UNIT 1: INTRODUCTION TO CHEMISTRY.

Unit Checklist.

Meaning of chemistry

- Matter
- States of matter
- Properties of the different states of matter.
- Conductors and non-conductors

Role of chemistry in society.

Laboratory rules.

- Necessity for laboratory rules.
- The laboratory rules and regulations.
- Laboratory safety symbols.

Common apparatus used in a chemistry laboratory.

The Bunsen burner.

- Structure
- Functions of the various parts.
- Types of Bunsen burner flames.

Methods of gas collection.

- Upward delivery
- Downward delivery
- Over water
- Using a collecting syringe.

Drying of gases

- Using concentrated sulphuric acid.
- Using anhydrous calcium chloride
- Using calcium oxide.

Drugs and drug abuse.

Meaning of chemistry.

- It is a branch of science which deals with the composition and properties of matter.

Matter

- Matter is anything that has mass and occupies space.

States of matter.

- Matter exists in three different states: solids, liquids and gases.

Properties of the different states of matter.

	Solid	**Liquids**	**Gases**
Mass	Definite	Definite	Definite
Shape	Definite	**Indefinite:** they take the volume of the container in which they are in.	**Indefinite**: they take the shape of the container in which they are.
Volume	Definite	Definite	**Indefinite**: volume may increase with increase in temperature; and decrease with decrease in pressure

Note: Conductors and non-conductors:

- The flow of electric current through materials is called electrical conductivity.
- Solid substances which allow electric current to flow through them are called **conductors**.
- Solid substances that do not allow electric current to flow through them are called **non-conductors**.

Role of chemistry in a society.

- Chemistry has enabled extraction of chemicals from plants.
- It is used in the manufacture of substances such as soap, glass, plastics, medicine, rubber, textiles etc from naturally occurring substances.
- Purification of substances from natural raw materials.
- It forms a basis for entry into careers e.g. teaching, medicine, chemical engineering etc.

Laboratory rules and safety symbols.

Necessity of laboratory safety rules.

1. To avoid accidents and injuries during practical experiments in the laboratory.
2. To avoid damage to and breakage of apparatus and laboratory fittings.
3. To avoid wastage of laboratory chemicals.

The laboratory safety rules and regulations.

- Never run while in the laboratory;

Reason: You may injure others or yourself in the laboratory.

- Never taste or eat something in the laboratory;

Reason: to avoid poisoning.

- Always consult your teacher before trying out any experiment; so as to avoid accidents.
- Label all chemicals in use so as to avoid confusion.
- Always use a clean spatula for scooping a substance from a container to minimize contamination.
- Always hold test-tubes and boiling tubes using test tube holder when heating; to avoid being burned.
- When heating a substance never let the open end of the tube face yourself or anybody else, because the liquid may spurt out and cause injury.

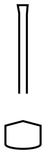
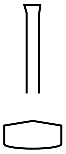
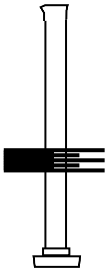
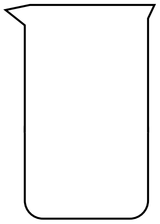
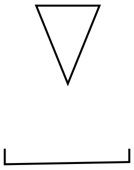
- Never look directly into flasks and test tubes where reactions are taking place, because the chemicals may spurt into your eyes and cause injury.
- Never smell gases directly. Instead, waft the gaseous fumes near your nose with your hand.
- Experiments in which poisonous gases and vapours are produced must be carried out in a fume cupboard or an open space outdoors.
- Always keep flammable substances away from flames because they easily catch fire.
- Always report any accidents to the teacher or the laboratory technician immediately for necessary action.
- In case of an accident do not scramble for the same exit, because it may hinder easy escape.
- Always put off flames that are not in use in order to avoid accidents and minimize fuel wastage.
- If a chemical gets on your skin or mouth rinse it immediately with a lot of water.
- Always dispose off the chemicals already used safely to avoid explosions and contaminations.
- Always work on a clean bench. After completing your experiment, clean all the pieces of apparatus you have used and return them to their correct storage places.
- Always read the label of the reagents before using them.

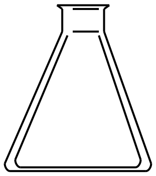
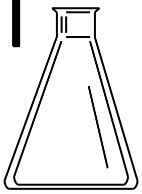
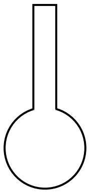
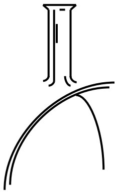
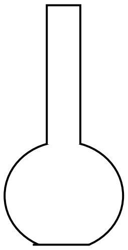
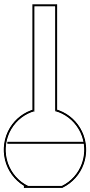
Safety symbols.

- These are signs found on the labels of bottles or cartons containing dangerous chemicals.
- The common safety symbols are as follows:

Symbols	**Meaning.**
	Toxic: are very poisonous and can easily kill if swallowed, inhaled or on contact with the skin. **Examples:** Chlorine and mercury;
	Harmful: Less harmful (dangerous) than the toxic substances; are only likely to cause pain and discomfort. **Examples:** copper (II) sulphate, lead (II) oxide
	Highly flammable: are substances that catch fire easily and must not be handled near open fire. **Examples:** ethanol, hydrogen
	Oxidizing: rapidly provide oxygen and can cause fire to burn more fiercely. **Examples:** potassium manganate (VII), hydrogen peroxide
	Corrosive: are substances that cause burns to skin and fabric; and can also react with other substances such as metals **Examples:** nitric (V) acid, conc. sulphuric acid, bromine.
	Irritant: can cause blisters or reddening of the skin; usually irritate the respiratory tract, skin, eyes etc. **Examples:** calcium chloride and zinc sulphate

Common Chemistry laboratory apparatus and their uses.

Name of apparatus	**Diagram.**	**Use.**
Test tubes		- General laboratory experiments; like heating solids; qualitative analysis etc
Boiling tubes		- Mainly used for heating small amounts of solids and liquids.
Test tube holder		Used for holding test tubes and boiling tubes during heating experiments.
Measuring cylinder		- Measuring accurate volumes of liquids in the laboratory
Beaker - Lipped glass or plastic vessels of various capacities.		- Glass beakers are used for boiling liquid substances; - Holding solutions during chemistry experiments.
Filtering funnel		- Directing liquids into containers with small narrow mouths; - Holding filter papers during filtration;

Name of apparatus	**Diagram**	**Use**
Stirring rod		
Watch glass		
Thermometer		- Measuring temperatures during experiments.
Conical flask	 	- Normal laboratory experiments like titration. - May be used for measuring volumes if graduated.
Round-bottomed flask	 	- Used when heating liquid substances because heat is supplied uniformly.
Flat-bottomed flask	 	- Used for general laboratory experiments.
Evaporating dish		- Used when evaporating liquids.

Name of apparatus	**Diagram**	**Use**
Crucible		- Use when heating solid substances that require strong heating.
Pestle and mortar		- Crushing substances while the mortar holds the substances being crushed.
Pie clay (ceramic) triangle		- Supporting crucibles during heating.
Tripod stand		- Supporting beakers and flasks in which liquids are being heated.
Wire gauze		- Used when glass apparatus are being heated; to facilitate even distribution of heat when heating substances in beakers or flasks
Clamp		- Supporting and holding pieces of apparatus during experiments.

Name of apparatus	**Diagram**	**Use**
Deflagrating spoon		- Holding burning substances.
Spatula		- Scooping solid substances from containers
Crucible tongs		- Holding solid chemicals.
Condenser		
Separating funnel		- Separating immiscible liquids.
Thistle funnel		- Delivering liquid substances into other containers like flasks during reactions.

Name of apparatus	**Diagram**	**Use**
Wash bottle		- Holding water for rinsing apparatus
Dropping funnel		
Test tube rack		- Holding boiling tubes and test tubes.
Teat pipette (dropper)		- Sucking liquid chemicals and placing them in another container dropwise.
Burette - It consists of a long narrow tube with a tap and a jet at the bottom.		- Delivering accurate volumes of liquids

Name of apparatus	**Diagram**	**Use**
Pipette		- Delivering a specified volume of liquid accurately.
Gas jar		- Gas collection.
Trough		- Holding some amount of water for some experiments e.g. gas preparation.
Reagent bottles		- Storing chemicals in liquid state.
Desiccator		- Drying substances or keeping substances free from moisture.

Name of apparatus	**Diagram**	**Use**
Spirit lamp **Note:** less preferred for heating because their flames are not hot enough; and they deposit soot on apparatus making them dirty hence difficulty in observing changes during experiments		- Heating substances in the laboratory.
Kerosene stove **Note:** less preferred for heating because their flames are not hot enough; and they deposit soot on apparatus making them dirty hence difficulty in observing changes during experiments		- Heating substances in the laboratory.
Electric heater		- Heating substances in the laboratory.
Candles **Note:** less preferred for heating because their flames are not hot enough; and they deposit soot on apparatus making them dirty hence difficulty in observing changes during experiments		- Heating substances in the laboratory.
Bunsen burner		- Heating substances in the laboratory.

Name of apparatus	**Diagram**	**Use**
Stop watch (clock)		- Measuring time particularly in determination of reaction rates.
Beam balance		- To take accurate weight measurements
Electronic balance		- Take accurate weight measurements; and can take extremely low weight measurements.
Volumetric flask		

Note: most chemistry apparatus are made of glass:
Advantages of glassware:
- It does not react with most chemicals
- Glass is transparent and hence reactions can easily be observed as they progress.
- Glass materials are easy to wash and rinse after experiments.
- They can be used comfortable in heating experiments.

Disadvantages:
- Have higher chances of breakages in case they fall during experiments.
- They are comparatively expensive to plastics

- Some materials like beakers may however be made of plastics.

Advantages of plastic apparatus.

- Have lower chances of breaking.
- They are relatively cheaper to buy.

Disadvantages:

- Plastics tend to react with some laboratory chemicals
- may not be transparent and hence reactions cannot easily be observed as they progress.
- Glass materials are difficult to wash and rinse after experiments.
- They cannot be used in heating experiments.

The Bunsen burner.

- is the most common heating apparatus in the laboratory.
- Was invented by a German scientist known as Wilhelm Bunsen hence the name.
- It uses natural (cooking) gas for heating.

Parts of a bunsen burner

- Chimney
- The air hole
- Collar
- The jet
- Base
- Gas inlet

Diagram: structure of a bunsen burner.

Diagram: Bunsen burner-components detached

Functions of the various parts:

- **The chimney:**

- It is a hollow metallic cylinder with an air hole near its lower end.

Function:

- allows air and the laboratory gas from the jet mix before they start burning at the upper end of the chimney.
- Its upper opened end provided a site where the gas burns.

- **The air hole:**

- Is a small aperture found at the lower end of the chimney and smaller than the collar.

- Its diameter (size) is regulated by the collar.
Function:
- Allows air (oxygen) to enter and mix with the laboratory gas in the chimney.

- **The collar:**

- Is a metallic ring with an air hole whose diameter is the same size as that of the air hole in the chimney.
- It fits into the lower part of the chimney; and can rotate around the chimney opening or closing the air hole.
Function:
- Regulates the amount of air entering the chimney.

- **The jet:**

- It is a very tiny opening just below the air hole, that connects the gas inlet to the chimney.
Function:
- allows the laboratory gas (methane) into the chimney at high pressure.

- **The gas inlet:**

- It is a hollow metal connected to the base and extending into the jet.
- Its external opening is usually connected to a rubber tubing that is connected to a gas tap.
Function:
- Allows laboratory gas from the cylinders (reservoirs) in the lab; through the gas taps into the Bunsen burner.

- **The base:**

- A thick heavy metal, that is usually circular or oval.
Function:
- It supports the Bunsen burner on the bench.

The Bunsen burner flames.
- A flame is a mass of burning gases.
- A bunsen burner can produce two types of flames depending on the size of the air hole and hence amount of air entering the chimney.

Types of flames.

- Luminous flame.
- Non-luminous flame.

(a). The luminous flame.
- It is a large bright yellow flame produced when the air hole is closed and hence no air enters the chimney.

Characteristics of a luminous flame.
- It is large, quiet and bright yellow.
- Colour is not uniform and it ahs four zones.
- It produces less heat.

- It gives a lot of light to the glow of unburnt hot carbon particles
- It produces soot.

Diagram: the luminous flame.

Parts of a luminous flame.

- **The thin outer zone:**

- Is a fairly visible, narrow zone on the outer surface of the flame.
- At this point methane (lab gas) mixes with sufficient air from the outside and burns completely to carbon (IV) oxide and water.

- **The inner bright yellow zone:**

- It is a large bright yellow zone that lies beneath the thin outer zone.
- Here, air supply is insufficient resulting to **incomplete combustion** of the gas.
- Consequently the gas burns producing **tiny carbon particles** instead of carbon (IV) oxide.
- The white **hot carbon particles glow brightly** and are responsible for the **yellow colour** and the emission of light.
- On cooling the carbon particles form **soot**, which blackens the bottom of the apparatus being heated.

- **The almost colourless inner zone.**

- Is found below the yellow inner zone; and consists mainly of unburnt gases.

- **The blue zone (region)**

- Is found on the outer side of the base of the flame.
- Here, air near the flame rises rapidly due to convection currents and mixes with the burning gas.
- This makes burning more complete than in the two upper parts above it.

Advantages of the luminous flame:

- Can be used for lighting purposes; because it produces more light.

Disadvantages.

- Produces less heat hence inefficient in heating.
- Due to production of soot it blackens apparatus thus preventing better observations of experiments.

(b). The non-luminous flame.

- It is a small blue flame produced when the air hole is completely open and hence a lot of air enters the chimney.

Characteristics of a non-luminous flame.

- It is small, noisy and blue.
- Colour is uniform and it ahs three regions.
- It produces comparatively more (a lot of) heat.
- It does not produce soot, due to complete combustion hence no carbon particles remain.
- It produces less light due to lack of white-hot carbon particles.

Diagram: the luminous flame.

Parts of a non-luminous flame.

- **The outer pale blue region.**

- It is a large light blue zone.
- Here, there is a lot of air coming up the chimney from the air hole and from the outside.
- The air gas mixture thus burns completely to carbon (IV) oxide and water.
- No soot formation because there are no carbon particles.

- **The middle greenish blue region.**

- consists of partially burnt gas-air mixture, due to insufficient air supply.
- However as the mixture rises up the pale blue region, it undergoes complete combustion due to plenty of air (from outside)

- **The inner almost colourless region.**

- Is located at the base of the flame.
- It consists of unburnt gas-air mixture.

Advantages of a non-luminous flame.

- Gives out a lot of heat hence very efficient in hating.
- It does not form soot hence will leave apparatus clean even after experiment (heating).

Disadvantages:
- It uses a lot of laboratory gas in burning.
- cannot be used for lighting purposes since it produces very little light.

Differences between a luminous and a non-luminous flame.

Luminous flam	**Non-luminous flame**
Bright yellow in colour	Blue in colour
Produces a lot of light	Produces a lot of light.
Large and unsteady	Small and steady
Produces soot	Does not produce soot
Has four zones	Has three zones
Burns quietly	Burns noisily
Moderately hot	Very hot

Experiments on Bunsen burner flames.
1. To investigate the heating effects of the luminous and non-luminous flames.
Apparatus:
- Bunsen burners, 250 ml beakers, lighter, stopwatch, tripod stand, wire gauze.

Procedure
- 100 cm^3 of water is put into ach of the two 250 ml beakers.
- One beaker is put over a luminous flame while the other is simultaneously put over a luminous flame
- Time taken for water to boil is noted for each set up.
- The bottom of ach beaker is observed for any changes.

Apparatus

Observations.
- Water heated over the non-luminous flame boiled ion a shorter time than the same amount of water heated over a non-luminous flame.
- The bottom of the beaker heated over the non-luminous flame remained clear but the one heated over the luminous flame was covered with black deposits of soot.

Explanations.
- The non-luminous flame is hotter than the luminous flame; hence boils the water faster
- The hottest part of the luminous flam is the outer blue zone.
- Incomplete combustion in the luminous flame leads to production of carbon particles, which when hot glow yellow and on cooling forms black soot on the beaker;

- Incomplete combustion in a non-luminous flame leads to production of carbon (IV) oxide and steam only, hence no soot formation.

Conclusions.

- The non-luminous flame is hotter than the luminous flame.
- The non-luminous flame is cleaner than the luminous flame.

2. To investigate the hottest part of a non-luminous flame.

Requirements

- Bunsen burner, stiff white paper (cardboard), wooden splint.

Procedure

- A bunsen burner is ignited with the air hole open to get anon-luminous flame.
- A piece of white paper (cardboard) is slipped into the flame in region marked X as shown below.
- The piece of paper is removed quickly before it catches fire.
- A fresh piece of paper is then slipped into region marked Y as shown below; then again quickly removed before it catches fire.
- The experiment for each of the regions marked X and Y is then repeated using wooden splints.
- The splints should be held long enough for some of their parts to get charred

Apparatus

Observations.

- **Using pieces of paper.**

- In region X, the part of the paper that was in contact with the flame was charred uniformly as shown below.
- In region Y, the part of the paper in contact with the flame had a charred ring with an unburnt part in the middle of the ring as shown below

Diagrams

- **Using wooden splints.**

- In region X, the part of the splint in contact with the flame was charred uniformly as shown below.
- In region Y, the part of the splint in contact with the flame had an unburnt part in between two charred regions as shown below.

Diagrams:

Explanations.

- Regions which become charred indicate that they are the hottest part of the flame.
- Region X corresponds to the outermost blue region of a non-luminous flame.
- Region Y is the almost colourless region of the non-luminous flame, which is however surrounded by the middle greenish blue and the outer pale blue zones.
- Thus in region X, the uniform charring of the paper and splint indicate that the outer pale blue zone is the hottest pat of the flame.
- Similarly the charred ring for experiment in region Y show that the parts in contact with the outer pale blue zone gets burnt faster before the parts in contact with the almost colourless or the greenish blue zones.

Conclusions.

- The hottest part of the non-luminous flame is the outermost pale blue zone.
- During heating the object being heated should not be placed nearer the chimney; these parts are less hot.
- For efficient heating the object being heated should be placed at the outermost region of the flame.

3. To show the presence of unburnt gases in a Bunsen burner flame.

Apparatus:

- Bunsen burner, tongs, narrow hard glass tubing.

Procedure

- A bunsen burner is lit and adjusted to get a non-luminous flame.
- A narrow hard glass tubing is held with a pair of tongs and one of its end s is placed in the colourless zone of the flame.
- A match is lit and placed at the free end of the glass tubing.

Apparatus

Observations.

- A flame is obtained at the free end of the glass tubing.

Explanations.

- The tubing trapped unburnt gases at the almost colourless zone of the flame.

- The trapped gases combined with atmospheric air (oxygen) at the other (free) end of the tubing hence the flame.

Conclusions.
- The almost colourless region contains unburnt gases.

4. To show the hottest part of the flame.

Apparatus:
- Bunsen burner, match stick

Procedure
- A matchstick is placed at the top of the bunsen burner chimney using a pin.
- A bunsen burner is lit and adjusted to get a non-luminous flame.
- The match stick is observed fro sometime for any changes.
- If no observable changes are made, the matchstick is then slowly raised towards the blue zone and observed keenly.

Apparatus

Observations.
- The matchstick did not ignite while it was at the bottom of the flame (resting on top of the chimney).
- It ignited as it was being raised towards the outer pale blue zone.

Explanations.
- The bottom of the flame (just on top of chimney) corresponds to the almost colourless zone.
- This zone contains unburnt gases, hence no burning occurs and is thus least hot to cause ignition of the matchstick.
- As the matchstick is raised upwards it moves past the greenish blue zone (where there is partial combustion) then to the outer pale blue zone where there is complete combustion and hence most heat.
- The heat in this region is adequate to cause ignition of the matchstick.

Conclusions.
- The outer pale blue zone is the hottest part of the non-luminous flame, and is thus the correct position to place an object during heating.

Methods of gas collection.

- Various chemical reactions produce gases; some of which are colourless while others are coloured.
- Additionally some gases are poisonous to the human body, while others are major causes of environmental pollution.

