Illustrative CHEMISTRY FOR SECONDARY SCHOOLS

FORMS 1 & 2

JOHN MESHACK MANDIGA

Dip(Ed), Bsc(Ed) with Honours

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P.O.Box 30484
Vikuge - Kongowe
Kibaha - Pwani
Mobile:
+255 (0)765 883 640
+255 (0)654 130 099
+255 (0)788 573 277
Email: mandiga1976@yahoo.com

Cover design, photos and typsetting and graphics design by
Michael Ivan Senyonga
Melany Stationaries
Dar es salaam - Tanzania.
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TABLE OF CONTENTS

Chapter 13 - Periodic classification

Chapter 14 - Formula, bonding and nomenclature

PREFACE

Illustrative Chemistry for Secondary Schools: Forms 1 and 2, is a book that has been written in such a way that is easily readable and understandable to learners. This book covers Chemistry topics for forms 1 and 2 as prescribed in Tanzania syllabus.

Chemistry knowledge and skills can easily be acquired through reading this book. Experiments, activities and exercises are well covered in this book to cater for learner-centered approach. This book covers the following topics:

1. Introduction to Chemistry
2. Laboratory techniques and safety
3. Heat sources and flames
4. The scientific procedure
5. Matter
6. Elements, Compounds and mixtures
7. Air, Combustion, rusting and fire fighting
8. Oxygen
9. Hydrogen
10. Water
11. Fuels and energy
12. Atomic structure
13. Periodic classification
14. Formula, bonding and Nomenclature

I wish you all the best as you interact with this book.
John Meshack Mandiga,
Kibaha- Pwani,
June - 2013.

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Kibaha - Pwani,
June - 2013

Chapter 1

Introduction to Chemistry

1. 1 THE CONCEPT OF CHEMISTRY

The word ***Chemistry*** comes from the word ***Alchemy.*** Alchemy is an earlier set of practices that encompassed elements of Chemistry, metallurgy, philosophy, astrology, mysticism and medicine. The word alchemy in turn is derived from the Arabic word ***"al-kimia"*** or "***kimi"*** which is Egyptian name, meaning "***cast together"***.

By definition, Chemistry is the branch of science which deals with composition, decomposition, structure and properties of matter. The people who study Chemistry are called ***Chemists***. Hundreds of years ago people were interested in what we call Chemistry. These people are called ***alchemists***.

1.2. BRANCHES OF CHEMISTRY

There are many branches of Chemistry. In this level of study the five (5) major branches of Chemistry will be considered,

- ***Organic Chemistry:*** is the study of carbon and its compounds.
- ***Inorganic Chemistry:*** is the study of other substances not containing carbon.
- ***Analytical Chemistry:*** is the study of separation, identification and composition of materials and the development of tools used to measure properties of matter.
- ***Physical Chemistry:*** is the study of the physical characteristics of materials and the mechanisms of their reactions.
- ***Biochemistry:*** is the study of chemical processes that occur inside of living organisms.

1.3 APPLICATION OF CHEMISTRY

Most of the products we use are produced through application of Chemistry. There is no way in life we can interact with industrial and non-industrial products and do away with the fact that the knowledge of Chemistry is applied in order to get those products. Examples of products made through application of Chemistry are fertilizers, pesticides, drugs, vaccines, detergents, toothpaste, insecticides, soft drinks, common salt, paints, cement, clothes, fuels, lubricants and grease.

Chemistry is an important subject that is applied in different fields such as mining, agriculture, medicine, manufacturing, education, food and beverage industry, home care and cosmetics industry, film industry, water treatment and purification. The places where Chemistry is applied among others are hospitals, homes, factories, laboratories, research centres, universities,water treatment plants and mining centres. In our homes, Chemistry may be used in baking, cooking and washing clothes with soap.

1.4 IMPORTANCE OF CHEMISTRY

The application of Chemistry knowledge directly relates with the importance of Chemistry. Chemistry is applied in different fields to produce very important products or rendering helpful services.

1. WATER TREATMENT

- The knowledge of Chemistry is used to produce chemicals like ***water guard*** that kill germs present in water.
- ***Liquid chlorine:*** is another chemical which is added in water to kill harmful bacteria. Chlorine is a useful disinfectant that is used in swimming pools to kill bacteria. Potassium Aluminum Sulphate is a chemical which when added into water fine particles found in water settles down to allow the process of ***sedimentation***.

2. TRANSPORT AND COMMUNICATION

- Liquid fuels like gasoline (petrol), diesel and kerosene are used by different means of transport. These fuels are produced by chemical processes
- The gaseous fuels like natural gas (methane), liquefied petroleum gas (LPG) are also processed by using the knowledge of Chemistry.
- Non-petroleum fossils like biodiesel and alcohols are produced through application of Chemistry.
- Chemical processes are applied to produce papers, and wires. These products are very essential in communication.

3. AGRICULTURE

- Agriculture is the science of livestock keeping and production of crops.
- Products like ***pesticides, insecticides, herbicides, fungicides, fertilizers, hormones, and growth agents*** are important in the field of agriculture. All the named products are produced through chemical processes. Farmers use the products

in order to get better agricultural yields.

- Other products used by farmers are ***weed killers*** and ***animal vaccines.***

4. FOOD AND BEVERAGE INDUSTRY

- Biochemical products like carbohydrates, lipids, proteins, vitamins, minerals, enzymes, food additives, flavors and colours are produced through chemical processes.
- The food industry has also benefited from chemical processes like ***food preservation.*** Food preservation can be employed industrially in canning and bottling of foods.
- Soft drinks like coca cola and Pepsi are produced through chemical processes like ***carbonation***.
- Carbonation is also involved in production of ***beer, wine, tonic water*** and many others.

5. MANUFACTURING INDUSTRY

- Manufacturing is the production of goods for use or sale using labor and machines, tools, chemical and biological processing or formulation. Manufacturing relies on Chemistry and its chemical processes. The raw materials and products depend much on Chemistry knowledge.
- Products like cement, cars, plastic containers, chemicals, textiles, paper, rubber, glass, computers, mobile phones and many others are produced by application of Chemistry knowledge.

6. HOME CARE AND COSMETICS INDUSTRY

- Home care products like soaps, detergents, disinfectants, air fresheners, paints, polish and vacuum cleaner are used to make the home and its surrounding cleaner and more comfortable to live. All these products are made through application of chemical processes.
- Cosmetics like ***lip stains, face powder, bronzer, eye lash curler, nail polish, eye liner, lotions and creams*** are produced chemically. Other beauty products are ***deodorants*** and ***primers.***

7. MEDICINE

- Chemically produced substances like drugs, vaccines and food supplements are very important in our lives. Prevention of illness and diseases, treatment of diseases ensures our well-being.
- In general the field of medicine backed up by Chemistry knowledge grants us with healthier living.

8. FILM INDUSTRY.

- A film camera catches the picture using chemicals on film. The first popular photographs were captured on copper plates in the 1840s.
- Video recorder, digital cameras and video tapes are produced by the aid of Chemistry knowledge.
- A piece of film consists of a light sensitive ***emulsion*** applied to a tough, transparent base. The emulsion consists of silver halide grains suspended in a gelatin colloid, in the case of colour film.
- Development chemicals applied to an appropriate film can produce either a positive (showing the same densities and colours as the subject) or negative image (with dark highlights, light shadows).

9. MINING INDUSTRY

- Mining is the extraction of valuable minerals or other geological materials from the earth. Extraction metallurgy is the practice of removing valuable metals from an ore and refine the extracted raw materials into pure form. Chemical processes especially electrolysis are very essential in converting metal oxides or sulphides into a pure metal.
- The knowledge of Chemistry is very important in all these processes.

