is introduced from low down into the blast furnace. As the hot air passes through the mixture of roasted ore, coke and limestone, the coke burns to form carbon dioxide (is oxidized) in an exothermic process.

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

As the carbon dioxide produced rises through the furnace, it is reduced by the excess hot coke to produce carbon monoxide.

$$CO_2(g) + C(s) \longrightarrow 2CO(g)$$

The carbon monoxide formed reduces the iron ore at a high temperature (about 1000˚C) to form iron metal.

$4CO(g) + Fe_3O_4(s) \longrightarrow 3Fe(s) + 4CO_2(g)$

The molten iron formed sinks to the bottom of the blast furnace where it is tapped and solidified into blocks of pig iron.

The role of limestone

Limestone removes Silicon(IV) oxide which is the main impurity in the iron ore. Limestone at high temperature decomposes to form calcium oxide and carbon dioxide. The calcium oxide formed combines with silicon (IV) oxide (impurity) to form molten calcium silicate (slag).

$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$

$CaO(s) + SiO_2(s) \longrightarrow CaSiO_3(s)$

The slag being less dense than iron forms a separate layer above iron and thus are tapped separately.

The slag protects the molten iron against any further oxidation by oxygen in the hot air in the blast furnace. The carbon dioxide produced as a bi-product in the furnace is quickly reduced by hot coke to carbon monoxide which is required as a reducing agent.

By products from the blast furnace include: calcium silicate/slag (used for making roads, manufacturing cement, manufacturing glass); waste gases e.g. carbon dioxide, carbon monoxide, steam and unreacted nitrogen; calcium phosphate (used as an in organic fertilizer).

There are three types of iron which are classified based on their percentage purity. The percentage purity also determines the strength and use of the iron. The types of iron are cast iron (pig iron), wrought iron and steel.

Cast iron

This is an impure iron which contains relatively high proportions of carbon (4%) and small proportions of other substances such as silicon, phosphorus and sulphur. Such impurities make cast iron to be hard, brittle and to have a lower melting point than pure iron. It cannot be welded and has little tensile strength.

Cast iron can be used to make hot water pipes, Bunsen burner bases, cookers, in railings and other purposes where little strain is imposed.

Wrought iron

This is the purest form of iron (contains about 0.3% carbon) and is obtained from cast iron by heating it with iron(III) oxide in a furnace by a process known as “puddling”. The oxygen of the iron oxide oxidizes carbon and sulphur to their respective gaseous oxides, phosphorus to Phosphates(V) and silicon to silicates which form slag. The semi molten mass is then hammered and rolled so that the slag is squeezed out and a mass of almost pure iron remains.

It is very tough, malleable and ductile and is there fore used to make iron nails, sheeting, ornamental work, horse shoes and agricultural implements. Wrought iron is some times referred to as low carbon steel.

Steel

Steel is an alloy of mainly iron with carbon and other elements like manganese, chromium, silicon, cobalt and some time tungsten. The quality of steel depends on the amount of carbon present and this in turn determines its intended use.

Steel is generally used in the construction of buildings, bridges, ships, car bodies, cutting and boring tools, crushing machines and stainless cutlery such as knives, forks e.t.c.

COPPER

The principle/ chief ores of copper are: copper pyrites ($CuFeS_2$); cuprite (Cu_2O); copper (I) sulphide (Cu_2S) and malachite ($CuCO_3,Cu(OH)_2$).