Examples:

- **Coloured gases:**

Chlorine (green-yellow); nitrogen (IV) oxide (brown); bromine (red-brown)

- **Colourless gases:**

Oxygen; carbon (II) oxide; carbon (IV) oxide; sulphur (IV) oxide; hydrogen; ammonia etc.

Factors affecting method used in collecting a gas.

- Density
- Solubility in water.
- Colour
- Toxicity

Summary on collection methods.

Method	**Apparatus**	**Characteristic of gas**
Upward delivery - Also called downward displacement of air		- Must be less dense than air. **Examples:** Hydrogen, ammonia gas. **Note:** being lighter the gas is supported by the denser air from below; - When used for colourless gases, it is not possible to know when the gas jar is full;
Downward delivery (upward displacement of air)		- Must be denser than air. **Examples:** carbon (IV) oxide; nitrogen (IV) oxide; chlorine gas; **Note:** - The gas displaces air and settle at the bottom of the collecting vessel - Unless the gas is coloured, it is difficult to know when the container is full
Over water		- insoluble or only slightly soluble in water; - does not react with water **Examples:** carbon (IV) oxide; hydrogen; carbon (II) oxide; **Note:** - with this method it is easy to tell

		when the gas jar or collecting tube is full of gas; - This method cannot be used when the gas is required dry;
Collecting syringe - the gas produced is collected in a syringe;		- Mainly for poisonous gases; since the gases are confined and leakages are limited; **Note:** this method allows collection of small volumes of gases; - It also allows direct measurement of volume of gas produced;

Drying of gases.

- Is the process by which the moisture in a gas being prepared is removed prior to collection.
- This is done by passing the gas through chemicals that absorb moisture.
- Such chemicals are called drying agents.
- The drying agents should not react with the gases being dried.

Examples of drying agents.

- Anhydrous calcium chloride
- Concentrated sulphuric acid.
- Calcium oxide.

Apparatus and drying agents

Collection and drying of some gases

Gas	Collection method	Drying agent
Oxygen	Over water	Concentrated sulphuric (VI) acid; anhydrous calcium chloride.
Hydrogen	Over water, upward delivery	Concentrated sulphuric (VI) acid; anhydrous calcium chloride
Nitrogen	Over water	Concentrated sulphuric (VI) acid; anhydrous calcium chloride
Carbon (IV) oxide	Over water, downward delivery	Concentrated sulphuric (VI) acid; anhydrous calcium chloride
Ammonia	Upward delivery	Calcium oxide

Drugs and drug abuse.

Drug: is a chemical substance that alters the functioning of the body.

Types of drugs

(i). Medicinal drugs (medicines):

- Are drugs mainly used for treatment and prevention of diseases.
- Are also classified into two: over the counter drugs and prescription drugs.

- **Over-the-counter drugs.**

- Are medicinal drugs that can be bought at a pharmacy or retail shop without written instructions from a doctor.

Examples: Mild painkillers like aspirin, panadol, paracetamol, drugs for flu etc.

- **Prescription drugs:**

- Are strong medicines which should only be taken upon a doctors instruction (prescription).
- In this prescription, the doctors give a dosage, which indicates the amount and the rate at which it should be taken.

(ii). Leisure drugs.

- Are drugs that are usually taken for pleasure.
- Are classified into two:

- **Mild drugs:** alcohol, tobacco;
- **Narcotic drugs**: marijuana, cocaine, heroin, mandrax etc.

Drug abuse:

- Is the indiscriminate use of a drug for purposes which it is meant for; or administration of an overdose or underdose of a drug; as well as use of drugs for leisure purposes.

Note:

- The worst form of drug abuse is the taking of drugs for leisure purposes; and the most commonly abused drugs are the leisure drugs.

Effects of commonly abused drugs.

- **Alcohol:**

- Affects the brain and the nervous system
- Damages the liver, and is a common cause of liver cirrhosis.
- Poor health due to loss of appetite.
- Time for working is wasted in drinking and hence less productivity and even lose of jobs; which results to poverty and family disintegrations.

- **Tobacco.**

- Bad breath, discoloured fingers and teeth
- Cause diseases such as bronchitis and tuberculosis.
- damages the lungs and is a common cause of lung cancer due to chemicals found in the cigarettes.
- Smoking during pregnancy is a common cause of miscarriages or still births.
- It is expensive: money used for other better uses is wasted in cigarette smoking.

- **Narcotic drugs:**

- Interferes with the functioning of the brain.
- Results to addiction and drug dependency.
- Some are administered directly into the blood through syringes and hence common routes of transmission of HIV/AIDS.

General effects of drug abuse on the society.

- Drug abuses spend most of their money on drugs and hence neglect their family leading to misery and societal breakdown.
- Drunk drivers cause accidents.
- People who are drunk with a drug are unreasonable and cannot make logical decisions; and hence cannot be productive at that time.
- Drug abuse has resulted into loss of morals leading to higher rates of rapes, violent crimes, murders, prostitution etc.
- Drug abuse has fueled the spread of sexually transmitted diseases and HIV/AIDS.

Note:

Drug addiction:

- Is a situation in which an individual becomes dependent on a particular drug such that he cannot function normally without it; and lack of it result to some discomfort.

UNIT 2: SIMPLE CLASSIFICATION OF SUBSTANCES

Unit Checklist:

1. Elements compounds and mixtures.
2. Mixtures
    - Types of mixtures
    - Separation of mixtures
    - Basic concepts
    - Method of separation of mixtures.
        - Decantation
        - Evaporation
        - Condensation
        - Filtration
        - Crystallization
        - Separating funnel separation
        - Distillation
        - Sublimation
        - Chromatography
        - Solvent extraction
3. Criteria for purity
    - Effects of impurity on melting point
    - Effects of impurity on melting point
4. Nature of matter and kinetic theory of matter.
    - Effects of heat on matter
        - Melting

- Evaporation
- Condensation
- Freezing
- Freezing
- Sublimation

5. Permanent and non-permanent changes
6. Constituents of matter
   - Atoms
   - Elements
   - Molecules
   - Compounds
   - Mixtures.
7. Names and symbols of common elements
8. Simple word equations.

Elements compound and mixtures.

(a). Element:

- Is a pure substance that cannot be split up into simpler substances by chemical means.

Examples: copper, hydrogen, carbon.

(b). Compound:

- A pure substance that consists of two or more elements that are chemically combined.

Examples:

Compound	**Elements in the compound**
Calcium carbonate	Calcium, carbon and oxygen
Sodium chloride	Sodium and chlorine
Ammonium nitrate	Nitrogen, hydrogen, oxygen
Iron (II) sulphate	Iron, sulphur, oxygen

(c). Mixture:

- A substance that consists of two or more elements or compounds that are not chemically combined
- Some mixtures can be naturally occurring while some are artificial.

Examples

Naturally occurring mixtures.

Mixture	Components
Air	Nitrogen, oxygen, carbon (IV) oxide, water vapour, noble gases etc
Sea water	Water and various salts like chlorides of sodium, potassium and magnesium
Crude oil	A mixture of hydrocarbons like methane, petrol, bitumen, etc
Magadi soda	Sodium carbonate, sodium hydrogen carbonate and sodium chloride

Artificial mixtures.

Mixture	Components
Soft drinks	Water, citric acid, sugar, carbon (IV) oxide, stabilizers, sodium benzoate
Black ink	Blue, black, yellow dyes and solvent
Cement	Oxides of aluminium, iron, silicon, calcium and calcium carbonate.

Types of mixtures:

- There are two types of mixtures;

- Homogenous mixtures
- Heterogenous mixtures

(i). Homogenous mixtures.

- Is a mixture with a uniform composition and properties throughout its mass.
- The parts (components) of the mixture are uniformly distributed throughout the mixture

Examples:

Tea with sugar solution.

(ii). Heterogenous mixture:

Is a mixture without uniform composition throughout its mass.

Examples:

- Soil, rocks and sand mixture.

Separating mixtures.

A mixture can be separated into its various components (constituents) by appropriate physical means, depending on type of mixture.

Basic concepts:

Residue: solid that remains on the filter paper during filtration

Filtrate: liquid that passes past the filter paper during filtration

Solute: a solid that dissolves in a particular liquid

Solvent: the liquid in which a solute dissolves.

Saturated solution: a solution in which no more solute can dissolve at a particular temperature

Unsaturated solution: a solution that can take more of the solute (solute) at a particular temperature.

Miscible liquids: liquids that can mix together completely.

Immiscible liquids: liquids that cannot mix together completely.

There are various methods that can be used to separate mixtures.
These include:

- Decantation
- Evaporation
- Condensation
- Filtration
- Crystallization
- Separating funnel separation
- Distillation (simple and fractional)
- Sublimation
- Chromatography
- Solvent extraction

1. Decantation:

- Is a method used to separate insoluble solids from liquids; a heterogenous mixture.

Procedure:

- The solid-liquid mixture is allowed to stand in a container.
- The insoluble solid settles at the bottom and the upper liquid portion poured out with care.

Apparatus.

Examples:

- Separation of sand-water mixture
- Separation of maize flour-water mixture.

Limitations (disadvantages) of decantation.

- It is not efficient as some fine suspended solids may come long with the liquid during pouring.

2. Filtration.

- Is the separation of an insoluble solid from a heterogenous mixture (liquid) using a porous filter that does not allow the solids to pass through.
- Upon filtration the undissolved solid is left on the filter paper and is called the residue.
- The liquid that passes the filter paper is called filtrate.

Examples: separation of sand from water.

(i). Procedure.

- The filter paper is folded into ¼ and opened to from a cone.

Diagram: folding a filter paper.

- It is carefully placed inside a filter funnel.
- The apparatus are then arranged as shown blow.
- The sand-water mixture is then poured into the filter paper in the filter funnel.
- The collecting liquid is directed into a conical flask.

(ii). Apparatus.

Applications of filtration.

- Filtration of domestic water.
- Extraction of medicinal substances from plants.
- Extraction of sugar from sugarcane.
- Operation of a vacuum cleaner.
- Fuel filters in automobile engines.

3. Evaporation.

- Is used to separate a soluble solid from its solution.
- Such solutions are usually homogenous mixtures.
- The solid is called a **solute** while the liquid is called a **solvent**.

Example: separation of salt from salt solution.

(i). Procedure:

- The salt solution is poured in an evaporating dish.
- The set up is then arranged as in the apparatus shown below.
- The solution is boiled under steam or sand bath until all the water in the salt solution evaporates and salt crystals remain in the dish.

(ii). Apparatus.

(iii). Observations and explanations.

- Upon heating the solution, water evaporates because it has lower boiling point than the salt.
- The solution is boiled until salt crystals start appearing on a glass rid dipped into the solution.
- This shall indicate that the solution is saturated.
- The saturated solution is allowed to cool and crystallize.
- The mother liquor (liquid that remains with the crystals) is poured and the salt (solid) dried between absorbent papers.

Note:

A crystal: is a solid that consists of particles arranged in an orderly repetitive manner.

- It is advantageous to boil the solution under a steam or sand bath rather than directly.

Reason:

- The steam or sand bath prevents the mixture from splashing out (spitting) of the evaporating dish.
- It also reduces chances of the evaporating dish cracking.

Applications of evaporation:

- Extraction of soda ash from Lake Magadi.

4. Crystallization and recrystallization.

(a). Crystallization:

- is the process of formation of crystals from a solution.
- It involves evaporation of the solution to form a concentrated solution.

Example: crystallization of potassium nitrate from its solution.

(i). Procedure:

- About 5g of powdered potassium nitrate is added to $10cm^3$ of water in a boiling tube.
- The solution is heated until all the solid dissolves and then allowed to cool and crystallize.

Note:

- More potassium nitrate dissolves in hot water than in cold water.

- The resultant solution is then heated until crystals start appearing; and this can be confirmed by dipping a glass rod into the solution and feeling for crystals.
- This is called a saturated solution i.e. a solution that cannot take in any more of the solute at a given temperature.
- The saturated solution is then allowed to cool and crystallize.

(ii). Observations:
- The resultant solid particles have definite shapes.
- Some are needle-like while others are flat and sharp-edged.
- These are the potassium nitrate crystals.

(b). Recrystallization:
- Is used in obtaining pure crystals from a soluble solid containing impurities.
- Involves filtration and evaporation.

Examples:
Obtaining pure copper (II) sulphate crystals from impure copper (II) sulphate.
Purification of rock salt.

Note: The process can be enhanced by suspending a small piece of pure crystal into the saturated solution.

Diagram: recrystallization of copper (II) sulphate.

Applications of crystallization.
- Separation of Trona from sodium chloride in Lake Magadi.

5. Distillation.
- Is the vapourisation of a liquid from a mixture and then condensing the vapour.
- Is used in the purification of liquids and separation of liquids from a mixture.
- It utilizes the differences in boiling points of the components of the mixture.
- Are of two types:

- Simple distillation
- Fractional distillation.

(i). Simple distillation.
- Is mainly used for purification of liquids containing dissolved substances.
- It is also useful in separating two miscible liquids with widely differing boiling points

Note:
Miscible liquids: Liquids that mix to from a uniform a uniform homogenous solution
- The liquid with the lower boiling point usually distills over first, and is collected.

Example: To obtain pure water from sea water.

(i). Procedure:

- Salty sea water is poured into a distillation flask.
- A few pieces of pumice or porcelain is added to the solution.

Reason:

- To increase the surface area fro condensation and evaporation.

- The solution is heated until it starts boiling, then the burner removed so that the liquid boils gently.
- The boiling goes on until the liquid (distillate) starts collecting in the beaker.

(ii). Apparatus.

(iii). Observations and explanations:

- The water boils and the resultant steam is passed through the **Liebig condenser**.
- As the vapour passes through the condenser, it is cooled by circulating cold water through the jacket of the condenser.
- The cold water enters through the lower bottom and leaves through the top upper part.

Reason:

- To provide more time for the cold water to condense the vapours.

- The distillate is collected in the beaker while the residue remains in the distillation flask.

Applications of simple distillation.

- Manufacture of wins and spirits.
- Desalinization of sea water to obtain fresh water.

(ii). Fractional distillation:

- Is a method used fro the separation of miscible liquids with very close boiling points.

Examples:

Ethanol and water.

- It is a modification of simple distillation in which the fractionating column is inserted on top of the distillation flask.
- All the components must be volatile at different extents in order for separation to be possible.

The fractionating column.

- Is usually an elongated (glass) tube, packed with pieces of glass beads or pieces of broken glass.

Role of glass beads.
- To increase the surface area for vapourisation of the various components of the mixture and allow the separation of the vapours to occur.
- Thus the more the glass beads in the fractionating column, the higher the efficiency of separation.

Note:
- The efficiency of the fractional distillation so s to get more pure components can also be done by:

1. Increasing the length of the fractionating column (making it longer)
2. Making the fractionating column narrower (decreasing the diameter)

Volatile liquids:
- Are liquids with the ability to change into vapour.
- More volatile liquids vapourize and condense faster than the less volatile liquids.

Note:
- During fractional distillation, the components of a mixture are collected at intervals, one at a time with the most volatile (lowest boiling point) coming out first.
- Each component collected in the receiver is called a fraction.

Example: separation of ethanol and water.
(i). Apparatus.

(ii). Procedure:
- Water-ethanol mixture is poured into a round-bottomed flask.
- The apparatus is then connected and set up as shown below.

Note:
- The thermometer bulb must be at the vapour outlet to the condenser.

Reason:
- For accurate determination of the vapourisation temperature for each fraction.

- The mixture is then strongly heated until the first fraction comes out of the distillation flask into the conical flask.
- Collection of the fractions should be done in a conical flask other than in a beaker.

Reason:

- To reduce the rates of evaporation of the fractions, especially the highly volatile ones (in this case ethanol)

- For this particular separation the first temperatures recorded by the thermometer should not exceed 80^{o}C; to ensure that the first fraction is only ethanol.

(iii). Discussion.

- Ethanol boils at 78^{o}C and water boils at 100^{o}C.
- When the mixture is heated, ethanol and water evaporate and pass through the fractionating column which is filled with glass beads to offer a large surface area
- The large surface area encourages evaporation of ethanol and condensation of water vapour.
- Water can be seen dropping back into the distillation flask.
-Ethanol vapour passes through the condenser and warm liquid ethanol is collected in the conical flask.

Note:

- The first portion is almost pure ethanol (about 97%) and burns quietly with a blue flame.
- It also has the characteristic smell of alcohols.

Industrial applications of fractional distillation.

- Separation of air into various components in BOC gases Kenya limited.
- Separation of crude oil into paraffin, petrol, kerosene diesel and other components in the Kenya oil refinery.
- Distillation of ethanol from molasses at Muhoroni Agro-chemicals company.

6. Sublimation.

- Is the process by which a solid changes directly to gaseous state upon heating.
- It is used to separate a mixture in which one of the components sublimes on heating.

Note:

Solid $\rightleftharpoons$ Gas

Solids that sublime have very weak forces of attraction between the atoms and hence are easily broken on slight heating.

Examples of solids that sublime on heating.

- Iodine; sublimes to from a purple vapour.
- Ammonium chloride; sublimes to from dense white fumes;
- Solid carbon (IV) oxide (dry ice);
- Anhydrous iron (III) chloride; sublimes to give red brown fumes.

Examples: separation of iodine from sodium chloride.

(i). apparatus:

(ii). Procedure:
- The iodine-common salt mixture is poured into a beaker and placed in a tripod stand.
- A watch glass full of cold water is placed on the beaker.
- The beaker is heated gently until some dense purple fumes are observed.

(iii). Observations:
- A purple vapour appears in the beaker.
- A dark-grey shiny solid collects on the bottom of the watch-glass.
- White solid remains in the beaker.

(iv). Explanations:
- Upon heating the mixture iodine sublimes and condenses on the cold watch glass to form a sublimate of pure iodine.

Note:
- Solid carbon (IV) oxide (dry ice) is used a s refrigerant by ice cream and soft drink vendors.

Reason:
- It sublimes on heating; as it sublimes it takes latent heat from ice cream (soft drinks) thus leaving it cold.
- It is also advantageous as it does not turn into liquid, which could be cumbersome to carry and would mess up the ice cream.

7. Chromatography.
- Is the separation of coloured substances using an eluting solvent.
- It is also used to identify the components of a coloured substance.
- It involves the use of a moving liquid (**eluting solvent**) on a material that absorbs the solvent.
- It involves **two** major processes:

- **Solubility:**

  The tendency of a substance to dissolve in a solvent.

- **Adsorption:**

  The tendency of a substance to stick on an adsorbent material.

Examples:

1. Separation of components of black ink.

(i). Procedure:
- A filter paper is placed on the rim of an evaporating dish or a small beaker.
- A drop of the black ink is placed at the centre of the filter paper; allowed to spread out and dry.
- A drop of water (ethanol) is then added to the ink and allowed to spread.
- After complete spread of the drop, a second drop is added.
- Water drops are added continuously until the disc of coloured substances almost reaches the edge.

(ii). Observations:

Note: The dry filter paper showing the separated components of a mixture is called a **chromatogram.**

(iii). Explanations:

- Water is the eluting solvent since ink is soluble in it.
- The various dyes in the black ink move at different distances from the black spot hence the bands.

Reasons:

- The dyes have different **solubilities** in the solvent; the more soluble the dye, the further the distance it travels on the absorbent paper
- They have different rates of **adsorption** i.e. the tendency of the dyes to stick on the absorbent material; dyes with low rates of absorption travel far from the original spot.

Note:

Solvent front.

- Is the furthest distance reached by the eluting solvent on the filter paper.

Baseline:

- The point at which the dye to be separated is placed; i.e. it is the starting point of separation.

2. To verify contents of red, black and blue inks.

(i). apparatus:

(ii). Observations:

(iii). Explanations:

- The mixtures A to D have various components with varying solubilities in the solvent (ethanol)
- Mixture D is the most pure because it has only one spot.
- Mixture C is the least pure (most impure), as it has the highest number of spots indicating it is composed of so many dyes (four)
- Mixture C has the most soluble dye; its last component is the one nearest to the solvent front.
- Mixtures with similar dyes in their composition have spots at same levels; in this case A, B and D.