1.5 CHEMISTRY FOR PROFESSIONAL DEVELOPMENT

The study of Chemistry is very important in many professions or careers. Professionals like geologists, engineers, nurses, medical doctors, farmers, horticulturalists, floriculturists, pharmacists, laboratory technicians, researchers and science teachers depend on the knowledge of Chemistry in their careers. Through studying Chemistry, skills are acquired by such professionals

SUMMARY

(a). In Chemistry we study the substances which make up the earth, the living things and the universe in general.
(b). People who study Chemistry are called chemists.
(c). People who were interested in Chemistry hundreds of years ago are called alchemists.
(d). Chemistry deals with the composition, decomposition, structure and properties of matter.
(e). Major five branches of Chemistry are organic Chemistry, inorganic Chemistry, analytical Chemistry, physical Chemistry and biochemistry.
(f). The importance of Chemistry is revealed in such fields as medicine, agriculture, mining Industry, water treatment, film industry, home care and cosmetics industry, transport and communication, food and beverage industry and manufacturing industry.

REVIEW QUESTIONS

1. Match the items in list A with their corresponding statements in LIST B.

LIST A

i. Agriculture ....
ii. Medicine ....
iii. Manufacturing industry ....
iv. Transport industry ....
v. Food and beverage industry ....

LIST B

A. Clothes, dyes
B. Fertilizers, pesticides, weed killers, animal vaccines.
C. Fuels, lubricants, oil, grease, coolant, drugs, animal vaccines, food supplements.
D. Paints, chemicals, vanishes, cement, plastics
E. Detergents, beauty products, shoe polish, tooth paste, disinfectants, insecticides.
F. Drugs, vaccines, food supplements
G. Soft drinks, common salt, yeast, baking powder, canned food.

2. Write ***TRUE*** or ***FALSE*** for the following statements.

i. Matter is anything that has mass and occupies space ....
ii. The people who study Chemistry are called chemists............................
iii. We cannot apply Chemistry in our homes when we are baking, cooking

or even washing our clothes with soap ....

iv. The study of Chemistry is important in professional development ....

v. Few materials and objects are made by application of chemical processes...

3. Choose the most correct answer for the following questions.

(i) .... might be involved in finding minerals and crude oil in rocks and in their exploitation.

A. Biologists B. Mathematicians
C. Geologists D. Alchemists

(ii) Letters, newspapers and magazines are written on ....that is manufactured chemically.

A. Rocks B. Plastic
C. Black board D. Paper.

(iii) Chemically produced substances in the form of..................................and ....are used in the field of medicine.

A. Vaccine, medicines B. Medicines, immunity
C. Vaccines, fruits D. Fertilizers, vaccines

(iv) The method of protecting food from spoilage is called.............................

A. Food protection B. Food preservation.
C. Food additives D. Food canning.

(v). Chemical substances that are used to destroy unwanted plants which are harmful to crops are called

A. Pesticides B. Insecticides
C. Weed killers D. Animal vaccines.

4. Mention seven examples of substances made by applying chemical methods.

5. One of the major problems in the world today is the spread of HIV/AIDS. Chemists are working to develop new drugs that will combat HIV/AIDS. Make a list of other problems that we face in the world today. Can you suggest anything that Chemistry might be able to do to solve them?

6. Briefly explain how you can apply the knowledge of Chemistry at home.

Chapter 2

Laboratory Techniques and Safety

2.1 INTRODUCTION

The term ***laboratory*** come from the same Latin word as the English word labour, which means ***hard work***.

By definition, a laboratory is a special room or building that is designed and used for scientific experiments. Chemical experiments are carried out in a ***chemistry laboratory***.

2.2 RULES AND SAFETY PRECAUTIONS IN A CHEMISTRY LABORATORY

The laboratory is meant to be a quiet and safe place to work in . In view of this, safety in the laboratory is of great importance to both students and teacher. The laboratory can be very dangerous place if safety regulation is not followed.

LABORATORY RULES.

Laboratory rules are specified guidelines required to be followed when working with the laboratory. Here is a set of rules to be followed when dealing with the laboratory.

- Never enter the laboratory without the permission or presence of the teacher.
- Always dress appropriately for the laboratory activities
- Always keep the windows open for proper ventilation
- Always master the location of all exit
- Always read instruction carefully before you start any experiment or activity.
- Never run in the laboratory
- Never eat or drink anything in the laboratory
- Never quarrel or fight in the laboratory
- Never use laboratory apparatus for drinking or storing food
- Never taste or sniff chemicals unless advised on how it should be done
- Never throw any solid into the sink or water ways
- Always use the fume chamber when carrying out experiments where harmful

gases are produced.

- Always wash your hands with soap and water before you leave a laboratory
- Always perform the intended experiments.
- Always replace covers and stoppers on the container after using the chemicals.
- Never spill liquids on the floor.
- Always keep the gangways and exits clear.
- Always report any breakages or accident to the teacher immediately.
- Always keep your bench top clean, dry and well arranged.
- Always direct the mouth of the test tube away from you or others when heating substances.
- Always use a clean spatula to remove chemicals from containers.
- Always use a lighter or wooden splints to light burners, remember to strike the match before turning on the gas tap.
- Never touch electrical appliance with wet hands.
- Always turn off any gas or water taps that are not in use.
- Never use dirty, cracked or broken apparatus.
- Never heat flammable liquids with a Bunsen burner flame.
- Never remove chemicals or equipment from the laboratory.
- Always wash off any chemical spillage on your skin or clothes with plenty of water.
- Always keep inflammable substances away from naked flames.
- Never casually dispose of chemicals and wastes inappropriately.
- Always clean up the equipment and store properly after use.

LABORATORY SAFETY MEASURES

Laboratory safety measures are precautions undertaken in order to minimize risks when carrying out laboratory activities. Laboratory safety measures can be explained as follows:

1. The laboratory should be equipped with protective clothes like laboratory coat and safety glasses all the time
2. All people using the laboratory should not put themselves or anyone else in danger by either smoking, causing panic incase of an unwanted chemical reaction, damage or injury .The supervisor should be called in case of such situation.
3. The fume cupboard should be used whenever working with hazardous or toxic substances and for every experiment where easily evaporable chemicals are employed
4. All flammable substances should not be exposed to open flame
5. When working with concentrated acids or alkalis protective gloves

and safety glasses should be weared.

6. Immediately rinse with water and neutralize if necessary in case a chemical comes in contact with your skin. In case a chemical come in contact with your eyes wash with running water and call the supervisor.
7. A pipette bulb should be used instead of mouth suctions when pipetting harmful or evaporable substance.
8. Organic solvent of extremely hazardous and toxic substance should be disposed of into the designated waste container.
9. All flammable liquids should not be mixed with oxidizing substance as this can cause fire.
10. All chemicals that easily react with each other should be stored separately.
11. There should be equipment for monitoring contamination in order to give alert of any possible dangers in the laboratory.
12. Refrigerator and freezers should be used for working with laboratory chemicals only.
13. Adequate ***first aid*** kits should be available in the laboratory
14. There should be clear instruction on how to use fire extinguishers in case of fire.
15. All persons working in the laboratory should not accidentally get into contact with harmful chemicals by ensuring that cupboards, storage cabinet and drawers have locks.
16. There should be regular inspection and checking of stored chemicals in order to avoid using expired substances.
17. Chemical containers should be stoppered and regularly checked to ensure that they do not leak.
18. Accidental use of wrong substances should be avoided by labeling all chemicals.
19. All spills should be cleaned immediately.
20. There should be emergency exits and easily accessible.
21. Gas cylinders should be in good working condition always, they should be labelled, stored and well supported.

2.3 CHEMICAL WARNING SIGNS

Chemical warning signs are safety symbols found on chemical containers especially those used in the laboratory to warn about hazardous materials, location or object.