Extraction of copper from copper pyrites

The ore is first concentrated by a process of froth floatation and then it is roasted in air to produce copper(I) sulphide.

$$2CuFeS_2(s) + 4O_2(g) \longrightarrow Cu_2S(s) + 3SO_2(g) + 2FeO(s)$$

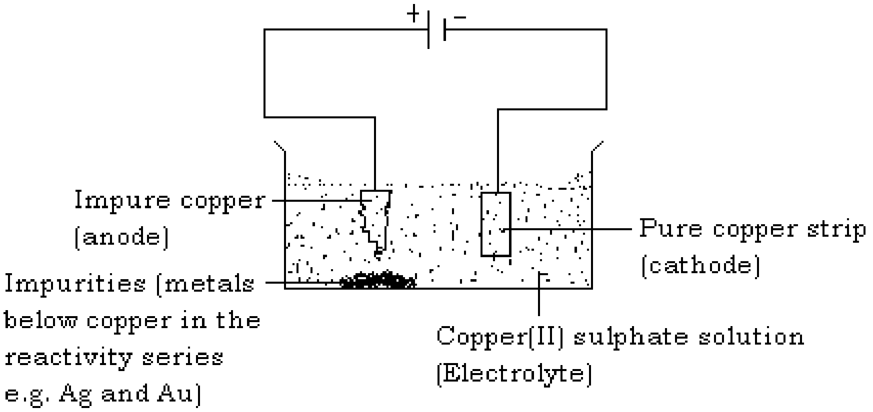
By adding silicon dioxide (sand) and heating the mixture in the absence of air, the solid iron(II) oxide impurity is converted into slag which is poured off with the reaction mixture leaving behind only copper(II) sulphide.

The copper(II) sulphide is reduced to metallic copper by heating in a regulated supply of air (oxygen).

$$Cu_2S(s) + O_2(g) \longrightarrow \underset{\text{(Impure copper)}}{2Cu(s)} + SO_2(g)$$

The metallic copper produced is impure copper (blister copper) and has to be purified. Purification of the copper is done by the process of electrolysis.

Purification of blister copper



The impure copper/blister copper is made the anode and strip of pure copper serves as the cathode. The electrolyte is acidified solution of copper(II) sulphate. During electrolysis, pure copper is transferred from the impure copper anode to the pure copper cathode. There fore, the anode dissolves and decreases in size as the cathode grows bigger.

At the anode (impure)

$Cu(s) - 2e^- \longrightarrow Cu^{2+}(aq)$

At the cathode (pure)

$Cu^{2+}(aq) + 2e^- \longrightarrow \underset{\text{(Pure copper)}}{Cu(s)}$

Ions above copper in the reactivity series remain in the solution.

Uses of copper

- For wiring electrical circuits since it is a good conductor of electricity and is relatively cheap. The copper used for this purpose must be very pure since impurities increase electrical resistance.
- It is used for making ornaments like ear rings and pins, bungles e.t.c. being little attacked by air.
- It is used for making alloys like bronze (copper and tin) and brass (copper and zinc),copper coinage (copper and tin), German silver (copper, zinc and nickel)
- Used for making water pipes and boilers.
- Copper is used as a roofing material because it weathers to acquire a coating of green basic copper carbonate, $CuCO_3.Cu(OH)_2.nH_2O$, which lends a colourful touch to the building.

FERTILIZERS

A fertilizer is a substance containing essential elements for the healthy growth of plants. When added to the soil, fertilizers normally improve soil fertility. The most important elements are: nitrogen (N), potassium (K) and phosphorus (P). Some times sulphur.

These elements are absorbed by plants as compounds especially as nitrates and phosphates. There are mainly two types of fertilizers i.e. artificial/inorganic fertilizer and organic fertilizers.

Artificial fertilizers

These are synthetic fertilizers that contain a high percentage of essential elements i.e. nitrogen, phosphorus and potassium; the fertilizers should be cheap and soluble in water. The examples of such fertilizers include: ammonium sulphate, ammonium phosphate, potassium nitrate, liquid ammonia and potassium chloride.

Note

1. Liquid ammonia is very soluble; it can easily be washed away by rain water and can easily evaporate off the soil.
2. A composite/balanced fertilizer contains nitrogen, phosphorus and potassium in certain correct proportions and it is normally labeled NPK.
3. Excessive use of ammonium fertilizers renders the soil acidic. It is there fore necessary to add quick lime (calcium oxide) to neutralize the acidity. The added lime also makes the soil more porous for proper air and water absorption.