Applications of chromatography.

- Purification of natural products such as hormones, vitamins and natural pigments.
- Detection of food poisons e.g. in canned foods and soft drinks.

8. Solvent extraction.

- Is the extraction of a solute from its original solvent by using a second solvent in which it has a higher solubility

Example: extraction of oil from nuts.

(i). Apparatus.

(ii). Procedure:

- Some nuts are crushed in a mortar using a pestle; to increase the surface area for solubility.
- A suitable solvent such as hexane or propanone (acetone) is added.
- The nuts are further crushed in the solvent.
- The resultant solution is decanted in an evaporating dish, and left in the sun to evaporate.
- The liquid remaining in the evaporating dish is smeared onto a clean filter paper.

(iv). Observations;

- A permanent translucent mark appears on the filter paper.

(v). Explanations:

- The nuts are crushed when in contact with the solvent to bring more of the oil in the nuts closer to the solvent.
- Upon evaporation oil is left behind because it has a higher boiling point than the solvent.
- A permanent translucent mark verifies the presence of oils.

Applications of solvent extraction.

- Used by dry-cleaners to remove dirt (grease) and stains from dry-clean-only clothes such as sweaters, suits, dresses etc.

9. Separating funnel separation.

- Is used fro separating a mixture with two or more immiscible liquids.
- Such liquids do not mix but instead form layers based on their densities.
- The heaviest liquid layer is found at the bottom of the separating funnel; while the lightest liquid is found at the top of the separating funnel.
- The liquids are drained one after the other by opening and closing the tap of the separating funnel.

Example: Separation of oil from water
Apparatus:
Beakers, separating funnel, paraffin oil, distilled water, rubber stopper.

Apparatus.

Procedure.

- The tap of the separating funnel is closed.
- Equal volumes of water and paraffin are put in a separating funnel until it is half full.
- The mouth of the funnel is closed with a stopper and the mixture shaken.
- The mixture is allowed to stand until two distinct layers are formed.
- The stopper is removed and the tap opened to allow the bottom layer to drain into the beaker.
- The tap is closed after most of the bottom layer has drained off.
- The beaker is removed and the rest of the bottom layer is drained into a separate container and discarded; to ensure that no part of the top layer (paraffin) gets into the beaker containing the bottom (water) layer.
- The other (top) layer is then drained into another beaker.

Observations:

- After the mixture has settled oil and paraffin separate into two layers.
- The first beaker contains only water; while the second beaker contains only paraffin.

Conclusion.

- Paraffin and water are immiscible.
- The top layer contains water which is denser while the top layer contains oil (paraffin) which is lighter.

Practical application:

- Extraction of useful substances from complex mixtures.

Example: Separation of iodine from sodium chloride.

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10. Use of magnets.

- Is used to separate solid mixtures, one of which is magnetic (usually iron).
- The iron is picked with a magnet leaving the other components of the mixture behind.

Examples:

- Separation of powdered iron from iron powder-sulphur mixture.

Practical applications:

-In the extraction (mining of iron); where magnetic iron ore is separated from other materials in the crushed ore.

- Separation of scrap iron from non-magnetic materials like glass and plastics in recycling plants.

Criteria for purity.

- **Pure substance:**

- Is a substance that contains only one type of compound or element.

Determination of purity

(a). Solids:

- Purity of solids is determined by measuring the melting point.
- Pure solids melt sharply over a narrow temperature range.

Examples:

- Naphthalene melts at $80^{o}C$ $81^{o}C$.
- Water melts at $0^{o}C$.

Effects of impurity on melting point.

- Impurities lower the melting point of a substance making it melt over a wide range of temperatures i.e. the melting point is not sharp.

Applications:

1. Extraction of metals.

- Impurities are added to purified metal ores to lower their melting points, hence save energy and extraction costs.

2. To prevent knocking of engines due to freezing of water in car radiators during cold seasons, impurities like ethylene glycol or salt (NaCl) is added.

3. Defrosting of frozen roads and sidewalks in temperate countries, by sprinkling a salt such as sodium chloride.

(b). Liquids:

- Purity of a liquid is determined by measuring its boiling point.
- A pure liquid has a sharp boiling point.

Examples:

Pure water boils at $100^{o}C$ at 1 atmospheric pressure.

Pure ethanol boils at $78^{o}C$.

Effect of impurity on boiling point.

- Impurities raise the boiling point of a liquid.

Example:

- Sea water boils at a higher temperature than pure water due to the presence of dissolved salts.

States of matter.

- **Matter:**

- Matter is anything that occupies space and has mass.
- It is composed of either pure substances or a mixture of substances.

States of matter:

Matter exists in three states:

- Solid
- Liquid
- Gas.

Kinetic theory of matter:

- States that matter is made up of small particles which are in continuous random motion.
- The continuous random motion of particles in matter is called Brownian motion.
- The rate of movement of particles in matter depends on the state

(a). Solid state:

- Have closely packed particles held by strong forces of attraction.
- Particles do not move from one point to another but vibrate about a fixed position.
- Upon heating, they start to vibrate vigorously.

Reason:

- Due to increase in their kinetic energy.

(b). Liquid state:

- Particles are further apart from one another than those in the solid state
- Forces of attraction between the particles are weaker than those in the solid.
- The particles remain close, but are free to move from one position to another.
- On heating they gain energy and move rapidly.

(c). Gaseous state:

- The particles are far apart and free to move randomly in all directions.
- Consequently they lack definite shape and volume, but occupy the whole space within a container.
- Forces of attraction between the particles are very weak.

The effect of heat on substances.

- A substance can change from one state to another upon heating or cooling.
- These are physical changes and can be reversed.
- There are **five** processes involved in changes of state:

- Melting
- Evaporation
- Condensation
- Freezing
- Sublimation

The processes involved in change of states of matter.

(i). Melting:
- Is a change of state from solid to liquid.

- **Melting point:**

- Is the constant temperature at which the melting takes place.

Examples:
- Ice melts at 0^{o}C.
- Sodium chloride melts at 800^{o}C.

- During melting the energy supplied to the particles is used to weaken the forces of attraction so that particles can move about.

(ii). Vapourisation (evaporation)
- Is change of state from liquid to gas.

- **Boiling point:**

- Is the constant temperature at which a liquid changes from solid to a gas.

- During boiling, the energy supplied is used to break the forces of attraction in the liquid thus moving the particles far away from each other.

Examples:
Water boils at 100^{o}C, while ethanol boils at 78^{o}C at one atmospheric pressure.
Reason:
The forces of attraction between the water particles are stronger than those of ethanol.

(c). Condensation:
- Is the change of state from a gas to a liquid.
- Is a change due to decrease in temperature.
- When the temperature of gas is decreased, the particles lose kinetic energy to the surroundings to move slowly.
- The attractive forces become stronger, and the sample changes to a liquid.

Note:
- The temperature at which condensation occurs is the same as the boiling point.

(d). Freezing:
- Is the change from a liquid to a solid.
- It is also due to decrease in temperature.
- When a liquid is cooled, the particles lose energy and move very slowly.
- They attract one another strongly, and ultimately remain in fixed positions.

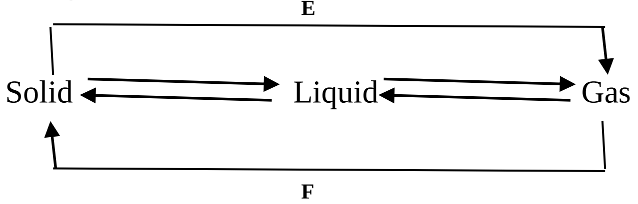
Note:
- The freezing point is the same as the melting point.

(e). Sublimation.
- Is the process whereby a solid does not melt when heated, but changes directly to the gaseous state.

Example:
- Iodine solid changes to purple vapour when heated to 70^{o}C.
- Dry ice (solid CO_2), used to cool ice cream, evaporates without leaving a liquid.

Note:
- The reverse of sublimation, whereby a gas changes directly to solid is called deposition.

Summary on changes of state.



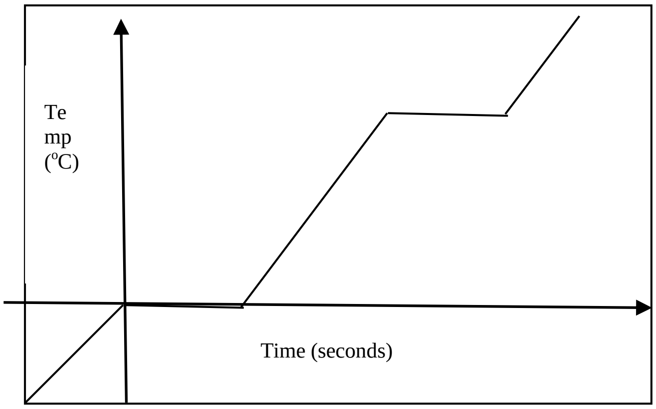
Key: A: melting; **B**: vapourisation; **C**: Freezing; **D**: Condensation; **E**: Sublimation; **F**: Deposition;

Experiment: Investigating changes in temperature when ice is heated.
(i). Procedure:
- A 250ml beaker is half-filled with dry ice, and the initial temperature recorded.
- The ice is heated, while stirring with a thermometer and the temperature recorded every 30 seconds.
- Heating and recording is done until the resultant water starts to boil.
- A graph of temperature against time is plotted.

(ii). Results:

Temperatures (^{o}C)	-10	0	30	60	90	120	150	180	210
Time (seconds)									

(iii). Graph:
Effect of heat on pure ice



(iv). Explanations:

- **Point A-B:**

- As the ice is heated the temperature rises steadily from -10^{o}C to 0^{o}C.
Reason:
- The heat supplied increases the kinetic energy of the ice (solid water) molecules; collisions between them hence increased temperature.

- **Point B-C:**

-The temperature of the ice remains constant even as heat is applied.
Reason:

- Heat supplied is used to break the forces of attraction between the water molecules in ice.
- This is the **melting point** hence at B-C the ice melts.

- **Points C-D:**

- At C, all the ice has already melted (turned to water).
- Between C and D, the temperature of the water increases as heating continues.

Reason:

- The heat supplied increases the kinetic energy of the water molecules; their rate of collision increases hence increased temperatures.

- **Points D-E:**

- Temperature of the water remains constant even as heat is being supplied.
- Point D-E is the **boiling point** i.e. 100°C.

Reason:

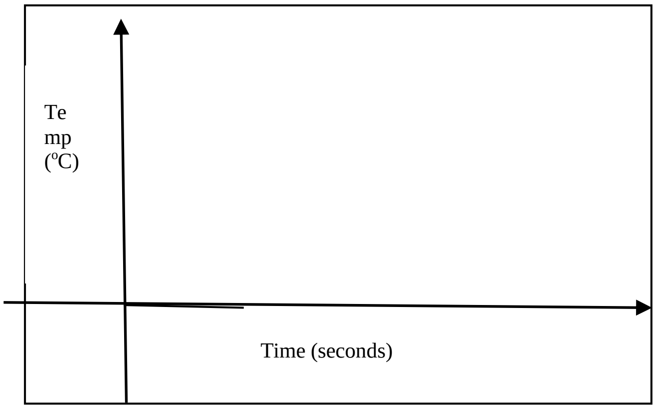
- The heat supplied is used to break the forces of attraction between water molecules in the liquid.

- **Points E-F:**

- At point E, all the liquid water has turned into vapour.
- Thus between E and F, the temperature of the vapour rises as heat is applied.

Note: Heating curve for an impure solid.

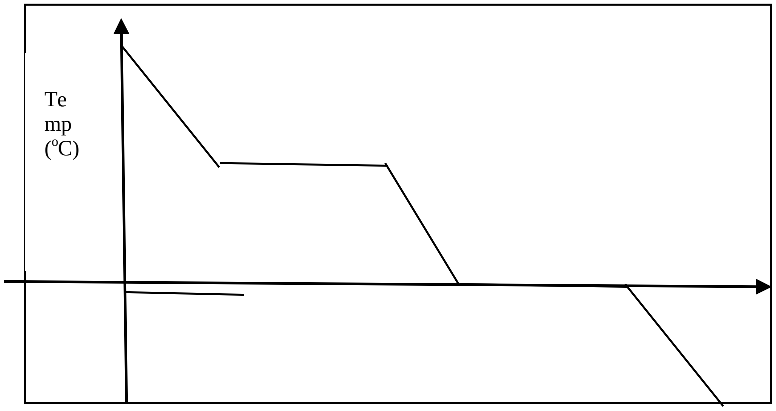
Effect of heat on impure ice



The cooling curve.

- Is a curve that shows how the temperature of a substance changes with time as it is cooled from a gas into a solid.
- It is the opposite of a heating curve.

Example: the cooling curve of water



Time (seconds)

Explanations:

- **Points A-B:**

- Gaseous state; temperature is declining.
- Heat loss results into decrease in kinetic energy of the gaseous molecules.

- **Points B-C:**

- This is the **condensation point**;
- The water vapour condenses to the liquid state; bonds are formed as the hest is lost.

- **Points C-D;**

- Temperature of the **liquid water** is declining;
- Heat loss results into decrease in kinetic energy of the liquid molecules.

- **Points D-E;**

- This is the **freezing point**;
- The temperature remains constant as the heat is being lost;
- The heat being lost results into bond formation; as the liquid forms a solid.

- **Points E-F;**

- The water is now in **solid state**.
- The temperature of the solid declines as heat is being lost.

Permanent and temporary changes.

- Heat causes matter to change.
- Changes due to heat can either be permanent or non-permanent (temporary).

(a). Permanent changes.

- Are also called chemical changes.
- Involves substances that are relatively less stable to heat.

Characteristics of permanent (chemical) changes.

(i). New substances are formed.
(ii). Involves considerable heat changes; energy is either given out or absorbed.
(iii). The mass of the substance changes.

Types of permanent (chemical) changes.

- Chemical changes are of two main types:

- Reversible permanent changes.
- Irreversible (non-reversible) permanent changes.

(i). Reversible permanent changes.

- Are chemical changes in which the final new products can recombine to form the original substance, under certain conditions.

Generally:

$$\underbrace{A + B}_{\textit{Reactants}} \rightleftharpoons \underbrace{C + D}_{\textit{Products}}$$

Examples:

1. Effect of heat on hydrated blue copper (II) sulphate.

(i). Apparatus:

- Test tubes, Bunsen burner, test tube holder
- Hydrated copper (II) sulphate

(ii). Apparatus:

(ii). Procedure:

Dry crystals of hydrated blue copper (II) sulphate are put in a clean dry test tube.

- The apparatus are arranged as above.
- The copper (II) sulphate is heated until no further change.
- The delivery tube is removed from the collected liquid while heating continues.

Reason:

- To avoid sucking back of the condensing liquid which would otherwise rehydrate the anhydrous copper (II) sulphate

- The test tube is allowed to cool and the remaining solid is divided into two portions.
- To one portion of the powder, add distilled water, while to the other potion add the condensed liquid.

(iii). Observations:

- A white solid/ powder remains in the test tube after heating.
- A colourless liquid condenses in the test tube dipped into the ice cold water.
- The colourless liquid turns the white solid into blue.

(iv). Explanations:

- Hydrated blue copper (II) sulphate have water of crystallization, giving it the characteristic blue colour.
- During heating, the heat energy supplied is used to drive out the water molecules (particles) out of the crystals;
- Without water, the copper (II) sulphate turns white and thus called anhydrous copper (II) sulphate;
- The water driven out of the crystals condenses in the test tube immersed in the ice cold water.

Equation

Hydrated copper (II) sulphate $\xrightleftharpoons[\text{Cool}]{\text{Heat}}$ anhydrous copper (II) sulphate + Water

Blue ***White***

Chemically:

$CuSO_4.5H_2O_{(s)} \underset{\text{Cool}}{\overset{\text{Heat}}{\rightleftharpoons}} CuSO_{4(s)} + 5H_2O_{(g)}$;

Confirmatory test for water.

- It turns white anhydrous copper (II) sulphate to blue;
- It changes blue cobalt chloride paper pink;

Confirmatory test that the water is pure:

- It boils at 100^oC;
- It melts at 0^oC;
- It has a density of $1g\ cm^{-3}$;
- It has a refractive index of 1.33;

Note:

- The delivery tube is removed from the collecting liquid while heating is continued; to ensure that no water condenses back into the copper (II) sulphate as this would cause **rehydration;**

- Addition of water to the anhydrous copper (III) sulphate changes its colour from white to blue;

Conclusion:

The effect of heat on copper (II) sulphate is a reversible chemical change;

2. Effect of heat on ammonium chloride.

- When ammonium chloride is heated, it produces ammonia gas and hydrogen chloride gas.
- These are seen as dense white fumes.
- Reversible, when ammonia and hydrogen chloride are gases are reacted or allowed to cool, they produce ammonium chloride;

In summary:

Ammonium chloride solid $\underset{\text{Cool}}{\overset{\text{Heat}}{\rightleftharpoons}}$ ammonium chloride + hydrogen chloride;

White ***Dense white fumes***

Chemically:

$NH_4Cl_{(s)} \underset{\text{Cool}}{\overset{\text{Heat}}{\rightleftharpoons}} NH_{3(g)} + HCl_{(g)}$;

Ammonium chloride ***Ammonia gas*** ***hydrogen chloride gas***

3. Decomposition of calcium carbonate.

In summary:

Calcium carbonate solid $\underset{\text{Cool}}{\overset{\text{Heat}}{\rightleftharpoons}}$ Calcium oxide + Carbon (IV) oxide;

(ii). Irreversible chemical change.

- Are chemical changes in which the resultant products cannot recombine to form the original substance (reactants);
- Majority of the chemical changes are irreversible;

Generally:

$A + B \longrightarrow C + D$

Reactants ***Products***

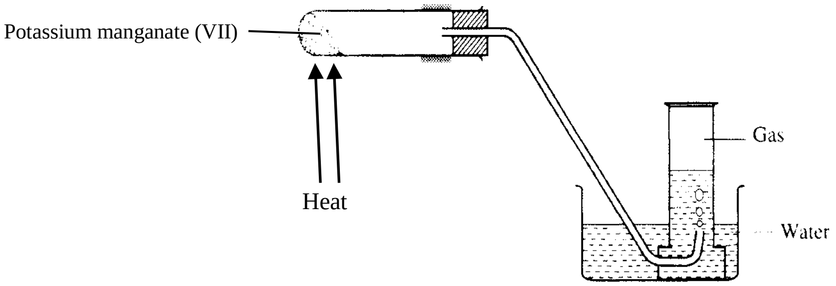
Examples:

1. Action of heat on potassium manganate (VII)

(i). Apparatus and chemicals.

- Bunsen burner, test tube, trough, wooden splint;
- Potassium managnate (VII);

(ii). Apparatus set up.



(iii). Procedure:

- 2 end-fulls of a spatula of potassium manganate (VII) are put in a hard glass test tube;
- The set up is assembled as shown above;
- The solid potassium manganate (VII) is heated, and the resultant gas collected over water;
- The resultant gas(es) is tested with a glowing splint;

(iv). Observations:

- The purple solid turns black;
- A colourless gas collects over water;
- The colourless relights a glowing splint;

(v). Explanations:

- Potassium manganate (VII), a purple solid was decomposed (splint up) on heating to yield (give potassium manganate (III) and oxygen.
- The potassium manganate (III) is the black residue;
- The colourless gas is oxygen; and relighting a glowing splint is the confirmatory test;

In summary:

Potassium manganate (VII) → Potassium manganate (III) + oxygen

Purple solid ***Black solid*** ***Colourless gas***

Note: It is not possible for oxygen and potassium manganate (III) to recombine back to potassium manganate (VII); hence the change is **irreversible;**

2. Thermal decomposition of copper (II) nitrate.

- The blue solid decomposes to form a **black solid**; copper (II) oxide, **red-brown fumes** of nitrogen (IV) oxide and a **colourless gas**, oxygen;

In summary:

Copper (II) nitrate → Copper (II) oxide + Nitrogen (IV) oxide + oxygen

Blue solid ***Black solid*** ***Brown (red-brown) fumes*** ***Colourless gas***

Note: Further examples of chemical changes

- The burning of any substance (except platinum);
- The rusting of iron;
- Addition of water to calcium oxide;
- Explosion of natural gas or hydrogen with air;
- Reacting of sodium in water;

Note: Exothermic and endothermic reactions:

(i). Exothermic reactions.