All chemicals are potential poisons. Some chemicals are dangerously lethal. Others harmful or toxic. The chemicals we meet in the school laboratory are not all completely safe. The containers of modern chemicals carry special chemical warning signs.

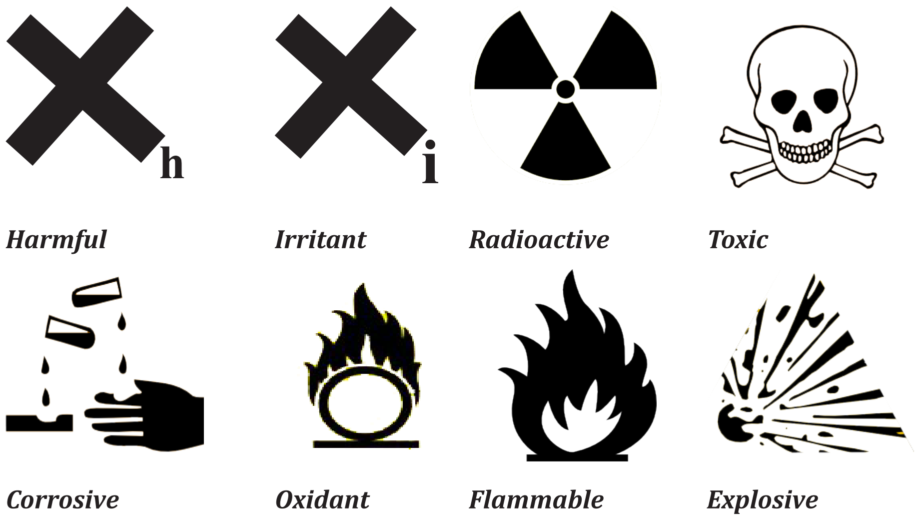


Fig 2.1. Chemical warning signs

EXPLOSIVE

Explosive substances are usually detonated before they explode. An explosion is forceful rapid reaction which involve throwing of particles at high speed. Explosives which can react without external detonation are particularly dangerous. Chemicals carrying the sign of explosive may cause explosion if they are not handled carefully and according to instruction. It is dangerous to keep explosives in glass containers. In case the explosion occurs glass particles will fly around and cause injuries to people.

TOXIC

Toxic substances can poison people. They can cause death immediately or after few days. They are most dangerous when they get into your body. They should not be allowed to enter your body through the mouth, nose, eyes, skin or ears. Some chemicals in this group are so dangerous that they can enter you body through the skin. If that happens, wash it out with plenty of water.

RADIOACTIVE

Radioactive substance can emit harmful radiations. Radioactive materials contain unstable elements such as uranium and plutonium. Radioactive material can be in the form of open sources or sealed source. An open source of radioactive material is normally used as a tracer in experiments and has the potent ion for spillage and release if not properly contained. A sealed source is in a form that is permanently bonded or fixed in a capsule or matrix designed to prevent release of radioactive material

FLAMMABLE

Flammable chemicals can catch fire easily. They should never be brought near open flames. If they have to be heated at all, an electric heater may be used. Any type of sparks may set these chemicals on fire, all burners must be put off before working with flammable chemicals. Flammable chemicals always evaporate fast. Their containers must be stoppered immediately after use.

HARMFUL OR IRRITANT

Harmful substances may make you sick or endanger your health. They will not kill you instantly but may affect you after a long exposure. There is always a danger of ignoring harmful chemicals because they are not instantly lethal. It is advised that they should be handled according to the safety instruction provided . Irritating substances can cause pains on your skin or eyes. They can endanger your health if they come in contact with your skin or eyes for too long

OXIDANT

This is a chemical or substance which helps a burning substance to burn faster. In presence of an oxidizing agent, small fires can be made very big. Heating a mixture of an organic material with an oxidizing agent may cause an explosion. An example is the heating of potassium permanganate with saw dust.

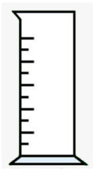
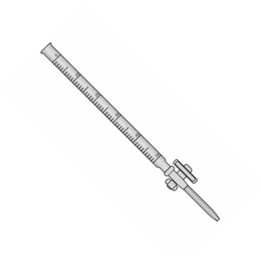
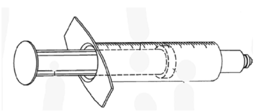
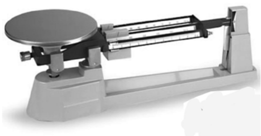
CORROSIVE

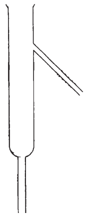
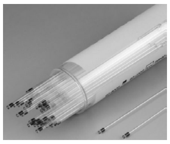
Corrosive substances can burn your skin. They can also corrode the floor and the desk top. You can turn blind if they get into your eyes. If by accident a corrosive substance comes into contact with your skin, go to the sink and wash with plenty of water. Corrosive substance may destroy metals .Concentrated mineral acids; H_2SO_4, HCl, HNO_3 and concentrated alkalis; NaOH, KOH, NH_3 are examples of very corrosive substances

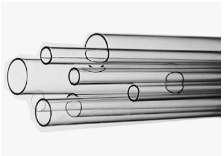
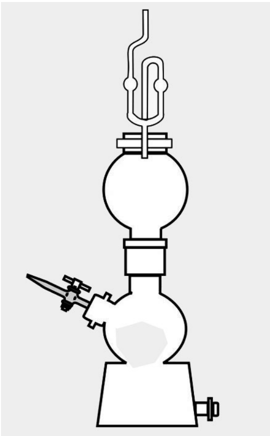
2.4 CHEMISTRY APPARATUS

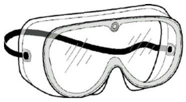
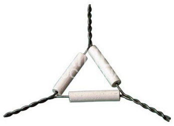
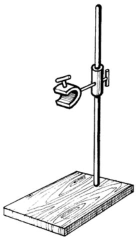
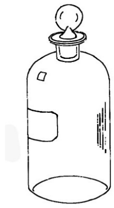
Apparatus are special tools and equipment that are used in the laboratory for various purposes such as heating, testing, measuring, filtering and grinding.

Table 2.1 showing chemistry apparatus

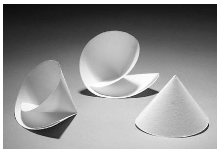
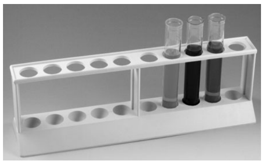
Apparatus	Diagram	Explanation
Pipette (Made of glass)		A pipette is a narrow glass tube into which small amounts of liquid are suctioned for transfer to other containers. It is used to measure specific volumes of liquid.
Measuring cylinder (Made of glass or plastic)		A measuring cylinder is a glass or plastic container that is marked to measure volumes of liquid
Thermometer (Made of glass)		A thermometer is an instrument that is used to measure and indicate the temperatures of substances
Burette (made of glass)		A burette is usually used to accurately measure and dispense liquids. It is commonly used in titrations.
Measuring Syringe (Made of plastic)		A measuring syringe is used for sucking and measuring specific volumes of liquids or gases.
Triple beam balance (Made of Iron)		A triple beam balance is a weighing instrument. It has three beams that carry weights. On one side there is a pan on which the object is placed. Each of the three weights is slid along the beams to weigh the object.