Organic fertilizers (Farm yard manure)

These are fertilizers from decomposed animal and plant waste matter such as cow dung, chicken droppings, dead animals, dead leaves, plant branches, coffee husks e.t.c.

Advantages of organic fertilizers

- It is cheap since it is prepared by the farmer
- No elaborate instruction for application
- They provide a suitable habitat for nitrifying bacteria
- They last for a long time in the soil thus improving soil texture
- They hold water thus keeping the soil moist even in dry season

Disadvantages

- They cannot be used selectively i.e. on soil which lacks a particular nutrient
- They harbor pests and diseases which attack the crop
- They have to be applied in bulk for effectiveness
- They act slowly and are not ready for use any time since they have to be dry first.

Advantages of artificial fertilizers

- They do not habour pests and diseases
- They are available for use whenever needed
- They are effective when used in small amounts
- Different fertilizers for different crops and soil types are usually available
- They normally give good yield

Disadvantages

- They are expensive in terms of production costs
- They do not give desired results if the instructions for application are not strictly followed
- They may exhaust soil fertility if used for a long time
- The crop/yield is of inferior quality to one produced using organic fertilizers.

Biogas production

Biogas is a gas with a high methane content which is produced by microbial fermentation of organic wastes.

Biogas is formed by bacterial activities on animal and vegetable wastes. It consists of mainly methane and others gases like; ammonia, hydrogen sulphide, carbon dioxide e.t.c.

A simple biogas generator consists of a container in which animal and vegetable wastes are mixed with a correct (limited) amount of water and then covered to prevent atmospheric oxides. A temperature of about 25˚C to 30˚C is maintained. The bacteria present in the wastes break down (decompose) the waste to form biogas.

The quality of biogas produced depends on the type of waste used. For instance, a mixture of cow dung, human excreta and bean stalks produce high quality biogas.

The biogas produced is mainly used as fuel for cooking purposes and lighting.

Advantages related to biogas production

- Biogas is easy and cheap to produce
- Sewage materials can be converted to biogas
- The solid waste from a biogas plant can be used as an organic fertilizer since it contains a high percentage of nitrogen.
- Forests and wild life would be conserved if institutions and rural communities used biogas instead of wood and charcoal for their fuel needs.

However the main disadvantage related to biogas production is the release of gases such as sulphur dioxide, carbon dioxide which pollute the air.

Sugar

Sugar is a food or drink sweetener.

Extraction of sugar

In Uganda, sugar is made from sugarcane plant. The sugarcane is cut into small pieces, then crushed and squeezed to force out the juice which contains mainly sucrose. A little lime is added to prevent the sucrose from hydrolyzing into simple sugars like glucose. The mixture is filtered and the clear filtrate obtained is diluted with correct amount of water. The diluted filtrate is then concentrated by evaporation to crystal formation point. The solution is then allowed to cool as sugar crystals form. This forms brown sugar crystals.

To obtain white sugar crystals, the brown sugar is dissolved in water and the solution boiled with animal charcoal. The mixture is filtered to remove the charcoal and the filtrate evaporated as before to form white sugar crystals.

Sample questions on metal extraction, Biogas and fertilizers

1. *Define the term ore. Explain why the extraction of metal is regarded as a reduction process. Name one principle ore of sodium and how sodium is extracted from the ore by electrolysis method. Write equations for the reactions at the electrodes. Outline five uses of sodium.*
2. *Name the principle ores of iron. Describe how iron is extracted from a named ore.*
3. *Copper can be extracted from copper pyrite. Describe how copper can be extracted from copper pyrite. Mention the uses of copper. Outline how copper is refined by electrolysis.*
4. *Describe the processes involved in the extraction of sugar from sugar cane. What is meant by biogas and how is biogas produced. Outline the advantages related to biogas production.*
5. *What is a fertilizer? Differentiate between the artificial and organic fertilizers. Give the advantages and disadvantages of artificial and organic fertilizers.*