- Are reactions in which heat is released // given out to the surrounding;
- Usually the final temperature of the reaction vessel // mixture (e.g. beaker is higher than initially;

Examples:

- Freezing;
- Condensation;
- Deposition // sublimation of fumes to solid;

(ii). Endothermic reactions.

- Are chemical reactions in which heat is absorbed from the surrounding;
- The final temperature of the reaction vessel or reaction mixture is usually lower than the initial i.e. they are accompanied by a drop in temperature;

Examples.

- Melting;
- Vapourization;
- Sublimation (of solid to gas)

Summary on chemical changes.

Reaction	Appearance of substance	Changes during reaction	New substance(s)	Type of change
Heating hydrated copper (II) sulphate	Blue	Blue crystals turn into a white powder; colourless liquid condenses on cooling;	Anhydrous copper (II) sulphate and water	Chemical
Heating potassium manganate (VII)	Shiny purple crystals	The purple solid turns black; evolution of a colourless gas;	Potassium manganate (III) and oxygen;	Chemical;
Heating ammonium	White solid //	Dense white fumed that cools to a	Ammonia gas and	Chemical;

chloride	powder;	white solid;	hydrogen chloride gas;	
Heating lead (II) nitrate;	White solid	The white solid turns into a red solid during heating which on cooling turns yellow; - Decrepitating sound; - Brown fumes; - colourless gas;	- Lead (II) oxide; nitrogen (IV) oxide and oxygen gas;	Chemical;
Heating lead (II) nitrate;	White solid	The white solid turns into a yellow solid during heating which on cooling turns white; - Decrepitating sound; - Brown fumes; colourless gas;	- Zinc (II) oxide; nitrogen (IV) oxide and oxygen gas;	Chemical;
Heating copper turnings	Brown turnings;	- Brown turnings // solid turn black;	- Copper (II) oxide;	Chemical;
Rusting of iron;	Grey solid	- Grey solid turns into a red brown solid;	- Hydrated iron (III) oxide;	Chemical;
Heating Copper (II) nitrate;	Blue solid	The blue solid turns into a black solid; - Brown fumes; - colourless gas;	- Copper (II) oxide; nitrogen (IV) oxide and oxygen gas;	Chemical;
Heating copper (II) carbonate	Green solid	The green solid turns into a black solid; - colourless gas;	- copper (II) oxide and carbon (IV) oxide;	Chemical;

(b). Temporary (non-permanent) changes.
- Are also called physical changes;
- They are changes that involve substances that are more stable to heat;
- On heating they do not decompose hence no new substances are formed;

Characteristics of permanent changes.
1. All are reversible upon changes in temperature;
2. No new substance is formed (instead there are only changes of state);
3. The mass of the substances do not change;

Examples:

Solid	**Original appearance**	**Observations during heating then cooling**
1. Candle wax	White sticky solid;	- The solid melts into a colourless viscous liquid; and on cooling solidifies to the original solid wax again;
2. Iodine solid.	Shiny dark-grey crystals;	- The solid turns directly to purple vapour (sublimation); - On cooling the purple iodine vapour (gas) changes directly to solid iodine (deposition); ie. Iodine solid ═ iodine vapour; ***Dark grey*** ***Purple***
3. Zinc oxide	White solid;	- The white solid turns yellow on heating and upon cooling changes back to the original white colour; ie. Zinc oxide ═ Zinc oxide; ***White (cold)*** ***Yellow (hot)***

4. Ice	White	- The solid water melts into liquid and on further heating the liquid vapourizes and turns into gas; - On cooling the gas condenses to liquid which then freezes back into solid; i.e Ice ═ Water ═ Gas
5. Platinum wire;		- A white glow of the metal is seen on heating, but on cooling the metal changes back to its original grey colour;
6. Lead (II) oxide;	Yellow	- The yellow solid turns red on heating and upon cooling changes back to the original yellow colour; ie. Lead (II) oxide ═ Lead (II) oxide; ***White (cold)*** ***Yellow (hot)***

Differences between Physical and Chemical changes.

Physical change	**Chemical change**
1. Produces no new kind of substance;	- Always produces a new kind of substance;
2. UIs usually (generally) irreversible;	Are generally irreversible; with only few exceptions (i.e. most are irreversible);
3. The mass of the substance does not change;	- The mass of the substance changes;
4. No energy is given out or absorbed i.e. are not accompanied by great heat changes;	- Energy is usually given out or absorbed i.e are usually accompanied by great heat changes;

Constituents of matter

- A detailed examination of matter reveals that it is built of very tiny units called toms;
- Presently about 115 atoms have been identified;
- The arrangement and number of atoms in a substance will result into other much larger constituents of matter;
- These are:

  - Elements;
  - Molecules;
  - Compounds;
  - Mixtures;

1. The atom;

- Is the smallest particle of matter that can take part in a chemical reaction;
- It is the smallest particle into which an element can be divided without losing the properties of the element;
- Atoms of various elements all differ from one another;

Examples:

- Copper is made up of many copper atoms;
- Sodium element is made up of many sodium atoms;

2. Elements.

- An element is a substance that cannot be split into anything simple by any known chemical means;
- An element consists of a single type of atom;
- There are about 155 known elements, 90 of which occur naturally.

- Elements are classified into two main groups;

- **Metals:**

- All are solids at room temperature (except mercury); and are good conductors of electricity;

- **Non-metals;**

- Exists as solids and gases;
- All are poor electric conductors except graphite;

Examples of elements.

(i). Metals:

- Sodium, magnesium, potassium, aluminium, lead, iron, zinc, silver, gold, tin, platinum, uranium, calcium, manganese etc.

(ii). Non-elements.

- Carbon, nitrogen, sulphur, oxygen, chlorine, fluorine, argon, neon, bromine, iodine, silicon, boron, xenon, krypton.

3. Molecule.

- Is the smallest particle of a substance that can exist independently;
- It is made when 2 or more atoms (similar or dissimilar) are chemically combined together;
- However atoms of noble/ inert gases exist as single atoms;

Note:

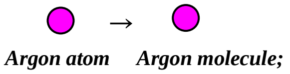
- Depending on number of atoms molecules can be categorized into:

(i). Monoatomic molecule;

- made up of only one atom;

Examples:

- Argon;



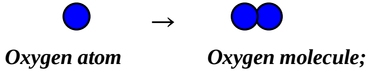
- Neon;
- Helium;

(ii). Diatomic molecules.

- Made up of 2 similar atoms; chemically combines;

Examples

- Oxygen gas;



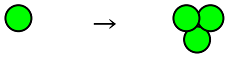
- Nitrogen gas;
- Hydrogen gas;
- Chlorine gas;

(ii). Triatomic molecules.

- Made up of 3 similar atoms; chemically combines;

Examples

- Ozone molecule;



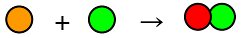
Oxygen atom ***Ozone molecule;***

Note:

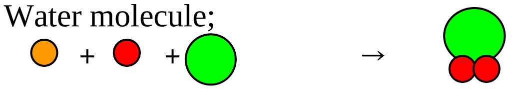
- Other molecules are also made from atoms of different elements chemically combined together;

Examples:

(i). Hydrogen chloride;



(ii). Water molecule;



4. Compounds.

- A compound is a pure substance consisting of two or more elements that are chemically combined.
- Compounds usually have different properties from those of its constituent elements;
- Properties of a compound are uniform throughout any given sample and from one sample to another;

Examples of compounds and their constituent elements.

Compound	**Constituent elements**
Colourless water liquid;	Oxygen and hydrogen;
Green copper (II) carbonate;	Copper, carbon and oxygen;
White sodium nitrate;	Sodium, nitrogen and oxygen;
Black copper (II) oxide;	Copper and oxygen;
Blue copper (II) nitrate;	Copper, nitrogen and oxygen;
Blue copper (II) sulphate;	Copper, sulphur, oxygen and hydrogen;
Ammonium chloride;	Nitrogen, hydrogen and chlorine;

Note:

- Carbonates are derivatives of (derived from or made of) carbon and oxygen;
- Nitrates are derivatives of nitrogen and oxygen;
- Sulphates are derivatives of sulphur and oxygen;
- Hydrogen carbonates are derivatives of hydrogen, carbon and oxygen;

5. Mixtures

- A substance that consists of two or more elements or compounds that are not chemically combined.

Characteristics of mixtures.

- It properties are the average of the properties of its elements;
- Its components can be separated by physical means e.g. filtration, magnetism, distillation etc.
- Its components are not necessarily if fixed positions;
- Are formed by physical means; i.e. there is usually no heat change during its formation.

Examples of mixtures:

1. Air:

- A mixture of oxygen, nitrogen, carbon (IV) oxide, water vapour, and noble gases.

2. Sugar solution.

- A mixture of sugar and water.

3. Sea water.

- Water, dissolved salts;

Experiment: To distinguish between an element and a compound.

(i). Apparatus.

- Watch glass, test tube, wooden splint, magnet, iron fillings, sulphur powder, dilute hydrochloric acid.

(ii). Procedure.

- Approximately 7g of iron fillings and 4g of sulphur are mixed in a test tube and the mixture strongly heated;

Observations:

- A red glow starts and spreads throughout the mixture forming a black solid.
- The black solid is iron (II) sulphide.
- The two products // substances in steps 1 and 2 are subjected to the following tests;

Test // Analysis	Observations	
	Iron-sulphur mixture.	**Iron (II) sulphide**
1. Colour: The colour of the substance is noted;	- The resultant substance is yellow-grey due to the yellow sulphur and the grey iron powder;	- The yellow-grey mixture changes to a black solid; iron (II) sulphide on heating;
2. Separation: A magnet is passed over the substances separately; - Alternatively, water was added to each substance;	- Before heating the iron could be separated from sulphur by use of a magnet or sedimentation; Note: These are physical methods;	- Magnetism and sedimentation have no effect on iron (II) sulphide;
3. Reaction with dilute hydrochloric acid: To each of the substances, a few drops of hydrochloric acid is added;	- Iron reacted with dilute hydrochloric acid to form a colourless gas that burns with a pop sound. - This is hydrogen gas; - Sulphur is not affected;	- Iron (II) sulphide reacted with hydrochloric acid to produce a colourless gas with a characteristic pungent (rotten egg) smell; - The gas is hydrogen sulphide;
4. Heat change.	No heat was produced or applied in mixing iron and sulphur;	- After heating the mixture, the formation of the new substance, iron (II) sulphide produced enough heat hence the bright red glow;

Explanations:

- These four experiments summarize the four main differences between compounds and mixtures.
- From the results, iron and sulphur powder is a mixture; while iron (II) sulphide is a compound.

Differences between a mixture and a compound.

Compounds	**Mixtures**
1. Components are in fixed positions;	- Components are in any positions;
2. Components can only be separated by chemical means;	- Components ca be separated by physical means;

which require large amounts of energy;	
3. The properties are different from those of the constituent substances;	- The properties are the average of the properties of the constituent elements;
4. Are formed by chemical means // methods; i.e. a new substance is formed and there is evolution of heat;	- Are formed by physical mans; no new substance is formed and there is no // negligible heat change;
5. Formation involves heat changes; either liberation or absorption;	- No heat change in the formation of a mixture;

Names and symbols of common elements.
Chemical symbols.
- Are chemical short hands, written to refer to elements.
- They are usually based on the letters of the element;
- Chemical symbols consist of one or two letters which are usually derivatives of the Latin or English name of the element;

Rules in writing chemical symbols.
1. The first letter must always be a capital letter;
2. The letters should not be joined with each other, like in handwriting; they must be printed.
3. If present, the second letter of a symbol must be a small letter;

Note: The abbreviations of the chemical symbols are mainly derivatives of English, Latin or German names.
Examples:
Copper is ymbolised as Cu; derived from Cuprum which is ltin;

- **Iron;** → Fe (Ferrum –Latin)
- **Potassium;** → K (Kalium –Latin)
- **Sodium;** → Na (Natrium –Latin)
- **Lead;** → Pb (Plumbum –Latin);

- The symbol of each element represents one atom of that element.
Example:
- Ag represents one atom of silver;
- 2Ag represents 2 atoms of silver;

Some common elements and their symbols.

Element	**Latin // Greek // German name**	**Symbol**
Carbon	-	C
Fluorine	-	F
Hydrogen	-	H
Iodine	-	I
Nitrogen	-	N
Oxygen	-	O
Phosphorus	-	P
Sulphur	-	S
Aluminium	-	Al
Argon	-	Ar
Barium	-	Ba
Calcium	-	Ca
Chlorine	-	Cl
Helium	-	He

Magnesium	-	Mg
Neon	-	Ne
Silicon	-	Si
Zinc	-	Zn
Copper	Cuprum	Cu
Iron	Ferrum	Fe
Lead	Plumbum	Pb
Mercury	Hydagyrum	Hg
Potassium	Kalium	K
Silver	Argentum	Ag
Sodium	Natrium	Na
Gold	Aurum	Au
Tin	-	Sn
Manganese	-	Mn

Importance of chemical names and symbols over common names.

1. Chemical names and symbols indicate the names in the compounds.

Examples:

- Compounds whose names end in -ide; contains only two elements;
    - Iron (II) sulphide: iron and sulphur;
    - Magnesium nitride: magnesium and nitrogen;
- Compounds whose names end in ate contain three elements and oe of them is oxygen;

2. Chemical names are universally known and accepted; hence provide mean of easy communication among chemists all over the world;

Simple word equations.

Equations:

- Is a linear summary of a chemical reaction, showing the reactants and products.

Examples:

Copper (II) oxide + hydrogen → copper + water;

Explanations:

- Substances on the left hand side are called **reactants;**
- Substances on the right hand side are called **products**;
- The addition sign (+) on the left hand side means "**reacts with**";
- The arrow (→) means to form;
- The addition sign (+) on the right hand side (products side) means "**and**".

Conclusion:

Copper (II) oxide reacts with hydrogen to form copper and water;

Note:

- Some chemical reactions are reversible and hence have two opposite arrows (   ) between reactants and products
- The arrows (   ) in chemistry means a reversible chemical reaction;

Further examples:

1. Copper (II) carbonate → Copper (II) oxide + carbon (IV) oxide;
2. Magnesium + oxygen → magnesium oxide;

UNIT 3: ACIDS BASES AND INDICATORS.

Unit checklist

1. Acids:
    - Meaning of acids;
    - Organic acids
    - Mineral acids;
2. Bases.
    - Meaning;
3. Indicators.
    - Meaning;
    - Preparation of acid-base indicators;
    - Commercial acid-base indicators;
    - Colour of indicators in acids and bases;
    - Classifications of substances as acids or bases using acid base indicators;
    - The universal indicator;
    - The pH scale;
    - pH values of various solutions in universal indiactor;
4. Properties of acids.
    - Physical properties;
        - Taste;
        - Effect on litmus papers;
        - Electrical conductivity;
    - Chemical properties.
        - Reaction with alkalis and bases;
        - Reaction with metals;
        - Reaction with carbonates and hydrogen carbonates;
5. Properties of bases
    - Physical properties;
        - Taste;
        - Texture;
        - Effect on litmus papers;
        - Electrical conductivity;
    - Chemical properties.
        - Reaction with acids;
        - Precipitation of some hydroxides;

    - Effect of heat

6. Uses of some acids and bases.

1. Acids:

- Are substances that dissolve in water to release hydrogen ions.
- Acids can either be organic acids or mineral acids;

(i). Organic acids:

- Are acids found in plants and animals;

Examples:

- lactic acid in sour milk;
- Citric acid in citrus fruits like oranges;
- Ethanoic acid in vinegar;
- Tartaric acid in baking powder;
- Methanoic acid in bee and ant stings;
- Tannic acid in tea;

(ii). Mineral acids.

- Are acids made from minerals containing elements such as sulphur, chlorine, nitrogen etc.
- Are formed from reactions of chemicals;
- Main examples include:
  - **Sulphuric (VI) acid (H_2SO_4)**; contains hydrogen, sulphur and oxygen;
  - **Hydrochloric acid**; contains hydrogen and chlorine;
  - **Nitric (V) acid (HNO_3)**; contains nitrogen, oxygen and hydrogen;

Note:

- Mineral acids are more powerful than organic acids; because they yield // release more hydrogen ions in water
- They are thus more corrosive.

2. Bases.

- Are substances that dissolve in water to yield // release hydroxyl ions;
- Just like acids they are bitter to taste;

Examples:

- Sodium hydroxide;
- Ammonium hydroxide;
- Calcium hydroxide;

Note:

- Some bases insoluble in water while some are soluble in water;
- Soluble bases are called alkalis;

3. Indicators.

- Are substances which give definite colours in acidic or basic solutions;
- Are substances which can be used to determine whether a substance is an acid or a base;
- Consequently they are called acid-base indicators;
- The determination is based on colour changes, where each indicator have particular colourations in acids and bases.
- Indicators can be commercially or locally prepared in the laboratory;

Indicators:

Experiment: preparation of simple acid-base indicators from flower extracts.

(i). Apparatus and chemicals.

- Test tubes;
- Pestle and mortar;
- Flower petals;
- Ethanol // propanone;
- Water;
- Various test solutions: sulphuric (VI) acid, hydrochloric acid, Ethanoic acid, sodium hydroxide, magadi soda, ammonia solution.

(ii). Procedure:

- Flowers from selected plants are collected and assembled e.g. bougainvillea, hibiscus etc;
- They are crushed in a mortar using a pestle and some ethanol added with continued crushing;
- The resultant liquid is decanted into a small beaker; and its colour recorded.
- Using a dropper, two to three drops of the resultant indicator are added to the test solutions.

(iii). Observations:

(a). Colour of extract in acids and bases

Plant extract	**Colour in hydrochloric acid**	**Colour in dilute sodium hydroxide**
1		
2		

(b). Result with various test solutions:

Test substance	**Colour (change)**	**Type of substance (acid/base)**
Lemon juice Wood ash Ammonia Sour milk Vinegar Nitric (V) acid Toothpaste Lime water Baking powder; Sugar Potassium hydroxide		

Note:

- Plant extracts acid-base indicators are not normally preferred in Chemistry experiments.

Reason:

- They dont give consistent (reproducible results because they are impure.
- Commercial indicators give more distinctive and reproducible results.

Commercial indicators.

- Are commercially prepared indicators which are sold in already purified forms.

Advantages of commercial indicators.

- They are relatively pure hence give consistent and reproducible results;
- They are readily available and easy to store in a Chemistry laboratory;

Main examples:

- Phenolpthalein;
- Methyl orange;
- Bromothymol blue;
- Litmus paper;

Note:

- Litmus is a blue vegetable compound which is extracted from plants called lichens;
- Litmus paper is an adsorbent paper which has been dipped in litmus indicator solution then dried;

Colours of various commercial indicators in acids and bases.

Indicator	**Colour in.**		
	Neutral	**Base**	**Acid**
1. Litmus;	Purple	Blue;	Red;
2. Phenolphthalein;	Colourless;	Pink;	Colourless;
3. Methyl orange;	Orange	Yellow	Pink;
4. Bromothymol blue	Blue	Blue	Yellow;

Classification of various substances as acids or bases using indicators.

Substance	**Colour in**				**Classification**
	Litmus	Phenolphthalein	Methyl orange	Bromothymol blue	
Hydrochloric acid					
Sodium hydroxide					
Omo (detergent)					
Soda					
Actal tablets					
Lemon juice					
Sour milk					
Bleach (jik)					
Fresh milk					
Wood ash					

The universal indicator.

- Is a full range indicator which gives range of colours depending on the strength of the acid or alkali.
- It is prepared by suitable mixing certain indicators;
- It gives a range of colour depending on the strength of acids and bases;
- Each universal indicator is supplied with a chart, to facilitate this classification.

The pH scale.

- Is a scale of numbers which shows the strength of acids or bases.
- It refers to the potential (power) of hydrogen;
- It ranges from 0 14;
- To determine the strength of an acid or base, the colour it gives in universal indicators solution is compared to the shades on the pH chart of the indicator;

Diagram: The pH scale.