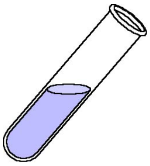
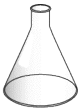
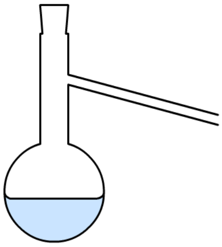
Apparatus	Diagram	Explanation
Electronic balance (Made of variety of elements)		An electronic balance is a scale that is usually used to measure the mass of chemicals. It gives more accurate readings than the beam balance.
Stop watch (Made of plastic or iron)		This is a special watch that is used to accurately time laboratory processes.
Fractionating column (made of glass)		A fractionating column is used to separate vapors of different densities.
Beehive shelf (Made of clay)		A beehive shelf is used to support a gas jar during collection of gas.
Bell Jar (Made of glass)		A bell jar is used in gas or air experiments.
Test tube brush (Made of Iron and nylon)		Test tube brush is a device made with nylon bristles attached to a twisted wire shaft, used to clean test tubes and flasks.
Corks (Made of wood)		Corks are used as stoppers for bottles
Capillary tubes (Made of glass)		Capillary tubes are used to enable small volumes of solutions to be drawn into the tube. It can also be used for stirring.

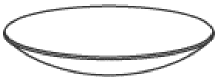
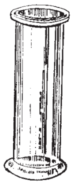
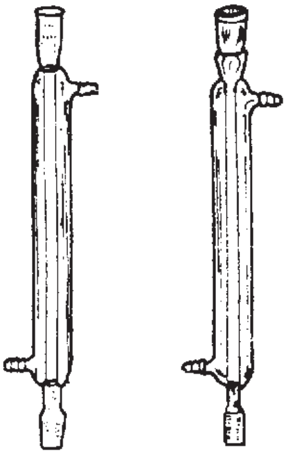
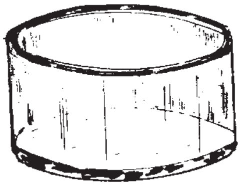
Apparatus	Diagram	Explanation
Rubber tubing (Made of rubber)		Rubber tubing are suitable for connecting apparatus and delivery of gases.
Glass tubing (Made of glass)		Glass tubing are used for blowing them in different shapes to suit in some laboratory preparations involving linkage between containers in experiments
Rubber bung (Made of rubber)		A rubber bung is used to fit in delivery tubes during experiments.
Desiccator (made of glass)		A desiccator is used for keeping substances dry
Glass rod (made of glass)		A glass rod is used for stirring substances during an experiment
Kipp's apparatus (made of glass)		A Kipp's apparatus is used for generating small amounts of gas continuously. It is also used to prepare poisonous gases.
Tweezers (Made of iron or wood)		Tweezers are tools used for picking up objects too small to be easily handled with the human hands.

Apparatus	Diagram	Explanation
Safety goggle (Made of glass or plastic)		Safety goggles are used to protect the eyes from chemical spills, strong light and harmful vapors in the laboratory.
Boss head (made of iron)		A boss head is a device that holds a clamp onto a retort stand.
Volumetric or graduated flask. (Made of glass)		A volumetric flask is used to make up a solution to a total final volume.
Pipe clay triangle (Made of Iron and clay)		A pipe clay triangle fits on the top of a tripod and support a crucible when it is heated.
Test tube holder (Made of iron and wood)		A test tube holder is an instrument that is used for holding a test tube while heating.
Retort stand and clamp (made of Iron)		Retort stand and clamp are used to hold apparatus such as burettes during experiments.
Tongs (Made of iron)		This is an instrument that is used to hold hot substances and apparatus.
Reagent bottle (Made of glass)		A reagent bottle is a bottle that is used to store different chemicals.

Apparatus	Diagram	Explanation
Plastic wash bottle (made of plastic)		This container is used to store distilled water.
Spatula (Made of iron or wood)		A spatula is used for scooping small quantities of powder or crystalline chemicals.
Thistle funnel (Made of glass)		A thistle funnel is a glass funnel with a wide top and a long stem. It is used to add reagents into flasks during experiments.
Bunsen burner (made of iron)		A Bunsen burner is a source of heat in the laboratory.
Boiling tube (Made of glass)		A boiling tube is a large test tube that is used to heat substances that should be heated strongly. It is also used when the amount of a substance is too large for a test tube.
Tripod stand (Made of Iron)		A tripod stand is a stand or support that has three legs. It is usually placed above the Bunsen burner when heating or boiling substances. A metal gauze is placed on top of the stand. The container that is used for heating is placed on the wire gauze.

Apparatus	Diagram	Explanation
Wire gauze (Made of iron)		A wire gauze is usually placed on a tripod stand. The flask or beaker is placed on the gauze during heating. It helps to spread out the flame and heat evenly under the container.
Crucible and lid (made of Porcelain)		A crucible is a container in which substances can be heated to very high temperature. It is made of porcelain or a non - reactive metal.
Evaporating dish (Made of porcelain)		An evaporating dish is a shallow bowl with a curled lip. It is used to heat and evaporate liquids and solutions. It can be heated to very high temperatures.
Deflagrating spoon (Made of iron)		A deflagrating spoon is a long - handled spoon used to heat small amount of substances inside a gas jar.
Filter funnel (Made of plastic or glass)		A filter funnel is a device that is wide at the top and narrow at the bottom. It is used to separate solids from liquids.
Filter paper (made of tissue paper)		A filter paper is usually folded into a cone and placed in filter funnel to separate solids from liquids.
Mortar and pestle (Made of porcelain)		A mortar is a small hard bowl. A pestle is a small heavy tool used for crushing things. A mortar and pestle are usually used for crushing or grinding substances.
Test tube rack (Made of wood or plastic)		A test tube rack is specially designed for placing test tubes so that they do not roll or break.

Apparatus	Diagram	Explanation
Beaker (Made of glass or plastic)		A beaker is a glass or plastic container that is used for holding, heating and mixing liquids. It can not measure liquids accurately but can be used to estimate their volumes.
Test tube (Made of glass)		A test tube is used for holding chemicals or for heating substances for short periods of time.
Dropper (made of glass or plastic)		A dropper is used to add liquids during an experiment, drop by drop.
Conical flask (made of glass)		A conical flask is used for holding liquids in experiments that heating is not necessary.
Flat - bottomed flask (made of glass)		A flat bottomed flask is used for holding liquids in experiments where heating is not necessary.
Round bottomed flask (made of glass)		A round bottomed flask is used to hold liquids in experiments where heating is needed.
Distilling flask (made of glass)		A distilling flask is used for holding mixtures to be separated in distillation process.

Apparatus	Diagram	Explanation
Watch glass (Made of glass)		This is a very shallow circular glass container that is used as a surface to evaporate some liquids, to hold substances that are being weighed or observed, or as a cover for a beaker.
Gas jar (Made of glass)		A gas jar is a glass container that is used for collecting gases during experiments.
Condenser (made of glass)		A condenser is used to cool gases down during distillation so they condense and form a liquid. The condenser consists of two tubes one inside the other. The hot gases pass through the middle tube and the cold water passes through the other tube.
Trough (Made of glass or plastic)		This is a heavy walled glass or plastic basin used to contain water and support a beehive shelf during preparation of gases.
Dispensing bottle (Made of glass)		Dispensing bottle is used for keeping aqueous chemicals for dispensing.

Experiment 2.1

Aim : To measure the temperature of liquids

Materials: Thermometer, beakers, tripod stand, wire gauze, stopwatch, pair of tongs, water.

Procedure

1. Pour some tap water into two beakers. Measure the temperature of the water by dipping a thermometer in each of the beakers for one minute.
2. Remove the thermometer and record the temperature
3. Place one beaker in a fridge, or in a bucket of cool water or ice cubes. Let it stand there for about ten minutes
4. Remove the beaker from the fridge or bucket and dip a thermometer in the water for one minute. Record the temperature.
5. Place wire gauze on a tripod stand.
6. Place a Bunsen burner under the tripod stand and light it.
7. Place the second beaker of water on the wire gauze and heat for five minutes
8. Turn off the Bunsen burner. Use tongs to remove the beaker from the wire gauze to avoid burning yourself.
9. Place the thermometer in the beaker containing hot water. Let it stand for one minute then remove it . Note the temperature

Discussion question

When the thermometer is placed in tap water, water from the fridge, and heated water, what will be the respective readings?