Note:

- The strongest acid has a pH of 1;
- The strongest alkali has a pH of 14;
- Neutral substances have a pH of 7;
- Any pH less than 7 is acidic solution; while any pH above 7 is for a alkaline / basic solution;

Colour and pH of various solutions in universal indicator;

Substance	Colour	pH on chart	Classification
Hydrochloric acid			
Sodium hydroxide			
Omo (detergent)			
Soda			
Actal tablets			
Lemon juice			
Sour milk			
Bleach (jik)			
Fresh milk			
Wood ash			

Properties of acids.

(a). Physical properties.

1. They have a sour taste.

Examples:

- The sour taste of citric fruits is due to the citric acid in them.
- The sour taste in sour milk is due to lactic acid;

2. They turn blue litmus to red;

- Red litmus will remain red in acidic solution; blue litmus will turn red;

3. Electrical conductivity;

- Acids conduct electric current when dissolved in water;
- This is because they dissolve in water to release hydrogen ions; which are the ones that conduct electric current;

4. Thy destroy clothing when strong; i.e. they at away clothing material leaving holes in it;

5. Strong acids are corrosive; hence able to burn plant and animal tissues;

Chemical properties.

1. Reaction with alkalis ad bases.

- Acids react with alkalis to form salt and water only;
- These types of reactions are called ***neutralization reactions***;
- The hydrogen ions of the acid react with the hydroxyl ion of the alkali to form water;
- The name of the salt is usually derived from the acid;

Examples:

Acid	**Derivative salt**
Sulphuric (VI) acid	Sulphates;
Hydrochloric acid	Chlorides
Nitric (V) acid	Nitrates;
Phosphoric acid	Phosphates;

Summary:

Acid + base (alkali) → salt + water; (a neutralization reaction);

Examples:

1. Sodium hydroxide + Hydrochloric acid → Sodium chloride + water;
2. Calcium oxide + Sulphuric (VI) acid → calcium sulphate + water;

2. Reaction with metals;

- Acids react with some metals to produce hydrogen;

Examples: Reaction with dilute hydrochloric acid and zinc metal;

Procedure:

- 2 cm3 of hydrochloric acid is put in a test tube;
- A spatula end-full of zinc powder is added.
- A burning splint is lowered in the test tube.

Observations.

- Effervescence of a colourless gas;
- The colourless gas burns with a pop sound;

Explanations.

- Zinc metal displaces the hydrogen ions in the acid which form the hydrogen gas;
- When a glowing splint is introduced into the hydrogen gas; it burns with a pop sound;
- This is the chemical test to confirm that a gas is hydrogen;

Conclusion;

- The gas produced is hydrogen gas;
- Thus, acids react with some metals to produce hydrogen gas, and a salt;

General equation:

Metal + Dilute acid → salt + Hydrogen gas;

Reaction equation:

Zinc + Hydrochloric acid → Zinc chloride + Hydrogen chloride;

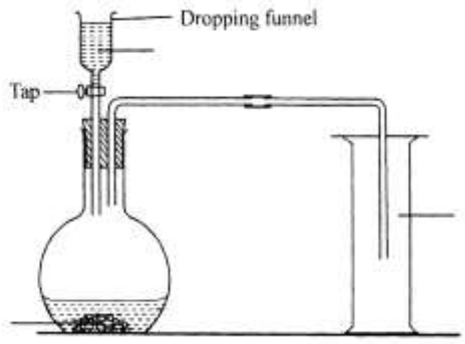
Further examples:
i. Magnesium + Dilute sulphuric (VI) acid → magnesium sulphate + hydrogen gas;
ii. Magnesium + Dilute Hydrochloric acid → magnesium chloride + Hydrogen gas;

3. Reaction with carbonates and hydrogen carbonates.
- Metal carbonates and hydrogen carbonates react with acids to form carbon (IV) oxide, water and a salt;

General equation:
Metal carbonate + Dilute acid → A salt + water + carbon (IV) oxide;
Metal hydrogen carbonate + Dilute acid → A salt + water + carbon (IV) oxide;

Examples: Reaction of sodium carbonate with dilute hydrochloric acid.
Procedure:
- About 2 cm^3 of dilute hydrochloric acid is put in a test tube;
- A spatula end-full of sodium carbonate powder is then added;
- A burning splint is carefully lowered into the test tube.

Apparatus.



Observations.
- An effervescence occurs (bubbles); and a colourless gas is produced;
- The colourless gas does not relight a glowing splint; showing that it is carbon (IV) oxide;

Properties of bases.
Note:
- Bases are substances that release hydroxyl ions when added to water;
- Soluble bases are called alkalis;

Examples:
Sodium hydroxide + Water → Sodium ions + hydroxyl ions;

(a). Physical properties of bases
1. They are bitter to taste;
2. They are slippery or soapy to feel;
3. They turn litmus blue;

4. They conduct electricity / electric current. This is because when they are added to water they release hydroxyl ions which are the ones that conduct electricity;

(b). Chemical properties.
1. Reaction with acids.
- They react with acids to form a salt and water as the only products;
- This is a neutralization reaction; and is used to cure indigestion;

Example:
- Actal tablets contain a base that neutralizes the stomach acid.

2. They precipitate some metal hydroxides.
- Addition of some alkalis to salt solutions results in formation of solids;
- Most of these are normally hydroxides;
- A solid that is formed when two solutions are mixed is called a precipitate;

Example:
- Copper (II) sulphate + Sodium hydroxide → Copper (II) hydroxide + Sodium sulphate solution;
Blue solid;

3. Effects of heat.
- Most metal hydroxides are decomposed by heat to form their oxides and water;

General equation:
Metal hydroxide $\xrightarrow{\text{Heat}}$ Metal oxide + Water;

Example:
Zinc hydroxide → Zinc oxide + water;

Applications of acids and bases.
Application of acids
1. Manufacture of aerated drinks;
2. Cleaning metal surfaces to remove oxide layer;
3. Sulphuric (VI) acid is used in car batteries to store and produce electricity;
4. Treatment of some insect bites.

Examples:
- Wasp and bee stings can be treated by applying vinegar (Ethanoic acid) or lemon juice;
- These acidic substances neutralize the alkaline insect stings;

Uses of bases.
1. Manufacture of anti-acid tablets to neutralize acid indigestions e.g. actal;
2. Calcium oxide ad calcium chloride are used to dry gases in the laboratory;

UNIT 4: AIR AND COMBUSTION.

Checklist.

1. Components
2. Determination of percentage of the active part of air.
   - Burning candle.
   - Heating copper turnings;
   - Heating magnesium turnings.
   - Smouldering of white phosphorus.
   - Rusting of iron;
3. Determination of presence of water and carbon (IV) oxide in water.
4. Fractional distillation of liquid air.
5. Rusting
   - Meaning and formula;
   - Conditions necessary for rusting;
   - Prevention of rusting;
6. Oxygen
   - Laboratory preparation of oxygen gas;
   - Chemical test for oxygen gas;
   - Alternative methods of oxygen preparation;
     - Addition of water to solid sodium peroxide;
     - Heating potassium manganate (VII) solid.
   - Use of oxygen
7. Burning substances in air;
   - Changes in mass;
   - Burning substances in oxygen;
     - Metals;
     - Non-metals;
   - Reactivity series;
8. Competition for oxygen among metals;
9. Applications for the competition for oxygen;
10. Atmospheric pollution.

Introduction:

- Air is a gaseous mixture constituted of several gases, water vapour and pollutants.

Combustion:

- Is the burning of substances, usually in presence of air // oxygen;
- During combustion only the oxygen component of air is used; .e the active part of air.

Percentage composition of air.

Component	**Percentage volume.**
Nitrogen	78%
Oxygen	21%
Carbon (IV) oxide	0.03%
Noble gases (argon)	About 1%
Water vapour	Variable
Smoke/dust particles	Variable;
Others	Trace

Note:

- From the noble gases argon is the most abundant, constituting about 0.93% of the entire 1%

Oxygen and combustion.

- When substances burn in air they consume oxygen.
- Thus the process of combustion utilizes mainly oxygen;
- The reactions in combustion are normally exothermic (give out heat) and often involve flame.

Note:

Combustion in which a flame is used is called burning;

- In combustion if all the oxygen in a given volume of air is used, the final volume of air reduces by about 21.0%;
- Since oxygen is the only constituent of air participating in combustion its termed the active part of air.

Experiments: Determination of the active part of air.

1. Burning candle in air.

Apparatus and requirements.

- Candle;
- Cork / evaporating dish;
- Sodium hydroxide solution;

Procedure:

- A candle about 3cm long is put on a wide cork/ evaporating dish;
- It is then floated in a dilute solution of sodium hydroxide solution just above the beehive shelf;
- It is carefully covered with a dry 100cm^3 measuring cylinder, during which the level of solution in the cylinder is noted and marked;
- The measuring cylinder is removed and the candle lit;
- The lighting candle is then covered with a measuring cylinder;
- The experiment is allowed to proceed until the candle goes off;

Observations:

- The candle went off after sometime;
- The sodium hydroxide level inside the gas jar rises;
- The sodium hydroxide level in the trough goes down;

Diagrams:

Explanations:

- The candle wax is made up of hydrogen and carbon, hence called a hydrocarbon;
- During burning it melts in air consuming oxygen and producing carbon (IV) oxide and water vapour;
- The water vapour condenses giving a negligible volume of water;
- The resultant carbon (IV) oxide is absorbed by the sodium hydroxide;
- Absorption of carbon (IV) oxide in the gas jar creates a partial vacuum within it;
- The sodium hydroxide in the trough rises to fill the resultant space; and hence a drop in the sodium hydroxide level in the tough;

Conclusion.

- Oxygen is the active part of air that is utilized during burning;
- Air is basically made up of 2 parts; an active part that supports burning and an inactive part that does not support burning;

2. Quantitative determination of percentage of oxygen in air.

Apparatus and chemicals.

- Tough;
- Beehive;
- Candle and gas jar;
- A 30ml ruler;
- Sodium hydroxide solution;

Procedure:

- The entire apparatus is arranged as shown below;
- An empty gas jar is inverted over the candle before lighting it;
- The initial height A, is measured and recorded;
- The gas jar is then removed; the candle lit and covered with the gas jar again;
- The set up is allowed to run till the candle extinguishes (goes off); and the final height (B) of the air column measured.

Diagrams

Calculations:
- Amount of air in the gas jar initially = A cm^3;
- Final amount of air remaining after burning; = Bcm^3;
- Amount of oxygen used; = (A B)

Thus;

$$\text{Percentage of oxygen in air:} = \frac{(A\ \ B) \times 10}{A}$$

$$= C\%$$

Sample data:

Volume of air in the gas jar before burning =

Volume of air in the gas jar after burning =

Volume of air used during burning =

$$\text{Percentage of air (by volume) used up} = \frac{\text{Volume used in burning}}{\text{Original (initial) volume}} \times 100$$

Substituting:

Conclusion:
- When candle burns in air, about 20% of air, which is oxygen used up;

3. Determination of percentage of air used up in rusting.

Apparatus:
- Gas jar;
- Trough;
- Beehive;
- Iron fillings//powder

Procedure:
- The gas jar is divided into five equal portions by marking around it using a waterproof marker;
- The gas jar is wet near the bottom and some iron fillings sprinkled on it;
- Some water is put in a trough and the jar with iron fillings // wool // powder inverted over it;
- The initial colour of iron fillings is noted;
- The set up is left undisturbed for a few days until the water shows no further change in rising;

Diagrams of apparatus set up:

Observations:
- The iron fillings change colour from a grey to form a brown solid;
- Water level in the gas jar rises // increases until the first mark in the gas jar;
- The water level in the trough decreases;

Explanations:
- The gas jar is moistened to make the iron fillings stick onto its surface so that the fillings do not fall in the water when the gas jar is inverted;
- The brown substance formed is called **rust** and its chemical name is **hydrated iron (III) oxide**;
- During rusting, oxygen is utilized, thus creating a partial vacuum in the gas jar;
- This causes the water level in the gas jar to rise up and the water level in the trough to go down;
- The rise in water level is equivalent to about 1/5 of the original air volume, which translates to about 20%;

Conclusion:
- When rusting occurs about 20% of air, which is oxygen, is used up;

4. Determination of percentage of air used up when air is passed over heated copper.
Apparatus and chemicals.
- Two 100cm^3 syringes labeled Y and Y;
- Hard glass test tube;
- Glass wool;
- Bunsen burner;
- Copper turnings;

Procedure:
- A small amount of copper turnings is put in a hard glass tube and glass wool put at both ends of the tube;
- All the air in the syringe Y is removed by pushing the plunger inside, upon which the syringe is tightly fixed at one end of the tube// glass tube;
- The plunger of syringe Z is pulled out to the 100cm3 mark; to fill it with air.

The apparatus is arranged as below.

Diagram:

- The tube containing copper turnings is strongly heated;
- Air is then passed over the hot copper turnings by slowly pushing the plunger Z to and fro for several times.

Reason:
- Ensure complete reaction between the hot copper turnings and oxygen (air);

- When no further air change in volume of air in the syringe occurs, the apparatus is allowed to cool;
- The volume of air left in syringe Z is recorded;

Observations:
- The brown solid (copper) turns into a black solid (copper (II) oxide);
- The plunger of syringe Z moves inwards to approximately 80 cm3 mark;

Explanations:
- The heated copper reacted with oxygen in air to form black copper (II) oxide;
- The percentage of oxygen that was in the air is approximately 20%, causing the plunger to move inwards to the 80 cm^3 mark;

Equation:

Copper + Oxygen → Copper (II) oxide;
(Brown) (Colourless) (Black)

$2Cu_{(s)} + O_{2(g)} \rightarrow 2CuO_{(s)}$
Brown (Colourless) (Black)

Conclusion:
- Burning of copper in air utilizes oxygen and produces black copper (II) oxide.

Sample results and calculations:
- Initial air volume in syringe Z = cm^3;
- Final air volume in syringe Z = cm^3;
- Volume of air used = (100 80) = 20 cm^3;

Percentage of oxygen in air = $\frac{20}{100}$ x 100 = 20%;

4. Determination of percentage of air used up when air is passed over heated Magnesium.
- When the same set up is used to investigate the percentage of air used up in combustion of magnesium the volume of air used up is relatively higher than the 20%.

Reason:
Magnesium produces a lot of heat during combustion and thus reacts with both **oxygen** and **nitrogen** to form two products; **magnesium oxide** and **magnesium nitride** respectively;

Observations:
- Magnesium glows giving a bright blinding flame;
- Formation of a mixture of two white powders.

Equations:

Reaction with oxygen:

Magnesium + Oxygen → Magnesium oxide;
(Grey) (Colourless) (White)

$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$
(Grey) (Colourless) (White)

Reaction with nitrogen:

Magnesium + Nitrogen → Magnesium nitride;
(Grey) (Colourless) (White)

$3Mg_{(s)} + N_{2(g)} \rightarrow Mg_3N_{2(s)}$
(Grey) (Colourless) (White)

Note:

- Sodium metal will also react with both oxygen and nitrogen during combustion; forming sodium oxide and sodium nitride respectively;

5. Smouldering of white phosphorus.

Apparatus and requirements.

- Graduated measuring cylinder;
- Water tough // pneumatic trough;
- Copper wire;
- White phosphorus;

Procedure:

- An empty measuring cylinder is inverted in a water trough and the water level noted;
- A small piece of white phosphorus is attached to the end of a piece of copper wire then put // inserted into the inverted measuring cylinder ensuring it is above the water;
- The set up is left undisturbed for 24 hours;

Precaution:

- Avoid contact with the phosphorus;
- Avoid inhalation of the fumes;

Observations:

- White fumes inside the cylinder at the start of the experiment;
- After 24 hours:
    - water level inside the measuring cylinder rises;
    - Water level in the trough drops;

Explanations:

- Yellow or white phosphorus smoulders in air; due to the fact that phosphorus reacts with oxygen to form phosphorus oxides;
-The phosphorus oxides are the white fumes;
- The phosphorus oxides then dissolves in water; forming acidic solutions of phosphoric acids;
- The water level rises inside the cylinder to occupy the volume of oxygen used up in reaction with phosphorus;

Equations:

Phosphorus + Oxygen → Phosphorus (V) oxide
White // yellow Colourless White fumes

$P_{4(s)} + 5O_{2(g)} \rightarrow 2P_2O_{5(g)}$;
White Colourless White fumes
Yellow

Phosphorus + Oxygen → Phosphorus (III) oxide
White // yellow Colourless White fumes

$P_{4(s)} + 3O_{2(g)} \rightarrow 2P_2O_{3(g)}$;
White Colourless White fumes
Yellow

Conclusion:

- Phosphorus smolders easily in air, reacting with oxygen (active part of air) to form phosphorus (III) or phosphorus (V) oxide;
- For this reason phosphorus is stored under water; to prevent it from reacting with atmospheric oxygen;

Note:

- This reaction can be made much faster by heating the copper wire; which will transmit heat to the piece of phosphorus at the tip, causing rapid burning of phosphorus to give dense white fumes of phosphorus (V) oxide // phosphorus (III) oxides;

Test for evidence of some components of air.

1. Water vapour.

(i). Formation of dew;

(ii). When white anhydrous copper (II) sulphate is left in the open overnight; it forms a blue solid of hydrated copper (II) sulphate;

Reason:

- The white anhydrous copper (II) sulphate absorbs atmospheric water vapour;
- Upon hydration the copper (II) sulphate turns blue;

(iii). Sodium hydroxide pellets form a colourless solution when left in the open air overnight.

Reason:

- The sodium hydroxide pellets absorbs atmospheric water vapour and dissolves in it forming sodium hydroxide solution;

(iv). When air is passed through anhydrous calcium hydroxide solid in a U-tube for sometime; there is formation of a colourless solution in the U-tube.

Apparatus:

Reason:
- The anhydrous calcium chloride absorbs atmospheric water vapour forming a colourless solution of calcium chloride;

Equation:
Calcium chloride + water → calcium chloride solution;
White ***Colourless solution;***

Note:
- Substances that absorb moisture from the air to form a colourless solution are called deliquescent substances.
- Other examples of deliquescent substances include: iron (III) chloride, magnesium chloride and zinc chloride;

2. Carbon (IV) oxide.
(i). Glass stoppers of reagent bottles containing sodium hydroxide solution tend to stick when left on for sometime in the laboratory;
Reason:
- The sodium hydroxide solution at the edges of the stopper is exposed to air; and thus reacts with atmospheric carbon (IV) oxide forming white sodium carbonate solid.

Equation:
Sodium hydroxide + Carbon (IV) oxide → Sodium carbonate + Water;
$2NaOH_{(aq)} + CO_{2(g)} \rightarrow Na_2CO_{3(s)} + H_2O_{(l)}$;

(ii). Bubbling atmospheric air through lime water (calcium hydroxide) to form a white insoluble salt of calcium carbonate.

Apparatus:

Reason:
- Atmospheric carbon (IV) oxide reacts with limewater (calcium hydroxide) to form a white insoluble precipitate of calcium carbonate salt;

Equation:
Calcium hydroxide + Carbon (IV) oxide → Calcium carbonate + Water;

$Ca(OH)_{2(aq)} + CO_{2(g)} \rightarrow CaCO_{3(s)} + H_2O_{(l)}$;

Note:
- When the air is bubbled on even after the formation of the white precipitate; the white precipitate dissolves after sometime to form a colourless solution;
Reason:
- Excess carbon (IV) oxide reacts with the calcium carbonate to form soluble calcium hydrogen carbonate.