Experiment 2.2

Aim: To measure the mass of different substances

Materials: Triple beam or electronic balance, various substances such as salt, sugar, watch glasses.

Procedure

1. Place an empty watch glass on the weighing balance. Note down its mass
2. Place the various items you have on the different watch glasses
3. Place the watch glass on the balance ,one at a time and note down the mass of each item

Discussion question

What are masses of different items you have measured?

2.5 FIRST AID

First aid is the help given to a sick or injured person before medical assistance from the hospital. Accidents happen always unexpectedly. Therefore the knowledge of first aid is important.

The following are importance of first aid:

1. First aid reduces the chance of infection, pain, bleeding and scarring.
2. First aid relieve suffering.
3. First aid preserve life.
4. First aid prevent the situation from worsening.
5. First aid promote recovery.
6. First aid prevents permanent disability.

CAUSES OF ACCIDENTS IN THE LABORATORY

The following are some of the causes of accidents in the laboratory:

1. Careless handling and storage of flammable and toxic substances may result in fire and explosions
2. Dealing with new or unfamiliar chemical substance may lead into existence of special hazard
3. Toxic gases, fumes or liquids may escape from their container or spill while being handled and cause poisoning ,allergies and respiratory problems.
4. Uncontrolled or unplanned chemical reactions can cause fire and dangerous explosions.
5. Wet, uneven or damaged floors can cause slips
6. Dropped or burst glassware can cause severe cuts.
7. Entanglement of clothes ,hair or fingers in rotating equipment such as centrifuges can cause bodily injury
8. Noise and vibration produced from equipment such as centrifuges and stirrers can cause hearing loss and stress
9. Lack of working understanding of hazards could lead to poisoning.
10. Improper or unintended use of equipment may cause damage to equipment or injure yourself
11. Inexperienced person working in the laboratory could lead to spillage of chemical and even fires.
12. Loss of attention to task may lead to use of wrong reagents.
13. Chemical spills and exposure could lead to burns and damage to body parts such as eyes.
14. Improper disposal of wastes may result in explosions, burns and fires
15. Poor ventilation in a laboratory may cause poisoning.

16. Electrocution could occur if electrical appliances are not plugged properly or are touched with wet hands

FIRST AID KIT

A first aid kit is a box that contains equipment and chemicals needed for the first aid. The main items in a first aid kit can be explained in the following table.

Table 2.2 Components of the first aid kit and their uses.

ITEM	USE
First aid book	Contains guidelines on how to use the items in the first aid kit.
Plaster or adhesive bandage	Covering small cuts or wounds.
Sterile gauze	Covering wounds to protect them from dirt and germs.
Iodine tincture	Cleaning fresh wounds to kill germs.
Soap	Washing hands, wounds and equipment
Pain killers	Relieving pain
Scissors or razor blade	Cutting dressing materials
Safety pins	Securing bandages
Bandage	Keeping dressings in place and immobilizing injured limbs
Cotton wool	Cleaning and drying wounds
Thermometer	Taking body temperature
Disposable sterile gloves	Covering the hands to avoid infecting wounds and to prevent direct contact with a victim's body fluids.
Petroleum jelly	Soothing chapped skin
Liniment	Reducing muscle pain
Torch	Source of light
Whistle	Blown to call for help
Mild antibiotics	Treating mild bacterial infections on the skin, ear, nose and mouth
Gentian violet (G.V)	For fungal infections of the skin and mouth. Also used for the treatment of serious heat burns.
Tweezers	Plucking hair from the face or eye brows
Instant ice packs	Treating injuries, sprains, bumps, bruises, insect bites and stings. Can also help ease the symptoms of headaches.
Emergency blanket	Preventing and countering hypothermia by reducing heat loss from person's body. Covering injured victims in order to help reduce shock.

2.6 FIRST AID PROCEDURES

The following are some situations that may require first aid and the procedure to follow in giving help.

BURNS

A burn is a type of injury to flesh or skin caused by heat, electricity, chemicals, light, radiation or friction. Burns usually cause blisters on the skin and if severe the skin becomes charred and peels off. Burns caused by hot liquids or gas are called ***scalds***

Procedure

1. Lay the victim down and protect the burned area from coming into contact with the ground if possible
2. Hold the burned area under cool running water for 10 or 15 minutes or until the pain subsides. If this is impractical, immerse the burn in cool water or cool it with cold compresses. Cooling the burn reduce swelling by conducting heat away from the skin. Don't put ice on the burn.
3. Check breathing and pulse and be prepared to resuscitate the victim ,if necessary
4. Gently remove any jewelry, shoes or burned clothing from the injured area. Loosen any tight clothing. Do not remove any clothing that is sticking to the skin.
5. Cover the burn with sterile gauze bandage. Don't use fluffy cotton or other material that may get lint in the wound. Wrap the gauze loosely to avoid putting pressure on burned skin. Bandaging keeps air off the burn, reduces pain and protects blistered skin.
6. Take an over-the-counter pain reliever. These include aspirin, naproxen or acetaminophen. Use caution when giving aspirin to children or teenagers. Though aspirin is approved for use in children older than 2,children and teenagers recovering from chickenpox or flu-like symptoms should never take aspirin. Seek medical help immediately.

Caution

- Do not use ice. Putting ice directly on a burn can cause a person's body to become too cold and cause further damage to the wound.
- Do not apply egg whites, butter or ointments to the burn. This could cause infection.
- Do not break blisters. Broken blisters are more vulnerable to infection
- Burns to the face and in the mouth or throat are serious, as they cause rapid inflammation of the air passage and may cause suffocation. Seek medical help immediately.

SUFFOCATION

Suffocation is the condition of being deprived of oxygen. Foam can appear at the mouth and nostril of the victim. The victim may eventually lose consciousness. Some chemical substances can cause suffocation when inhaled, leading to breathing difficulties

Procedure

1. Remove anything covering the mouth and nose. A plastic bag should be torn open and attempt to remove anything obstructing the airway.
2. Ensure the victim's airway is open for air to reach the lungs. Do this by placing the victim on his or her back. With one hand on the victim's fore head and the other on the chin, tilt the head backwards to open the air way
3. If the casualty is unconscious,place him or her in the lateral position, check the airway, breathing and pulse and begin expired air resuscitation (EAR) or cardiopulmonary resuscitation (CPR)if necessary. Continue until natural breathing is restored.
4. If the casualty is or becomes conscious ,monitor air way, breathing and pulse.
5. Keep the victim warm using a light blanket.
6. Seek medical help immediately

CHOKING

Choking is the mechanical obstruction of the flow of air from the environment into the lungs. Choking prevent breathing, and can be partial or complete. Signs of choking include difficulty in speaking and breathing.

Procedure

1. Encourage the victim to cough up the object.
2. If the object remain stuck deliver five back blows between the person's shoulder blades with the heel of your hand
3. If the object is still stuck ,perform the ***Heimlich manouvre***. This procedure involves the following
   a. Stand behind the person. Wrap your arms around the waist. Tip the person forward slightly
   b. Make a fist with one hand. Position it slightly above the person's navel.
   c. Grasp the fist with the other hand. Press hard into the abdomen with a quick upward thrust-as if trying to lift the person up .
   d. Perform a total of 5 abdominal thrusts, if needed. If the blockage still isn't dislodged, repeat the thrusts until the object comes out.

BRUISES

A bruise is a mark on the skin formed when small blood vessels break and leak their content into the soft tissue beneath the skin. When a blow breaks blood vessels near the skin surface leakage of small amount of blood lead to discoloration of the skin.