Equation:
Calcium carbonate + Water + Carbon (IV) oxide → Calcium hydroxide solution;

$CaCO_{3(s)} + H_2O_{(g)} + CO_{2(g)} \rightarrow Ca(HCO_3)_{2(aq)}$;

Fractional distillation of Liquefied air.
- Air is a mixture of gases;
- It can be separated into its constituents by fractional distillation of liquid air.
- During the process air is passed through a series of steps during which it is purified, some components eliminated then it is compressed into liquid prior to fractional distillation.
- The process can be divided into two main stages;

Purification and liquefaction;
Fractional distillation of air;

(a). Purification and liquefaction.
Step 1: Purification:
- The air is purified by removal of dust particles ;
- This is done through the following ways:

- Passage through filters; during which air is passed through a series of filters; the dust particles remain within the filters while dust free air passes on to the next stage;
- Electrostatic precipitation where air is passed through charged electrodes which trap oppositely charged dust particles;

Step 2: Removal of carbon (IV) oxide.
- The dust-free air is passed through a chamber containing calcium hydroxide solution;
- The sodium hydroxide solution dissolves the carbon (IV) oxide present in the air;
- During the reaction, sodium carbonate and water are formed;
- Over a prolonged time; the sodium carbonate absorbs more (excess) carbon (IV) oxide forming sodium hydrogen carbonate;

Equations:
Sodium hydroxide + carbon (IV) oxide → Sodium carbonate + Water;
$2NaOH_{(aq)} + CO_{2(g)} \rightarrow Na_2CO_{3(aq)} + H_2O_{(l)}$;

In excess;
Sodium carbonate + Water + Carbon (IV) oxide → Sodium hydrogen carbonate;

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

Step 3: Removal of water vapour;
- The dust-free, CO_2 – free air is then cooled to -25°C;
- This process solidifies the water vapour out as ice;
- This cooling process may be done at temperatures a s low as -80; so as to solidify any carbon (IV) oxide (freezing point -78°C) that may have escaped absorption by the sodium hydroxide;
- The removal of water vapour and carbon (IV) oxide are important because it prevents blockage of the pipes in the rest of the system;

Step 4: Liquefaction of air;
- The dry, dust-free and carbon (IV) oxide-free air is compressed to about 100 atmospheres of pressure; causing it to warm;
- The compressed air is cooled by refrigeration;
- The cold compressed air is made to expand rapidly by passage through a nozzle which cools it further;
- The repeated compression, cooling and expansion of air causes it to liquefy at about -200°C:

Note:
At this temperature only neon and helium whose boiling points re lower than -200°C remain in gaseous states;

(b). Fractional distillation;
- The liquid air now consists only of nitrogen, oxygen and noble gases (especially argon);
- The liquid air is fed at the bottom of a fractionating column;
- It is warmed to a temperature of -192°C;
-Nitrogen distils over fast at -196°C because it has a lower boiling point; and is collected at the top of the fractionating column;

Note:
- Any vapours of oxygen and argon which rise together with nitrogen vapour condense in the column and fall back as liquids;
- The nitrogen collected is 99% pure;

The small amounts of impurities include neon and helium;
- The liquids remaining at the bottom of the fractionating column after vaporization of all nitrogen is mainly oxygen and argon; with traces of krypton and xenon;

- The liquid is again warmed further to a temperature of -185°C; causing the vapourization of argon whose boiling point is -186°C;
- This is collected as a gas at the top of the fractionating column;
- The residue liquid is mainly oxygen with minute quantities of krypton and xenon which have even high boiling points;
- The oxygen is drained off and stored as pressurized oxygen in steel cylinders;

Uses of the products:
Oxygen;

- Used in hospitals with patients with breathing difficulties;
- It is used by mountain climbers and deep-sea divers for breathing;
- It is used to burn fuels;
- It is combined with acetylene to form oxy-acetylene flame which is used in welding;
- During steel making oxygen is used to remove carbon impurities;

Nitrogen:

- Manufacture of ammonia;
- Used in light bulbs; because of its inert nature it dies not react with the filament;
- As a refrigerant e.g. storage of semen for artificial insemination;

Rusting.

- Is the corrosion of iron in presence of oxygen and moisture to form brown hydrated iron (III) oxide;
- The chemical name **rust** is therefore **hydrated iron (III) oxide** with the formula $\mathbf{Fe_2O_3.2H_2O}$;
- Rust itself is a brown porous substance;

Disadvantage of rusting:

- It weakens the structure of the metal (iron) and hence eventually destroys them.

Experiment: To show the conditions necessary for rusting.

Experiment	Procedure	Observation	Explanation
1	- Two clean iron nails are put inside the test tube; - 10 cm^3 of tap water are then added; - Examine for two days;	- Iron nails turn brown implying there is rusting;	- There is presence of both oxygen and water;
2	- Two clean iron nails are added to the test tube; - 10 cm^3 of boiled hot water is added followed by about 3 cm^3 of oil; - Examine for two days;	- No rusting occurs;	- There is water but no oxygen so no rusting occur; - Boiling the water removes any dissolved oxygen; - Addition of the oil on top prevents dissolution // entry of any air containing oxygen into the water;
3	- Two clean iron nails are added to the test tube; - Push a piece of cotton wool half way the test tube; - Place some anhydrous calcium chloride on it and cork the tube tightly; - Examine for two days;	- No rusting occurs;	- There is no air // oxygen but no water; - Anhydrous calcium chloride absorbs any moisture form the air in the test tube; - Corking the tube tightly prevents more moisture from the atmosphere from getting into the tube as the calcium chloride may get saturated and allow moisture into the nails;
4	- Two clean iron nails are added to the test tube; - Examine for two days;	- Some little rusting occurs;	- Air contains and oxygen and some moisture that will facilitate rusting;
5	- Two clean iron nails are added to the test tube; - Add salty water; - Examine for two days;	Rusting occurs; and at a faster rate than the rest;	- Rusting occurs due to presence of both water and oxygen in the salty water; - Rusting is faster because the salty water contains ions which gain

			electrons hence facilitate faster oxidation of iron;

Summary diagrams

Further explanations:

- During rusting the first step is the oxidation of iron b7y xygen (in the air) to form anhydrous iron (III) oxide;

Equation:

Iron + Oxygen → Iron (III) oxide;

$4Fe_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_{3(s)}$;

- The anhydrous ion (III) oxide then undergoes hydration with water to form brown hydrated iron (III) oxide;

Equation:

Anhydrous iron (III) oxide + Water → Hydrated iron (III) oxide.

Black ***Brown;***

- Rusting occurs faster in salty conditions;

Reason:

- The initial step is the oxidation of iron, from iron (II) ions (Fe^{2+}) to iron (III) ions (Fe^{3+});
- During oxidation iron (II) ions give out electrons to undergo oxidation and form iron (III) ions;
- Salty water contains several dissolved salts whose ions easily accept electrons from the iron (II) ions and thus accelerating the oxidation of iron and hence rusting;

Prevention of rusting.

Note:

- Rusting destroys materials; equipment and roofs made of iron;
- Rust is porous and thus allows air and water to reach the iron beneath.
- Thus if not removed iron will continue corroding until it is all eaten up.

Methods of preventing rusting.

1. Galvanizing.

- Is the coating of iron with a small layer of zinc;

- Can be done by either dipping the iron object in molten zinc, spraying with a spray of molten zinc, or by electroplating (electrolytic deposition);
- On exposure to air the zinc acquires an inert layer of zinc oxide that is impervious to both air and water;
- The iron beneath is thus prevented from air and water and thus rusting.

Note:
- The iron is protected even if the zinc coating is scratched.

Reason:
- Upon scratching both the iron and zinc get into contact with air and water;
- Since zinc is more reactive than iron, air and water reacts with zinc at the expense (instead) of iron;

2. Electroplating:
- Refers to electrolytic coating of metals (iron) with less reactive (less corrosive metals);
- This is done through the process of electrolysis where floe of electric current causes the less reactive metal to coat the metal being protected from rusting;

Example:
- Most tin cans are in fact made yup of steel coated with a thin layer of tin.
- Other than being non-toxic tin is unreactive and rarely reacts with the contents of the can or air;

Note:
- Unlike in galvanizing, when an electroplated material gets scratched, the metal underneath (iron) rusts, and very fast;

Reason:
- Both iron and the electroplating metal (tin) are exposed to air and water;
- Since iron is more reactive than tin (the less reactive electroplating material) it reacts with oxygen and water in preference to tin;

- This explain why galvanization is more durable than electroplating;
- Other less reactive metals that are used to coat iron objects include chromium, silver and gold;
- Some electroplating metals such as gold and silver also increase the aesthetic value of the electroplated object;

3. Sacrificial protection.
- Blocks of a more reactive metal such as zinc or magnesium are attached to the iron structure;
- The more reactive metal will be corroded in preference to iron;
- To keep the iron structure from rusting, the block of reactive metal has to be replaced regularly;
- This metal is used for the protection of underground water pipes as well as ship hulls;
- The blocks of reactive metal are either attached directly to the iron structure or connected to it by a wire.

Diagrams:

4. Painting.
- The paint coats the metal surface and thus prevents contact with air and water hence no rusting;
- However if the paint is scratched, rusting occurs quickly;
- It is used mainly in ion railings, gates, bridges, roofs, ships ad cars;

5. Alloying.
- Alloys are mixtures of two or more metals;
- Thus to prevent iron from rusting it may be mixed with one o more metals resulting into a substance that does not rust;

Example:
- Stainless steel is an alloy of iron with chromium, nickel and manganese and it resistant to rusting.

7. Oiling and greasing;
- Oil is used in moving engine parts while grease is used I other movable metal joints;
- The oil // grease forms a barrier that prevents water and air from coming into contact with the metal surface and hence preventing rusting;
- Oiling and greasing are unique in the sense that they are the only methods that can be used to prevent rusting in movable car parts;

Oxygen.
- A very important constituent of air;
- Lavoisier (1743 1794), A French Chemist showed that it is the component of air used in respiration and also in burning fuels;
- It is the most abundant of all elements; occurring both freely as well as in combination with other elements;
- Freely it constitutes about 21% by volume of atmospheric air;

Laboratory preparation of oxygen gas.

Apparatus.
- Zinc; round-bottomed // flask bottomed flask; thistle // dropping funnel; rubber stopper, deliver tubes, rubber tubings, beehive shelf, trough, gas jars, wooden splint, hydrogen peroxide (20% by volume), manganese (IV) oxide.

Diagram;

Procedure.
- Some manganese (IV) oxide is placed into a flat-bottomed flask;
- The apparatus is set up as shown in the diagram above;
- Add hydrogen peroxide from a thistle funnel into the flask dropwise;
- The gas is collected as shown;

Observations:
- Bubbles of a colourless gas are released from the flask through the water then into the gas jar;
- The colourless gas collects on top of the water;

Explanations:
- Hydrogen peroxide decomposes slowly to oxygen and water under normal conditions;
- This process is however slow to collect enough volumes of oxygen;
- On addition of manganese (IV) oxide the decomposition is speeded up;
- Thus manganese (IV) oxide speeds up the decomposition of hydrogen peroxide and thus acts a s a **catalyst**;

Equation:

- **Without a catalyst:**

Hydrogen peroxide → Water + oxygen;

$2H_2O_{2(l)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$ *(slow process)*

- **With manganese (IV) oxide catalyst:**

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

$2H_2O_{2(l)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$ *(faster process)*

Note:

The first few bubbles of oxygen gas are not collected.

Reason: The gas is mixed with air which was originally in air and hence impure.

Method of collection;
- Over water collection.

Reason:
- It is insoluble in water and less dense than water

Physical properties of oxygen gas.
- It is colourless;
- It is odourless;
- Has a low boiling point of about -183^{o}C;
- Almost insoluble in water (hence collected over water);

Chemical test for oxygen gas.
- On inserting a glowing splint on a gas jar full of oxygen gas; it relights a glowing splint;

Drying of oxygen gas.

- The resultant oxygen is usually moist due to the fact that it is collected over water;
- If required dry the gas ca be died using either of the two methods:

(i). Using sulphuric (VI) acid.

- Bubbling the gas through a wash bottle containing concentrated sulphuric (VI) acid;
- The concentrated sulphuric (VI) acid absorbs moisture from the gas leaving it dry;
- The dry gas is then draw into collection syringe;

Diagram;

(ii). Using anhydrous calcium chloride.

- From the flask the gas is passed through a U-tube containing anhydrous calcium chloride;
- The anhydrous calcium chloride also absorbs moisture from the gas leaving it dry;
- The dry gas is then drawn into a collection syringe;

Diagram:

Alternative methods of oxygen preparation.

1. Addition of water to sodium peroxide.

Apparatus:

- Sodium peroxide; round-bottomed // flask bottomed flask; thistle // dropping funnel; rubber stopper, deliver tubes, rubber tubings, beehive shelf, trough, gas jars, wooden splint, water;

Diagram of apparatus.

Procedure.

- Some sodium peroxide is placed into a flat-bottomed flask;
- The apparatus is set up as shown in the diagram above;
- Add water from a thistle funnel into the flask dropwise;
- The gas is collected as shown;

Observations:

- Bubbles of a colourless gas are released from the flask through the water then into the gas jar;
- The colourless gas collects on top of the water;

Explanations:

- Sodium peroxide reacts with water to liberate oxygen;
- A solution of sodium hydroxide remains in the flask;
- This solution will turn litmus paper blue showing it is alkaline.

Equation:

Sodium peroxide + water → Sodium hydroxide + oxygen;

$2Na_2O_{2(l)} + 2H_2O_{(l)} \rightarrow 4NaOH_{(aq)} + O_{2(g)}$

Note:

The first few bubbles of oxygen gas are not collected.

Reason: The gas is mixed with air which was originally in air and hence impure.

Method of collection;

- Over water collection;

Reason:

- It is insoluble in water and less dense than water;

Chemical test for oxygen gas.

- On inserting a glowing splint on a gas jar full of oxygen gas; it relights a glowing splint;

2. Heating potassium manganate (VII) solid.

Apparatus:

- Ignition tube // boiling tube; means of heating; solid potassium manganate (VII); rubber stopper, deliver tubes, beehive shelf, trough, gas jars, wooden splint, water;

Diagram of apparatus.

Procedure:
- The apparatus is set up as shown above.
- Some solid potassium manganate (VII) is put in a hard ignition// combustion tube and strongly heated as shown above.
- The resultant gas is collected over water as shown above.

Observations;
- The purple solid forms a black solid (potassium manganate (II) solid);
- Bubbles of a colourless gas are evolved and collect over water;

Explanations:
- Upon heating potassium manganate (VII) decompose to manganese (VI) oxide; potassium

Equation:

Potassium manganate (VII) → Potassium manganate (II) + Oxygen gas

$KMnO_{4(s)} \rightarrow KMnO_{2(s)} + O_{2(g)}$;

Uses of oxygen.

1. Used in hospitals for breathing by patients with breathing difficulties;
2. It is used by mountain climbers and deep sea divers for breathing;
3. It is used to burn fuels e.g. burning fuels for propelling rockets;
4. Used in welding and cutting metals

Examples:
- It combines with hydrogen to form a very hot oxy-hydrogen flame that is used in welding and cutting metals;
- It combines with acetylene to form oxy-acetylene flame which is also used in welding and cutting metals;

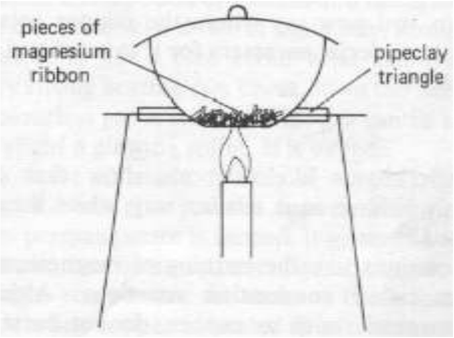
5. During steel making, oxygen is used to remove iron impurities.
- During this process oxygen is blown over hot impure iron.
- The oxygen react with carbon impurities forming carbon (IV) oxide which escapes laving pure iron which is steel due to its higher purity;

Burning substances in air.
- When substances burn in air they mainly react mainly with oxygen (the active part of air);
- Some metals however also react with nitrogen;
- During burning there is usually change in mass;

Experiment: To investigate burning substances in air.

Requirements:
- Metal (magnesium ribbon); crucible; tripod stand; pipe clay triangle; means of heating;

Apparatus.



Procedure:

- About 1g of magnesium is put in the crucible;
- The crucible (with the magnesium is then weighed)
- The apparatus is set up as above;
- The crucible is heated with the lid lifted occasionally; so as to allow in air;
- No content of the crucible is allowed to escape; to ensure all products of the burning are retained;
- After all the magnesium has burned the crucible is allowed to cool;
- The crucible and its contents are weighed again;

Observations

Mass of crucible + magnesium before burning = xg
Mass of crucible + contents after burning = yg
Change in mass = (x y) g;
Mass of product before burning is lower // less than the mass of the product after burning;

Explanations:

- When the magnesium is burned in a closed crucible in a closed container, most of the air is consumed;
- It is therefore necessary to allow in air so that the burning can continue;
- During burning the magnesium combines with air to form a new product;
- Magnesium combines with both oxygen and nitrogen in air to form magnesium nitride and magnesium oxide;

Equations

With oxygen:

Magnesium + oxygen → Magnesium oxide;
$Mg_{(s)} + O_{2(g)} \rightarrow MgO_{(s)}$;

With nitrogen:

Magnesium + Nitrogen →Magnesium nitride;
$3Mg_{(s)} + N_{2(g)} \rightarrow Mg_3N_{2(s)}$;

Conclusion:

- Generally when metals burn in air, there is increase in mass;
- All metals react with oxygen to form metal oxides;
- Only more reactive metals react with nitrogen in air;

Note:

- During burning if the product(s) of the burning is gaseous, then there would be decrease in mass.

Examples:

Phosphorus → Phosphorus (V) oxide;
Lead (II) nitrate → Phosphorus (V) oxide + Nitrogen (IV) oxide + Oxygen gas;
Calcium carbonate → Calcium oxide + carbon (IV) oxide;

Burning metals in air and in oxygen.

Requirements:

- Metals; deflagrating spoon; gas jar; source of heat;

Diagram of apparatus;

Procedure:

- A piece of sodium is warmed on a deflagrating spoon until it begins to burn;
- It is then lowered into a gas jar of air as shown above;
- The flame colour is noted;
- The gas jar is allowed to cool; some water added to the product(s) in the gas jar and shook well;
- Any gases produced are tested by smell and also with litmus papers;
- The experiment is then repeated with pure oxygen;
- The whole procedure is repeated with other metals;

Observations;

- When substances burn in oxygen they form only oxides; as opposed to burning substances in air where some react with both air and nitrogen;
- Different substances produce different flame colours;
- Many metals burn in air and in oxygen at different speeds; with more reactive metals burning more vigorously than the less reactive metals;
- Burning is faster in oxygen than in air;

Reason:

- Oxygen is pure but in air there are other constituents such as nitrogen, carbon (IV) oxide and noble gases which slow down the burning;
- In air products are generally oxides and in some few cases (magnesium and sodium) nitrides as well;

- Metals that tend to be more reactive are the ones that react with both oxygen and nitrogen;
- In oxygen products are strictly oxides;
- Some of then products are soluble in water while others are not.

Sample equations:
Magnesium:
With oxygen:
Magnesium + oxygen → Magnesium oxide;
$Mg_{(s)} + O_{2(g)} \rightarrow MgO_{(s)}$;

With nitrogen:
Magnesium + Nitrogen →Magnesium nitride;
$3Mg_{(s)} + N_{2(g)} \rightarrow Mg_3N_{2(s)}$;
Sodium:
With oxygen:
Sodium + oxygen → Sodium oxide;
$4Na_{(s)} + O_{2(g)} \rightarrow 2Na_2O_{(s)}$;

With nitrogen:
Sodium + Nitrogen →Magnesium nitride;
$6Na_{(s)} + N_{2(g)} \rightarrow 2Na_3N_{(s)}$;

Summary: burning metals in air.

Metal	How it burns	Appearance of product	Name of products	Solubility of product in water	Effect of solution on litmus paper
Magnesium	Burns with a bright white flame;	White powder	Magnesium oxide and magnesium nitride;	Slightly soluble; alkaline gas (ammonia) is produced during the process;	Turns blue;
Copper	Burns with a blue flame; surface turns black;	Black solid;	Copper (II) oxide;	Insoluble;	No effect;
Iron.	Glows to red hot; produces sparks;	Brown black (dark brown) solid;	Iron (II) oxide	Insoluble;	No effect;
Sodium	Buns very vigorously with a golden yellow flame;	White solid;	Sodium oxide and sodium nitride	Soluble; alkaline gas (ammonia) is produced in the process;	Turns litmus blue;
Calcium	Vigorous with a red flame;	White solid;	Calcium oxide and calcium nitride;	Slightly soluble; alkaline gas evolved in the process;	Turns blue;
Zinc		Yellow solid which cools to white;	Zinc oxide	Insoluble;	No effect
Lead		Red solid which cools to yellow;	Lead (II) oxide	Insoluble;	No effect;
Potassium	Very vigorously with a lilac flame;	White solid;	Potassium oxide and potassium nitride;	Soluble; alkaline gas evolved in the process;	Turns blue;

Note:

- When metals combine with oxygen, it forms metal oxides. In these reactions oxygen is added to the metals; hence the reaction is called **oxidation**.
- **Oxidation** refers to the addition of oxygen to a substance;
- The reactivity of various metals with oxygen differs.
- The arrangement of the metals in order of their activity forms the **reactivity series**;
- Metallic oxides generally turn litmus paper blue and are thus said to be **basic oxides**;
- Some metallic oxides however have both acidic and basic properties and are thus termed **amphoteric oxides** e.g. aluminium oxides;

The Reactivity series of metals;

Potassium;	↑	Most reactive;
Sodium;		
Calcium;		
Aluminium;		
Zinc;		*increasing reactivity;*
Iron;		
Lead;		
Copper;		
Mercury;		
Silver;		
Gold;		Least reactive

Burning non-metals in oxygen

Requirements:

- Non-metals; deflagrating spoon; gas jar; source of heat;

Diagram of apparatus;

Procedure:

- A piece of sulphur is heated on a deflagrating spoon until it begins to burn;
- It is then lowered into a gas jar of oxygen as shown above;
- The flame colour is noted;
- The gas jar is allowed to cool; some water added to the product(s) in the gas jar and shook well;

- Any resultant solution is tested with litmus papers;
- Any gases produced are tested by smell and also with litmus papers;
- The experiment is then repeated with pure oxygen;
- The whole procedure is repeated with other non-metals such as carbon and phosphorus;

Explanations:

1. Sulphur.

- Burns in oxygen with a blue flame to give a colourless gas with a choking irritating smell;
- The gas is sulphur (IV) oxide;

Equation:

Sulphur + Oxygen → Sulphur (IV) oxide;

$S + O_{2(g)} \rightarrow SO_{2(g)}$;

- The sulphur (IV) oxide dissolves in water to form sulphurous acid, which turns litmus rd;

$SO_{2(g)} + H_2O_{(l)} \rightarrow H_2SO_{3(l)}$;

2. Carbon

- Glows red to give a colourless gas that forms a white precipitate in lime water;
- The gas is Carbon (IV) oxide;

Equation:

Carbon + Oxygen → Carbon (IV) oxide;

$C + O_{2(g)} \rightarrow CO_{2(g)}$;

- The Carbon (IV) oxide dissolves in water to form weak carbonic acid, which turns litmus rd;

$CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(l)}$;

Note:

- In limited oxygen the carbon undergoes partial oxidation forming carbon (II) oxide;

Equation:

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

3. Phosphorus.

- Burns in oxygen with a white flame to give dense white fumes;
- The white fumes are either phosphorus (V) oxide or phosphorus (III) oxide;
- Both solids // fumes dissolve in water to form phosphoric acid;

Equations:

With limited supply of oxygen:

Phosphorus + Oxygen → Phosphorus (III) oxide;

$P_{4(s)} + 3O_{2(g)} \rightarrow 2P_2O_{3(g)}$;

With excess oxygen.

Phosphorus + Oxygen → Phosphorus (V) oxide;

$P_{4(s)} + 5O_{2(g)} \rightarrow P_2O_{10(g)}$;

Summary: effects of burning non-metals in air.

Non-metal	**How it burns in**	**Name of products**	**Appearance of the**	**Effect of solution on**

	oxygen	**formed**	**product**	**litmus paper**
Sulphur	Burns with a blue flame;	Sulphur (IV) oxide	White fumes;	Turns ed
Carbon	Glows red	Carbon (IV) oxide	Colourless gas;	Turns red;
Phosphorus	Burns with a white flame	Phosphorus (V) oxide and phosphorus (III) oxide	White fumes	Turns red;

Precautions:

The experiment should be done in a fume cupboard.

Reason:

Products of burning sulphur and phosphorus in air are poisonous.

Note:

- Most non-metallic oxides are acidic in nature and therefore turn litmus red and are thus referred to as acidic oxides;
- Some non-metallic oxides form oxides which are neither acidic nor basic and are thus termed neutral oxides; e.g. carbon (II) oxide and water (hydrogen oxide);

Competition for combined oxygen among elements.

Apparatus:

- Metal oxides, source of heat, metals.

Procedure:

- A spatula end full of copper (II) oxide in a bottle top.
- Magnesium powder and mixed well;
- Record the observations;
- The experiment is repeated using other metal oxides with various other metals like zinc, iron etc.

Observations:

Metal Metal oxide	Magnesium	Zinc	Iron	Lead	Copper
Magnesium oxide (white)	No reaction	No reaction	No reaction	No reaction	No reaction
Zinc oxide (white)	White magnesium oxide and grey zinc metal	No reaction	No reaction	No reaction	No reaction
Iron (III) oxide	White magnesium oxide and grey iron metal;	White zinc oxide and iron;	No reaction	No reaction	No reaction
Lead (II) oxide (yellow when	White magnesium oxide and	White zinc oxide and lead;	Iron (III) oxide and lead;	No reaction	No reaction
Copper (II) oxide (Black	White magnesium oxide and brown copper	White zinc oxide and brown copper metal;	Brown iron (III) oxide and brown copper metal	Yellow lead (II) oxide and brown copper metal	No reaction

	metal;				

Explanations:

- A more reactive metal takes away oxygen from a less reactive metal;
- This is because a more reactive metal reacts more readily with a less reactive metal;
- These reactions are called **displacement reactions**;
- Some metals can displace other metals from their oxides upon heating;
- Metals which are higher in the reactivity series can displace metals which are lower in the reactivity series from their oxides;
- From the table none of the metals can displace magnesium from its oxide, while copper can be displaced from its oxides by all the metals.
- Thus from the list **magnesium** is the **most reactive** while **copper** is the **least reactive**.
- Such results of displacement reactions can also be used to develop a reactivity series of the metals (elements) concerned.

Selected equations:

1. Copper (II) oxide + Magnesium → Magnesium oxide + Copper

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

Black Grey White Brown.

2. Zinc (II) oxide + Magnesium → Magnesium oxide + Zinc

White when cold Grey White Grey.
Yellow when hot

$ZnO_{(s)} + Mg_{(s)} \rightarrow MgO_{(s)} + Cu_{(s)}$;

3. Copper (II) oxide + Zinc → Zinc oxide + Copper

Black Grey White when cold Brown.
Yellow when hot

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

4. Lead (II) oxide + Magnesium → Magnesium oxide + Lead

Yellow when cold Grey White Grey.
Redwhen hot

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

5. Copper (II) oxide + Zinc → Magnesium oxide + Copper

$CuO_{(s)} + Zn_{(s)} \rightarrow MgO_{(s)} + Zn_{(s)}$;

Black Grey White Brown.

Typical reactivity series from the results above:

Magnesium ↑ Most reactive
Zinc
Iron
Lead
Copper ↓ Least reactive

Note:

- Removal of oxygen is called **reduction;**
- Addition of oxygen is called **oxidation;**
- A substance that loses oxygen during a reaction is said to be **reduced** while a substance that removes oxygen from another is called **reducing agent**;

- A substance that gains oxygen during a reaction is said to be oxidized while a substance that loses / donates oxygen to another is called an **oxidizing agent**;

Examples:

1. Copper (II) oxide + Magnesium → Magnesium oxide + Copper

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

Black Grey White Brown.

Reducing agent: magnesium
Oxidizing agent: Copper (II) oxide
Oxidized species: Magnesium
Reduced species: Copper

2. Copper (II) oxide + Zinc → Magnesium oxide + Copper

$CuO_{(s)} + Zn_{(s)} \rightarrow MgO_{(s)} + Zn_{(s)}$;

Black Grey White Brown.

Reducing agent: Zinc
Oxidizing agent: Copper (II) oxide
Oxidized species: Zinc
Reduced species: Copper

- In the above reactions both reduction and oxidation take place at the same time;
- A reaction in which both reduction and oxidation occur at the same time is called a **redox reaction**; red from reduction and ox from oxidation;

Application of Redox reactions:

1. Extraction of metals;

- Ores of metals such as zinc, iron lead etc are roasted in air to form corresponding metal oxides;
- The metal oxides are then reduced to corresponding metals using common reducing agents like carbon and carbon (II) oxide.

Examples:

Zinc (II) oxide + Carbon (II) oxide → Zinc + carbon (IV) oxide;

Atmospheric pollution and percentage composition of air.

- Human activities have changes the normal composition of air in some places;
- This has not only altered the percentage composition of the main components but also added other components into the air.

Examples:

- Mining increases the amount of dust particles in the air;
- Geothermal power drilling may result into emission of gases like hydrogen sulphide, sulphur (IV) oxide into the air;
- Industrial processes like manufacture of nitric (V) acid, contact process etc may add gases ilke sulphur (IV) oxide, nitrogen oxides into the air;

These gases and emissions cause atmospheric pollution:
- Gases like sulphur (IV) oxide and nitrogen (IV) oxide dissolve I rain water to form acidic rain that causes corrosion of buildings, iron sheet roofing, bleaching of plants; irritation in bodies and respiratory surfaces in animals etc;
- Dust particles may block stomata in plants; cause smog formation hence reducing visibility (leading to more cases of road accidents).

Uses of oxygen:

1. Used for breathing in hospitals fro patients with breathing difficulties.
2. Used for breathing by mountain climbers and deep sea divers.
3. Used to burn fuels e.g. burning fuels to propel rockets.
4. Manufacture of the oxy-acetylene flame that is used in welding and cutting of metals;
5. Removal of iron impurities during steel making i.e. oxygen is blown through impure iron; the oxygen then reacts with carbon impurities forming carbon (IV) oxide which escapes leaving behind pure iron (steel).

UNIT 5: WATER AND HYDROGEN.

Checklist:

1. Introduction
2. Burning candle wax in air.
3. Reaction of water with metals
    - ✔ Sodium
    - ✔ Potassium
    - ✔ Calcium
4. Reaction of metals with steam.
    - ✔ Calcium
    - ✔ Magnesium
    - ✔ Zinc
    - ✔ Iron
    - ✔ Aluminium
5. Hydrogen
    - ✔ Hydrogen;
    - ✔ Reduction property of hydrogen
    - ✔ Burning hydrogen in air;
6. Uses of hydrogen

Introduction:

- Water is the most abundant substances on earth;
- It covers about 71% of the earths surface;
- Main sources of water include seas, lakes, rivers, oceans.

Burning candle wax in air.

Apparatus:

Procedure.

- The candle is lit under the funnel and the suction pump turned on.
- The set up is left undisturbed for about 15 minutes.

Observations;

- The candle continues to burn.
- Droplets of a colourless liquid in the tube A;
- The colourless liquid turns white anhydrous copper (II) sulphate to blue and blue anhydrous cobalt (II) chloride into pink;
- A white precipitate forms in the calcium hydroxide in tube B;
- Deposits of a black solid on the inner sides of the funnel;

Explanations;

- The suction pump ensures continuous supply of air hence the candle continues to burn;
- Candle wax buns in oxygen to form carbon (IV) oxide and steam;
- The carbon (IV) oxide is sucked out through the apparatus by the suction pump;
- Carbon (IV) oxide forms a white precipitate of calcium carbonate when bubbled through lime water (calcium hydroxide)
- Incomplete combustion of the carbon in the candle wax produces carbon particles which cools and deposits as black solids;

Equations:

As the candle burns:

Carbon + Oxygen → Carbon (IV) oxide;

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

Hydrogen + Oxygen → Carbon (IV) oxide;

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

For the formation of the black deposits (soot)

Carbon + Oxygen → Carbon + Carbon (IV) oxide;

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

In the calcium hydroxide:

Carbon (IV) oxide + calcium hydroxide → Calcium carbonate + Water

Colourless ***Colourless*** ***White precipitate*** ***Colourless***

$CO_{2(g)} + Ca(OH)_{2(aq)} \rightarrow CaCO_{3(s)} + H_2O_{(l)}$;

- The steam condenses into water in the boiling tube;
- Water turns white anhydrous copper (II) sulphate to blue and blue anhydrous cobalt chloride paper into pink;

General equation:

Hydrocarbon + Oxygen → Water + Carbon (IV) oxide;

Conclusion:

- Candle wax is a compound of carbon and hydrogen only; and such compounds are defined as **hydrocarbons**;
- When burned in air (oxygen) hydrocarbons produce carbon (IV) oxide ad steam (water);
- Other examples of hydrocarbons include: petrol; diesel; kerosene etc;

Note: Effects of repeating the same experiment without a suction pump.

Apparatus:

Observations:

- The candle went off;
- Deposition of black solid on the inner sides of the funnel;
- No colourless liquid in tube A;
- No white precipitate in tube B;

Explanations:

- The carbon (IV) oxide and steam produced would accumulate in the filter funnel hence making the flame to go off;
- Incomplete combustion of the candle would produce carbon particles which cool as soot;
- Only negligible amount of water and carbon (IV) oxide would pass through the apparatus;

Reactions of water with metals:

1. Potassium.

Procedure:

- A small piece of potassium metal is cut and dropped into a trough containing water;
- The resultant solution is tested with litmus paper;

Diagram of apparatus:

Observations and explanations:

- The metal floats on the water surface; because it is less dense than water;
- A hissing sound is produced; due to production of hydrogen gas;
- It explosively melts into a silvery ball then disappears because reaction between water and sodium is exothermic (produces heat). The resultant heat melts the potassium due to its low melting point.
- It darts on the surface; due to propulsion by hydrogen;
- The metal bursts into a lilac flame; because hydrogen explodes into a flame which then burns the small quantities potassium vapour produced during the reaction;
- The resultant solution turns blue; because potassium hydroxide solution formed is a strong base;

(b). Reaction equations.

Equation I

$2K_{(s)} + 2H_2O_{(l)} \rightarrow 2KOH_{(aq)} + H_{2(g)}$;

Equation II

$4K_{(s)} + O_{2\,(g)} \rightarrow 2K_2O_{(s)}$;

Equation III:

$K_2O_{(s)} + H_2O_{(l)} \rightarrow 2KOH_{(aq)}$

Effect of resultant solution on litmus paper;

- Litmus paper turns blue; sodium hydroxide formed is highly soluble in water; releasing a large number of hydroxyl ions which result into alkaline conditions // high pH;

2. Sodium.

Procedure:

- A small piece of sodium metal is cut and dropped into a trough containing water;
- The resultant solution is tested with litmus paper;

Diagram of apparatus:

Observations and explanations:

- The metal floats on the water surface; because it is less dense than water;
- A hissing sound is produced; due to production of hydrogen gas;
- It vigorously melts into a silvery ball then disappears because reaction between water and sodium is exothermic (produces heat). The resultant heat melts the sodium due to its low melting point.
- It darts on the surface; due to propulsion by hydrogen;
- The metal may burst into a golden yellow flame; because hydrogen may explode into a flame which then burns the sodium;
- The resultant solution turns blue; because sodium hydroxide solution formed is a strong base;

(b). Reaction equations.

Equation I

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

Equation II

$4Na_{(s)} + O_{2\,(g)} \rightarrow 2Na_2O_{(s)}$;

Equation III:

$Na_2O_{(s)} + H_2O_{(l)} \rightarrow 2NaOH_{(aq)}$

Effect of resultant solution on litmus paper;

- Litmus paper turns blue; sodium hydroxide formed is highly soluble in water; releasing a large number of hydroxyl ions which result into alkaline conditions // high pH;

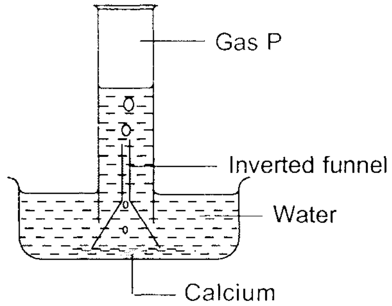
3. Calcium.

Procedure:

- A small piece of calcium metal is cut and dropped into a trough containing water;
- A filter funnel is inverted over it;
- A test tube filled with water is then inverted over the funnel;
- The gas given out is collected as shown in the apparatus below.
- The resultant gas is then tested with a burning splint;

- The resultant solution in the trough is tested with litmus paper.

Diagram of apparatus:



Observations and explanations:

- Calcium sinks to the bottom of the beaker; because it is denser than water;
- Slow effervescence of a colourless gas; due to slow evolution of hydrogen gas;
- Soapy solution formed; due to formation of alkaline calcium hydroxide;
- A white suspension is formed; because calcium hydroxide is slightly soluble in water;

Reaction equation:

$Ca_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$;

Effect of resultant solution on litmus paper;

- Litmus paper slowly turns blue; calcium hydroxide formed is slightly soluble in water; releasing a small number of hydroxyl ions which result into alkaline conditions // high pH;

4. Magnesium and other metals.

- Magnesium reacts with atmospheric oxygen to form magnesium oxide that coast the metal surface;
- Thus before reacting it with water this oxide layer has to be removed e.g. by polishing metal surface using sand paper;
- Reaction between magnesium and cold water is generally very slow; with very slow evolution of hydrogen gas;
- Zinc and iron metals do not react with cold water;

Reaction of metals with steam.

Note:

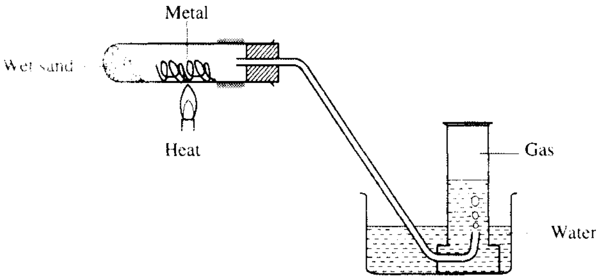
- Metals that react with cold water would react very explosively with steam and thus their reactions with steam should not be attempted in the laboratory;
- However some metals which react only sparingly with cold water or do not react with cold water at all react with steam to produce respective metal oxide and hydrogen gas;

1. Magnesium

Procedure:

- A small amount of wet sand is put at the bottom of a boiling tube;
- A small piece of magnesium ribbon is cleaned and put in the middle of the combustion tube;
- The magnesium ribbon is heated strongly then the wet sand is warmed gently;
- The delivery tube is removed before heating stops; and the gas produced is tested using a burning splint;

Diagram of apparatus:



Observations and explanations.

- Magnesium burns with a bright blinding flame;
- Grey solid (magnesium) forms a white solid; due to formation of magnesium oxide;
- Evolution of a colourless gas that burns with a pop sound; confirming it is hydrogen;

Reaction equation.

Magnesium + Steam → Magnesium oxide + Hydrogen gas;

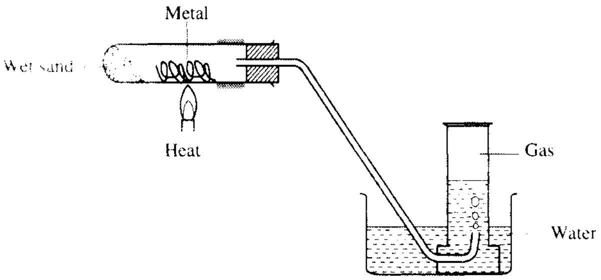
$Mg_{(s)} + H_2O_{(g)} \rightarrow MgO_{(s)} + H_{2(g)}$;

2. Zinc

Procedure:

- A small amount of wet sand is put at the bottom of a boiling tube;
- A small piece of zinc put in the middle of the combustion tube;
- The zinc is heated strongly then the wet sand is warmed gently;
- The delivery tube is removed before heating stops; and the gas produced is tested using a burning splint;

Diagram of apparatus:



Observations and explanations.