Procedure

1. Elevate the injured area
2. Apply ice or a cold pack several times a day for a day or two after the injury
3. Rest the bruised area, if possible
4. Consider pain relief drugs and others which reduce swelling

Caution

For the first 48 hours after an injury, avoid things that might increase swelling such as hot showers, hot tubs, hot packs or alcoholic beverages

SHOCK

Shock is a life –threatening condition that occurs when the body is not getting enough blood flow. This can damage multiple organs. Shock requires immediate medical treatment and can get worse very rapidly. Shock ultimately leads to cellular death, progressing to organ failure, and finally if untreated whole body failure and death. In general when a person is in shock, his or her vital organs like heart , the lungs and the brain are not getting enough blood or oxygen.

A person experiencing shock may have pale or gray skin ,the pulse is weak and rapid, the person may be nauseated and vomit, the eyes lack luster and may seem to stare (sometimes the pupils are dilated) the person may be unconscious. If conscious, the person may feel faint or be very weak or confused. Shock sometimes causes a person to become overly excited and anxious

Procedure

1. Have the person lie down on his or her back with feet about a foot higher than the head. If raising the legs will cause pain or further injury, keep him or her flat. Keep the person still.
2. Check for signs of circulation (breathing, coughing or movement) and if absent; begin cardiopulmonary resuscitation (CPR).
3. Keep the person warm and comfortable by loosening any belts or tight clothing and covering the person with a blanket. Even if the person complains of thirst, give nothing by mouth.
4. Seek medical help immediately.

ELECTRIC SHOCK

An electric shock is a condition that occurs when a person comes into contact with an electric energy source. An electric shock is usually caused by contact with poorly insulated wires or ungrounded electrical equipment, by using electrical equipment while in contact with water, or by being struck by lightning.

Procedure.

1. Turn off the source of electricity.
2. Move the source of electric shock away from the person using a dry, non conducting object made of cardboard, plastic or wood.
3. Check for signs of circulation (breathing, coughing or movement). If absent, begin cardiopulmonary resuscitation (CPR) immediately.
4. If the person is breathing but unconscious, put him or her in the recovery position.
5. Administer first aid for burns, shock or other injuries the victim may have sustained.
6. Seek medical help immediately.

Caution.

- Do not touch the person with bare hands if he or she is still in contact with the electrical current.
- Do not get near high – voltage wires until the power is turned off. Stay at least 20 feet away – farther if wires are jumping and sparking.
- Do not move a person with an electrical injury unless the person is in immediate danger.

FAINTING.

Fainting is a sudden and temporary loss of consciousness resulting from a reduction in blood flow to the brain. The episode lasts less than a couple of minutes and a victim recovers from it quickly and completely. A victim may feel light headed or dizzy before fainting. A longer, deeper state of unconsciousness is often called a coma.

Procedure.

1. Check the person's airway and breathing. If necessary begin rescue breathing through cardiopulmonary resuscitation (CPR)
2. Loosen tight clothing around the neck.
3. Raise the person's feet above the level of the heart (about 12 inches).
4. If the person has vomited, turn onto his or her side to prevent choking.
5. Keep the person lying down for at least 10 – 15 minutes, preferably in a cool and quiet space. If this is not possible, sit the person forward with the head between the knees.

6. Seek immediate medical help if the victim does not recover in a few minutes.

BLEEDING.

Bleeding is the loss of blood from the circulatory system. Bleeding can occur internally, where blood leaks from blood vessels inside the body or externally, either through a natural opening such as the mouth, nose, ear, urethra, vagina, anus or through a break in the skin. Bleeding may be light or severe. Excessive loss of blood can cause death.

Light bleeding.

1. Place the victim in a comfortable position.
2. Elevate the injured part
3. Gently clean the wound using clean water and antiseptic or common salt. Cover the wound using sterile gauze. Gently clean the surrounding skin and dry it using sterile dressing.
4. Dress the wound and bandage it.
5. Take the person to hospital in case the bleeding continues.

Severe bleeding.

Procedure.

1. Have the injured person lie down and cover the person to prevent loss of body heat. If possible, position the person's head slightly lower than the trunk or elevate the legs and elevate the site of bleeding.
2. While wearing gloves, remove any obvious dirt or debris from the wound. Do not remove any large or more deeply embedded objects.
3. Apply pressure directly on the wound until the bleeding stops. Use a sterile bandage or clean cloth and hold continuous pressure for atleast 20 minutes. Maintain pressure by binding the wound tightly with a bandage or clean cloth and adhesive tape.
4. Don't remove the gauze or bandage. If the bleeding continues and seeps through the gauze or other material you are holding on the wound, don't remove it, instead add more absorbent material on top of it.
5. Squeeze a main artery if necessary. If the bleeding does not stop with direct pressure, apply pressure to the artery delivering blood to the area.
6. Immobilize the injured body part once the bleeding has stopped. Leave the bandages in place and get the injured person to the emergency room as soon as possible.
7. Seek medical help immediately.

Nose bleeding.

Procedure.

1. Let the victim sit upright and lean forward. Remaining upright reduce blood pressure in the veins of the nose. Sitting forward helps to avoid swallowing blood. Swallowing blood can irritate the stomach.
2. Let the victim pinch his/her nose by using the thumb and index finger. Let him/her breath through the mouth and continue to pinch for five to ten minutes. Pinching sends pressure to the bleeding point on the nasal septum and often stops the flow of blood.
3. Place a wet piece of cloth at the back of the victim's neck.
4. Let the victim prevent re-bleeding by avoiding picking or blowing the nose. Advice the victim not to bend down for several hours after the bleeding episode.
5. If re-bleeding occurs, take the victim to hospital.

POISONING

A poison is a substance which, if taken into the body in sufficient quantity, may cause temporary or permanent damage. Poisons can be swallowed, absorbed through the skin, inhaled, splashed into the eyes or injected. Once in the body, they may enter blood stream and can be carried swiftly to all organs and tissues.

Signs and symptoms of poisoning include burns or redness around the mouth and lips from drinking certain poisons, breath that smell like chemicals, vomiting, difficulty breathing, sleepiness and confusion.

Procedures

1. Call for medical assistance immediately
2. If the person has been exposed to poisonous fumes such as carbon monoxide, get him or her into fresh air immediately.
3. If the person swallowed the poison, remove anything remaining in the mouth.
4. If the suspected poison is a household cleaner or other chemicals, read the label and follow instructions for accidental poisoning.
5. If the poison spilled on the person's clothing, skin or eyes, remove the clothing. Flush the skin or eyes with cool or lukewarm water.
6. Make sure the person is breathing. If not, start cardio pulmonary resuscitation (CPR) and rescue breathing.
7. Take the poison container (or any pill bottle) with you to the hospital.

VOMITING.

Vomiting is the forceful expulsion of the contents of the stomach through the mouth sometimes the nose. Vomiting may result from many causes, ranging from gastritis or poisoning to brain tumors, or elevated intracranial pressure. The feeling that one is about to vomit is called nausea, which usually precedes, but does not always lead to vomiting.

Procedure.

1. Reassure the victim and advise him or her to take slow and deep breaths.
2. Use a damp cloth to wipe the face of the victim
3. Let the victim drink gradually large amounts of clear liquids, including oral rehydration drink.
4. Do not give the victim solid food until vomiting has stopped.
5. Provide the victim with bananas, rice, apple sauce without sugar, toast, potatoes and pasta after 24 hours.
6. Get medical assistance if
   a. Vomiting goes on for longer than one day.
   b. There is blood in the vomit.
   c. The vomiting is occurring because of a known injury, like head trauma or infection.
   d. The victim acts confused, lazy or lethargic and is less alert than usual.
   e. Vomiting and diarrhea are present.
   f. Severe abdominal pain is present.
   g. Severe headache or stiff neck is present.
   h. Feelings of nausea last for longer than one week.