- Zinc metal does not burn but rather glows;
- Grey solid (zinc) forms a yellow solid which cools to a white solid (zinc oxide);
- Evolution of a colourless gas that produces a pop sound when exposed to a burning splint; confirming it is hydrogen;

Reaction equation.

Zinc + Steam → Zinc oxide + Hydrogen gas;

Grey Colourless Yellow when hot Colourless
White on cooling

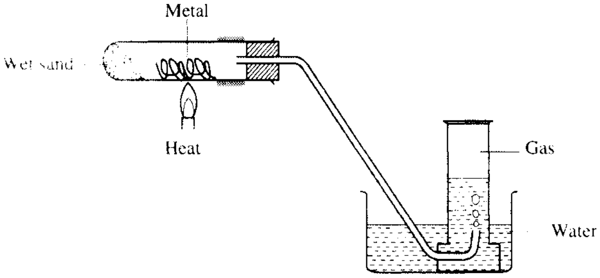
$Zn_{(s)} + H_2O_{(g)} \rightarrow ZnO_{(s)} + H_{2(g)}$;

3. Iron

Procedure:

- A small amount of wet sand is put at the bottom of a boiling tube;
- A small piece of iron put in the middle of the combustion tube;
- The iron is heated strongly then the wet sand is warmed gently;
- The delivery tube is removed before heating stops; and the gas produced is tested using a burning splint;

Diagram of apparatus:



Observations and explanations.

- Iron metal does not burn but rather glows;
- Grey solid (zinc) forms a **black solid**; due to formation of **tri-iron tetra-oxide**;
- Evolution of a colourless gas that burns with a pop sound; confirming it is hydrogen;

Reaction equation.

Iron + Steam → Tri-iron tetra-oxide + Hydrogen gas;

Grey Colourless Black Colourless

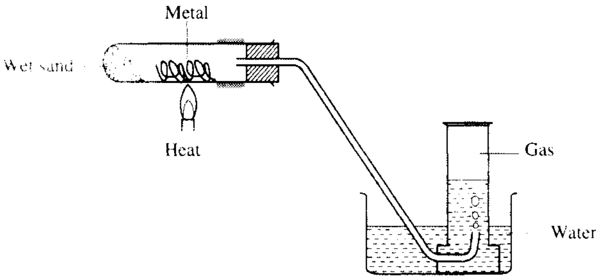
$3Fe_{(s)} + 4H_2O_{(g)} \rightarrow Fe_3O_{4(s)} + 4H_{2(g)}$;

4. Aluminium

Procedure:

- A small amount of wet sand is put at the bottom of a boiling tube;
- A small piece of aluminium put in the middle of the combustion tube;
- The aluminium is heated strongly then the wet sand is warmed gently;
- The delivery tube is removed before heating stops; and the gas produced is tested using a burning splint;

Diagram of apparatus:



Observations and explanations.

- Aluminium burns in steam but the reaction quickly stops; because the reaction forms a layer of aluminium oxide that coats the metal surface preventing further reaction;
- Grey solid (aluminium) forms a white solid of aluminium oxide;
- Slight evolution of a colourless gas that burns with a pop sound; confirming it is hydrogen;
- The production of the gas however stops soon after the reaction starts because the oxide layer stops further reaction;

Reaction equation.

Aluminium + Steam → Zinc oxide + Hydrogen gas;
Grey ***Colourless*** ***White*** ***Colourless***

$2Al_{(s)} + 3H_2O_{(g)} \rightarrow Al_2O_{3(s)} + 3H_{2(g)}$;

5. Other metals.

- Lead and copper do not react with steam;

Summary of the reaction between metals with cold water and steam

Metal	Action of metal on water	Action of metal on steam
Potassium	Violent	Explosive
Sodium	Violent	Explosive
Calcium	Moderate	Violent
Magnesium	Very slow	Rapid
Aluminium	No reaction	Slow
Zinc	No reaction	Slow
Iron	No reaction	Slow
Lead	No reaction	No reaction
Copper	No reaction	No reaction

Note:

- Metals can thus be arranged in order of their reactivities with water; resulting to a **reactivity series** similar to that obtained form reaction between metals with oxygen;

Reactivity series of metals:

Potassium; ↑ Most reactive;
Sodium;
Calcium;
Magnesium
Aluminium;
Zinc; *increasing reactivity;*
Iron;
Lead;
Copper; Least reactive;

Hydrogen

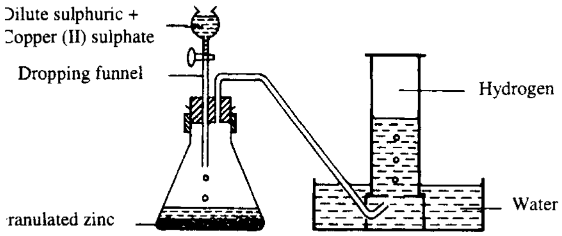
- An element that does not exist freely in nature;
- Generally exists in compounds such as water, sugars, fuels etc;

Laboratory preparation of hydrogen gas.

Note:

- Hydrogen gas is generally prepared by the reaction between dilute acids and metals;
- Most suitable acids are dilute hydrochloric acid and dilute sulphuric (VI) acid;
- Most suitable metal is zinc metal;

Apparatus:



Procedure:

- Zinc granules are added to dilute sulphuric (VI) acid;
- Small amounts of copper (II) sulphate are added to the zinc – acid mixture;

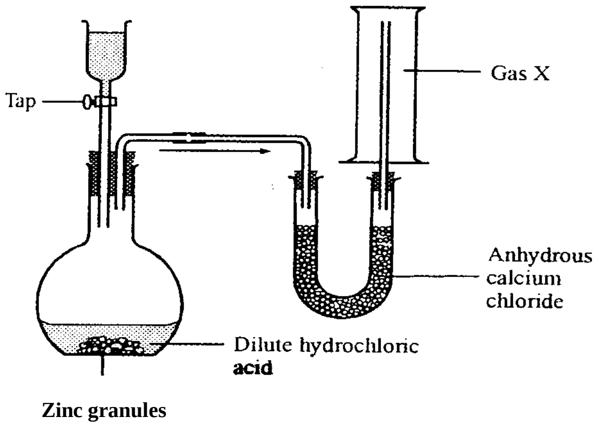
Reason: To act as a catalyst hence speed up the reaction;

- The resultant colourless gas is collected over water;

Reason: The gas is insoluble in water;

- If the gas is required dry, the gas is passed through concentrated sulphuric (VI) acid or a U-tube containing calcium chloride;

Diagrams for drying the gas:



- The dry gas is collected by upward delivery (downward displacement of air);
Reason: It is less dense than air (note that hydrogen is the lightest gas known);

Note:
- Nitric (V) acid is not used in preparation of hydrogen gas; except very dilute nitric (V) acid and magnesium
Reason: Nitric acid is a strong oxidizing agent hence the hydrogen formed is immediately oxidized to water
- Potassium, sodium, lithium and calcium are not used in laboratory preparation of hydrogen gas;
Reason: They react explosively with acids;
- Magnesium is not usually used for laboratory preparation of hydrogen;
Reason: It is expensive;
- Prior to using it for preparation of hydrogen; aluminium should be washed with concentrated hydrochloric acid;
Reason: To remove the protective oxide layer that usually forms on the aluminium surface on its exposure to air;
- Hydrogen gas produced from iron metal tends to have a foul smell;
Reason: Iron gives a mixture of gases due to impurities in the iron; the foul smell is usually due to production of hydrogen sulphide that results from sulphide impurities in the iron metal;

Reaction equations:
I. Iron + Sulphuric (VI) acid → Iron (II) sulphate + Hydrogen gas;
$Fe_{(s)} + H_2SO_{4(aq)} \rightarrow FeSO_{4(aq)} + H_{2(g)}$

II. Iron (II) sulphide + Sulphuric (VI) acid → Iron (II) sulphate + Hydrogen sulphide gas;
$FeS_{(s)} + H_2SO_{4(aq)} \rightarrow FeSO_{4(aq)} + H_2S_{(g)}$;

Other sources of hydrogen.
- Cracking of alkanes;
- Fractional distillation of petroleum;

Properties of Hydrogen gas;

Physical properties.
1. Colourless and odourless;
2. It is insoluble in water;
3. It is less dense than air (it is the lightest gas known);

Chemical properties.
1. It has no effect on litmus paper implying that it is neutral;
2. It burns with a pop sound when mixed with air;
3. Pure hydrogen burns quietly with a blue flame producing water;
4. Hydrogen does not support combustion but it burns in air;
5. It is a reducing agent;

Chemical test for hydrogen;

- When a burning splint is introduced into a gas jar full of hydrogen gas; the gas buns with a "pop sound"

Note:

- The intensity of the "pop sound" diminishes as the purity of hydrogen increases;

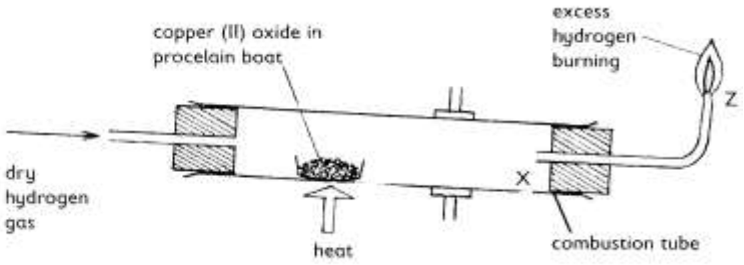
Reactions of hydrogen:

1. Passing hydrogen over heated copper (II) oxide;

Requirements:

Copper (II) oxide; porcelain boat; mean of heating; combustion tube; dry hydrogen gas;

Diagram of apparatus.



Procedure:

- Apparatus are arranged as shown above;
- Dry hydrogen gas is passed through the combustion tube for sometime prior to heating the oxide;

Reason:

- ✔ To drive out all the air from the apparatus which may otherwise re-oxidize the metal after reduction.
- ✔ A mixture of hydrogen and air will explode when the combustion tube is heated;

- The gas is continuously collected at the jet and tested; until the gas burns smoothly without a "pop" sound;
- The gas is then lit at the jet and the copper oxide heated;
- This is done until no further change;
- The apparatus is allowed to cool as hydrogen is still continuously allowed to pass through;

Reason:

- ✔ To prevent re-entry of atmospheric oxygen which will re-oxidize the hot metal back to the metal oxide;

- Excess hydrogen must be burnt at the jet so that excess gas is not allowed to escape into the air;

Reason:

- ✔ The mixture of hydrogen and oxygen is explosive when ignited;

Observations and explanations

- The black solid turns into a brown solid;
- Droplets of a colourless liquid on the cooler parts of the combustion tube;

Reason:

The hot black copper (II) oxide is reduced by hydrogen gas into brown copper metal while hydrogen gas is oxidized to water;

- The colourless liquid is confirmed to be water by:
  - ✔ Adding drops of it to anhydrous copper (II) sulphate which forms a blue solid;
  - ✔ Adding drops of it onto blue anhydrous cobalt (II) chloride which turns pink;

Reaction equations:

In the combustion tube:

Copper (II) oxide + Hydrogen → Copper + Water;

Black solid ***Colourless gas*** ***Brown solid*** ***Colourless liquid***

$CuO_{(s)} + H_{2(g)} \rightarrow Cu_{(s)} + H_2O_{(l)}$;

At the jet:

- Hydrogen burns with a blue flame producing water;

Equation:

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

Conclusion:

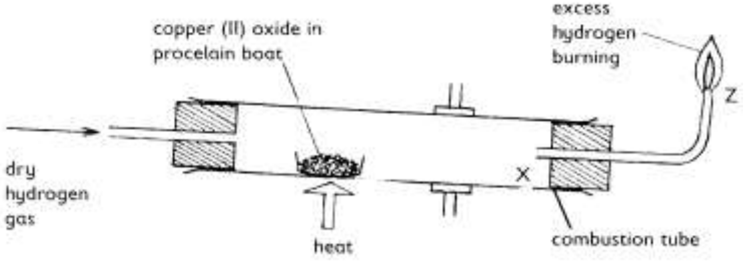
- Hydrogen is a reducing agent and thus reduces the copper (II) oxide to copper metals while hydrogen itself undergoes oxidation to form water;
- Hydrogen is the reducing agent;
- Copper (II) oxide is the oxidizing agent;

2. Passing hydrogen over heated Lead (II) oxide;

Requirements:

Lead (II) oxide; porcelain boat; mean of heating; combustion tube; dry hydrogen gas;

Diagram of apparatus.



Procedure:

- Apparatus are arranged as shown above;
- Dry hydrogen gas is passed through the combustion tube for sometime prior to heating the oxide;

Reason:

✔ To drive out all the air from the apparatus which may otherwise re-oxidize the metal after reduction.

✔ A mixture of hydrogen and air will explode when the combustion tube is heated;

- The gas is continuously collected at the jet and tested; until the gas burns smoothly without a "pop" sound;
- The gas is then lit at the jet and the copper oxide heated;
- This is done until no further change;
- The apparatus is allowed to cool as hydrogen is still continuously allowed to pass through;

Reason:

✔ To prevent re-entry of atmospheric oxygen which will re-oxidize the hot metal back to the metal oxide;

- Excess hydrogen must be burnt at the jet so that excess gas is not allowed to escape into the air;

Reason:

✔ The mixture of hydrogen and oxygen is explosive when ignited;

Observations and explanations

- The yellow solid turns red on heating then finally into a grey solid;
- Droplets of a colourless liquid on the cooler parts of the combustion tube;

Reason:

The hot red lead (II) oxide is reduced by hydrogen gas into grey lead metal while hydrogen gas is oxidized to water;

- The colourless liquid is confirmed to be water by:

✔ Adding drops of it to anhydrous copper (II) sulphate which forms a blue solid;

✔ Adding drops of it onto blue anhydrous cobalt (II) chloride which turns pink;

Reaction equations:

In the combustion tube:

Lead (II) oxide + Hydrogen → Lead + Water;

Yellow-cold; red - hot   Colourless gas   Grey solid   Colourless liquid

$PbO_{(s)} + H_{2(g)} \rightarrow Pb_{(s)} + H_2O_{(l)}$;

At the jet:

- Hydrogen burns with a blue flame producing water;

Equation:

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

Conclusion:

- Hydrogen is a reducing agent and thus reduces the lead (II) oxide to lead metals while hydrogen itself undergoes oxidation to form water;

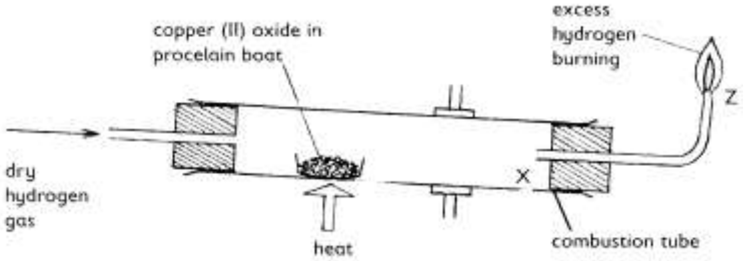
- Hydrogen is the reducing agent;
- Lead (II) oxide is the oxidizing agent;

3. Passing hydrogen over heated Iron (III) oxide;

Requirements:

Iron (III) oxide; porcelain boat; mean of heating; combustion tube; dry hydrogen gas;

Diagram of apparatus.



Procedure:

- Apparatus are arranged as shown above;
- Dry hydrogen gas is passed through the combustion tube for sometime prior to heating the oxide;

Reason:

- ✔ To drive out all the air from the apparatus which may otherwise re-oxidize the metal after reduction.
- ✔ A mixture of hydrogen and air will explode when the combustion tube is heated;

- The gas is continuously collected at the jet and tested; until the gas burns smoothly without a "pop" sound;
- The gas is then lit at the jet and the copper oxide heated;
- This is done until no further change;
- The apparatus is allowed to cool as hydrogen is still continuously allowed to pass through;

Reason:

- ✔ To prevent re-entry of atmospheric oxygen which will re-oxidize the hot metal back to the metal oxide;

- Excess hydrogen must be burnt at the jet so that excess gas is not allowed to escape into the air;

Reason:

- ✔ The mixture of hydrogen and oxygen is explosive when ignited;

Observations and explanations

- The Brown solid turns into a grey solid;
- Droplets of a colourless liquid on the cooler parts of the combustion tube;

Reason:

The hot brown iron (III) oxide is reduced by hydrogen gas into grey iron metal while hydrogen gas is oxidized to water;

- The colourless liquid is confirmed to be water by:

- ✔ Adding drops of it to anhydrous copper (II) sulphate which forms a blue solid;
- ✔ Adding drops of it onto blue anhydrous cobalt (II) chloride which turns pink;

Reaction equations:

In the combustion tube:

Iron (III) oxide + Hydrogen → Iron + Water;

Brown solid ***Colourless gas*** ***Grey solid*** ***Colourless liquid***

$Fe_2O_{3(s)} + 3H_{2(g)} \rightarrow 2Fe_{(s)} + 3H_2O_{(l)}$;

At the jet:

- Hydrogen burns with a blue flame producing water;

Equation:

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

Conclusion:

- Hydrogen is a reducing agent and thus reduces the iron (III) oxide to iron metal while hydrogen itself undergoes oxidation to form water;
- Hydrogen is the reducing agent;
- Iron (III) oxide is the oxidizing agent;

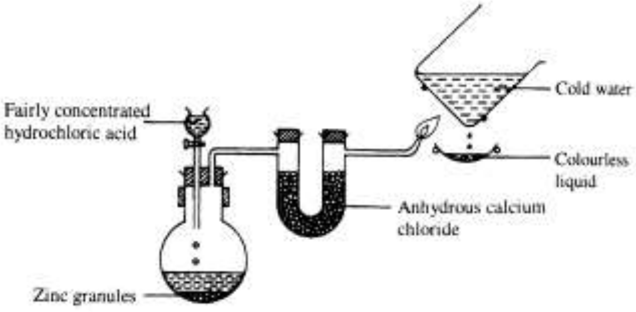
Note:

- Hydrogen does not reduce (remove oxygen) from oxides of metals above it in the reactivity series;

Products of burning hydrogen gas in air.

Apparatus and requirements:

- The apparatus is arranged as shown below.



Procedure:

- Apparatus is arranged as shown below.
- A stream of hydrogen is passed through anhydrous calcium chloride;
- The gas is tested for purity by collecting samples over the jet and testing with a burning splint.

Note: pure hydrogen gas should burn smoothly without the typical “pop sound”;

- The gas is then lit and the pump tuned on;

- The products of burning hydrogen is drown in through the apparatus using the pump for about 15 minutes;
- The product condensing in the test tube in cold water is tested with white anhydrous copper (II) sulphate and blue cobalt chloride paper;

Observations:
- Pure hydrogen burns with a blue flame;
- A colourless liquid condenses in the test tube immersed in cold water;
- The liquid turns white anhydrous copper (II) sulphate blue;
- The colourless liquid turns blue anhydrous cobalt chloride pink;

Explanations:
- The calcium chloride in the U-tube is used to dry the gas;
- The pure dry hydrogen gas burns with a blue flame to form steam which condenses into liquid water;
- Water turns anhydrous copper (II) chloride from white to blue; and turns blue anhydrous cobalt chloride paper into pink;

Reaction equations:
At the jet:
$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$;

Uses of Hydrogen
1. Large scale manufacture of ammonia in the **Haber process**;
2. **Hydrogenation** for the manufacture of margarine.
- This refers to hardening of oils into fats.
- In this reaction Hydrogen gas is bubbled into liquid oil in presence o0f nickel catalyst;
- The oil takes up hydrogen and is converted into fat;
3. Hydrogen is used in weather balloons because it is lighter than air;
- Usually a radio transmitter is connected to a weather balloon filled with air; as the balloon floats in air the transmitter collects information which is conveyed to weather stations for interpretation by meteorologists;
4. A mixture of hydrogen and oxygen forms the very hot oxy-hydrogen flame (with temperatures up to 2000^oC that is used in welding and cutting metals;
5. It is used in rocket fuels for propulsion of rockets;
6. Manufacture of hydrochloric acid; during which hydrogen is burnt in chlorine;