FALLING

Falling is an act of moving from a higher level to a lower level rapidly and without control. Falling may result after loss of balance and collapsing.

Procedure.

1. Cover the victim's body with a blanket to reduce the chances of going into shock.
2. Look for signs of a fractured skull such as unequal pupils in their eyes, bleeding from the ear or clear fluid running from the nose.
3. Check for swelling or limbs sticking out at an unusual angle. If a bone is broken, keep it still and wrap it with towels to support it.
4. If you cannot see any broken bones or you do not think the victim have got a head or neck injury, put the victim in recovery position.
5. If the victim is conscious and there are no obvious signs of serious injury apply

a cold compress made from a cloth dipped in water to the areas that were hit for 10 minutes to help with swelling.

6. If the victim is unconscious seek medical help immediately.

SUMMARY

(a). A laboratory is a special room or building that is designed and used for scientific experiments.
(b). Apparatus are special tools and equipment that are used in the laboratory.
(c). Laboratory rules are specified guidelines required to be followed when working with the laboratory.
(d). Chemical warning signs are safety symbols found on chemical containers, especially those used in the laboratory.
(e). First aid is the help given to a sick or injured person before medical assistance from the hospital.
(f). First aid kit is a box that contains equipments and chemicals needed for the first aid.

REVIEW QUESTIONS.

1. Choose the best answer from the choices given.

i. Temperature can be measured with
   (A) Vernier caliper (B) Measuring cylinder.
   (C) Measuring syringe. (D) Thermometer.

ii. ………… is good for measuring volumes quickly but not very accurately.
   (A) Measuring cylinder. (B) Beaker
   (C) Volumetrick flask. (D) Pipette.

iii.………………….. Is used to make up a solution to a total final volume.
   (A) Burette. (B) Gas syringe.
   (C) Volumetric flask. (D) Conical flask.

iv. …………….. Can be used to measure the volume of a gas produced during an experiment.
   (A) Gas syringe. (B) Round bottomed flask
   (C) Flat bottomed flask (D) Volumetric flask.

v. A condenser is used to cool gases down during …… so they will condense and form a liquid.
   (A) Evaporation (B) Decantation
   (C) Filtration (D) Distillation.

2. Indicate ***TRUE*** or ***FALSE*** for the following statements.
   i. A mortar and pestle is used to grind lumps into a powder………………..
   ii. The funnel is used when transferring a liquid from one container to another………….
   iii. A pipe clay triangle fits on the top of a tripod and supports a crucible when it is heated……………
   iv. A combustion spoon is used to burn a large amount of a solid, perhaps sulphur in a gas jar of oxygen………….
   v. Students are advised to pick up reagent bottles by holding the stopper………………..
3. Match the items in List A with their corresponding statements in LIST B.

List A	List B
i. Harmful	A. This substance is dangerous and can cause death.
ii. Irritant	B. This substance explodes easily
iii. Radioactive	C. This substance catches fire easily
iv. Corrosive	D. This substance reacts easily with oxygen.
v. Toxic	E. Corrodes surfaces as well as the human body.
vi. Oxidant	F. Emits harmful radiations.
vii. Flammable	G. Irritate parts of the body.
viii. Explosive	H. This substance is harmful or poisonous
	I. Emits harmful radiations and reacts easily with oxygen.
	J. Catches fire and explodes easily.

4. Define first aid and first aid kit respectively.
5. Some beakers, syringes, measuring cylinders and troughs are made of plastic. What are the advantages of plastic containers.
6. Chemical apparatus can be divided into that which holds chemicals (containers) and that used to support other pieces of apparatus (supporters) copy and complete the following table putting in the names of the following pieces of apparatus. Flask, tripod stand, test tube rack, test tube holder, beaker, trough.

Containers	Supporters

7. The following list contains the names of pieces of apparatus with the letters in each name jumbled up. Rearrange the letters to produce the names of ten pieces of apparatus.

A. Lupstaa
B. Lefnum
C. Tptipee
D. Romtmreehet
E. Bulceric
F. Faskl
G. Tertbeu
H. Rebake
I. Gotsn
J. Gothur

From the list you have obtained.

i. Name seven pieces of apparatus made of glass.
ii. Name four pieces of apparatus which can be used as containers.
iii. Name two pieces of apparatus which can be used to measure volumes of liquids.
iv. Name one instrument that measures temperature.

8. Copy and complete the following table.

Apparatus	Made of	Uses
Gas jar		
Measuring cylinder		
Funnel		
Tongs		
Forceps		
Watch glass		
Evaporating basin		
Burette		
Test tube rack		
Spatula		

Chapter 3

Heat Sources and Flames

3.1 HEAT SOURCES

Heat is energy transferred from one body to another by thermal interactions. It is very useful in our day to day life. It is important to have good sources of heat in the laboratory. Chemical reactions which would proceed slowly at normal temperatures will take place faster when they are heated.

Heat may be obtained from electricity or from natural fuels like coal, charcoal, gas, kerosene, spirit etc. For each type of fuel used, a different kind of burner is needed. The common types of burners used in laboratories are: Bunsen burner, spirit burner, kerosene burner, gas stove, charcoal burner and candle. The Bunsen burner is the best of them all because it is convenient to handle and it produces a range of flames including a very hot flame (approximately 1000˚C).

A spirit burner uses methylated spirit as a fuel. Candle may be used where the experiment does not need heating to a high temperature. Burning candle produces a lot of soot and smoke. A kerosene burner can reach high temperatures (1000˚C or more). A gas stove uses natural gas as a source of fuel. If the Bunsen burner and kerosene burner are not available, a charcoal burner is the next choice simply because a charcoal burner can burn for a long time. The burner can produce high temperatures. A tin lamp is a small kerosene burner which can be used if soot is not a major handicap. This burner can supply a high temperature with use of blow pipe.

A sand bath is common piece of laboratory equipment made from a container filled with heated sand. It is used to provide even heating for another container, most often during a chemical reaction.

A water bath can be used to keep a reaction on vessel at the temperature of boiling water until all water is evaporated. A thermometer is usually kept in the water to monitor the temperature.

Spirit lamp

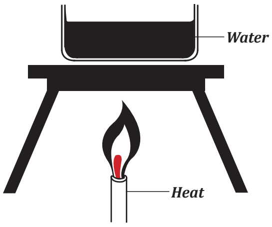
Bunsen burner

Kerosene burner

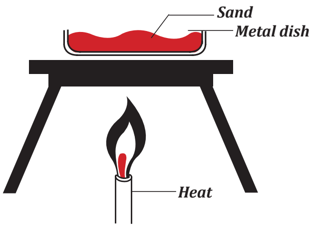
Charcoal burner

Gas stove

Tin lamp



Water bath



Sand bath

Fig. 3.1. Heat sources in the laboratory

3.2 THE BUNSEN BURNER

A Bunsen burner is a small laboratory heat source consisting of a vertical metal tube connected to a gas source and producing a very hot flame from a mixture of gas and air let in through adjustable holes. These holes allow air to enter the tube and mix with the gas in order to make a very hot flame.

The gas from the source can be natural gas (which is mainly methane) or a liquefied petroleum gas such as propane, butane or a mixture of both. The Bunsen burner is named after a German chemist and physicist, Robert Wilhelm Bunsen, who invented it in 1855. ***Consider the diagram of a Bunsen burner below;***

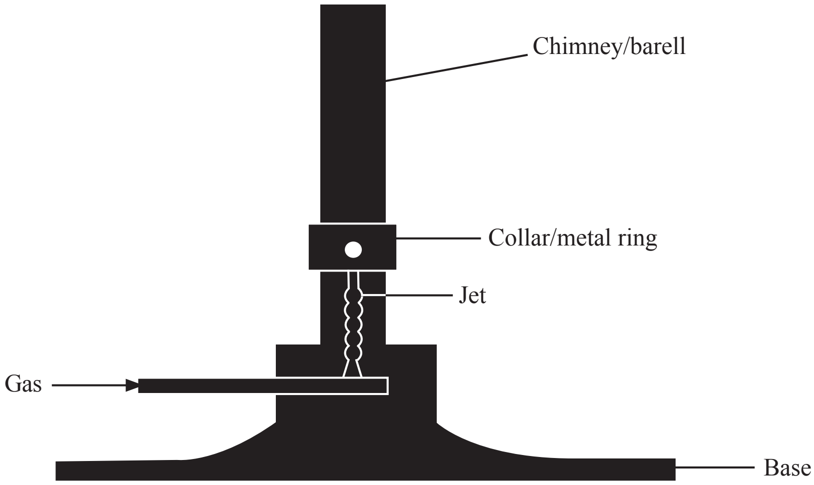


Fig: 3.2 Bunsen burner.

HOW THE BUNSEN BURNER WORKS

The gas enters the burner through a tube connected to a jet inside the base. Air enters the burner through the air holes at the base. The amount of air coming in can be varied by turning the collar. At the top of the barrel, the mixture of air and gas burns.

Activity 3.1; to light and adjust a Bunsen burner.

1. Dismantle the Bunsen burner. Start by unscrewing the barell or chimney. Next remove the collar or metal ring. Use diagram to identify the parts.
2. Reassemble the Bunsen burner.
3. Carry out the following instructions for lighting a Bunsen burner.
   a. Connect the Bunsen burner by the rubber tube to the gas mains.
   b. Close the air hole.
   c. Allow plenty of gas to enter the burner by turning the gas tap on.

d. Immediately bring a flame to the top of the barrel. You may use a lighted taper or a match stick. You will get a tall shapeless yellow flame. This is called the luminous or yellow flame.
e. Adjust the gas supply until the flame is about 10cm high.
f. Turn the collar until the air hole is fully open. You now have the non-luminous or blue flame. ***For safety reasons a lit Bunsen burner should always be left on a luminous flame.***

3.3 TYPES OF FLAME

A ***flame*** is a zone of burning gases that produces heat and light. It is visible gaseous part of a fire. The flame is formed as a result of burning a fuel. The colour and temperature of flame depend on the type of fuel burning and the source of the flame.

Flames can be classified as ***luminous*** or ***non-luminous***. A luminous flame is yellow. Although it uses oxygen when burning, the oxygen supply is usually not enough to completely burn up the fuel. Therefore it produces a black substance known as ***soot***.

The non-luminous flame is blue. This is because there is an adequate supply of oxygen and so the fuel burns efficiently. Such a flame doesn't produce any soot. It also produces more heat than the luminous flame.

Different heat sources produce different types of flames. Example, the candle and the tin lamp produce luminous flames, where as the gas stove produces a non-luminous flame.

On lighting the Bunsen burner, the type of flame obtained depends on how much oxygen is available for burning. When the air holes are closed, there is less oxygen and the flame is luminous. When the air holes are partly open, there is more oxygen and the flame is medium. When the air holes are fully open, there is enough oxygen and the flame is non-luminous.

When the air holes of the Bunsen burner are closed, the flame is wavy and yellow. This luminous flame is called a safety flame because it is easily seen and so is less likely to cause accidents. When the air holes of a Bunsen burner are half open, the flame becomes difficult to see. This type of flame can easily cause accidents in the laboratory since a person may not be aware that the burner is on. When the air holes of a Bunsen burner are fully open, the flame is cone shaped and burns with a roaring sound.

Consider below diagrams for luminous and non-luminous flames.

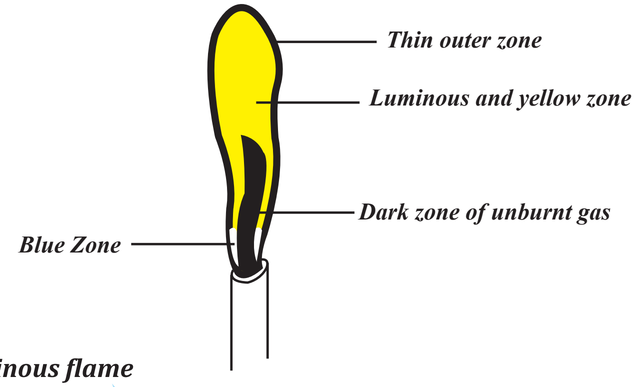


Fig. 3.3 Luminous flame

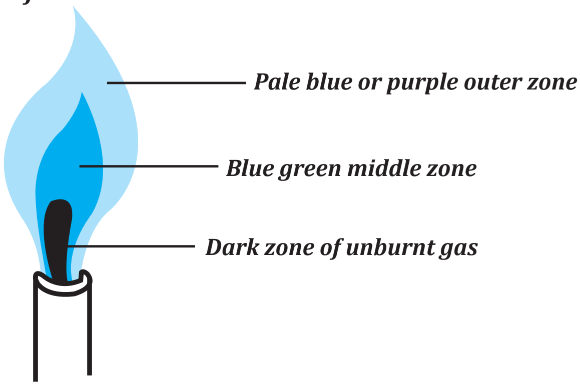


Fig. 3.4 Non-luminous flame

3.4 CHARACTERISTICS OF LUMINOUS AND NON-LUMINOUS FLAMES

Table 3.1; characteristics of luminous and non-luminous flames.

Luminous flames	Non - luminous flames
i. Yellow in colour.	i. Blue in colour.
ii. Produces soot.	ii. Does not produce soot.
iii. Has a wavy flame.	iii. Have a triangular flame.
iv. Produces less heat.	iv. Produces more heat.
v. Burns quietly.	v. Burns with a roaring noise.
vi. Produced when air holes of a Bunsen burner are closed.	vi. Produced when air holes of a Bunsen burner are open.
vii. Has four zones.	vii. Has three zones.

USES OF LUMINOUS AND NON - LUMINOUS FLAME

Luminous Flame.

The luminous flame is mainly used for lighting. It is yellow flame that easily brightens a room. Since it is not very hot, it is safer for lighting than the non-luminous flame. Some heat sources that can produce a luminous flame are candle, tin lamp and hurricane lamp.

Non-Luminous Flame.

The following are uses of non-luminous flame.

1. Non-luminous flame is used for heating purposes because it gives a lot of heat. Most heating in the laboratory is done using a Bunsen burner. Non-luminous flame of a Bunsen burner has three parts; the pale blue or purple outer zone, the blue green middle zone and the inner dark zone of unburnt gas.
2. The hottest part of the non-luminous flame is at the tip of the middle zone.
3. A non-luminous flame is used in the flame test of certain chemical substances. A flame test involves introducing a sample of the desired substance to the non-luminous flame and observing the colour change. The colour can indicate what components the substance is made of.
4. Welding is the joining together of metal pieces or parts by heating their surfaces and pressing them together. A non-luminous flame is suitable for this task because it is very hot.
5. A non-luminous flame is suitable for cooking since it gives off enough heat and it does not produce soot.

Activity 3.2

To compare the heating ability of luminous and non-luminous Bunsen flames.

1. Measure 100 cm^3 of water in a measuring cylinder and pour this into a 250 cm^3 beaker.
2. Put a thermometer into the water and take the temperature.
3. Stand the beaker on a tripod and gauze.
4. Light a Bunsen burner and adjust until you have a blue flame 10cm high. Do not touch the gas tap now throughout the experiment.
5. Put the lit Bunsen burner under the tripod and heat the water.
6. Stir the water constantly with a glass rod and take the temperature every half minute for 5 minutes.
7. Empty the water away and allow the apparatus to cool.
8. Repeat the experiment using a yellow flame.
9. Draw two graphs on the same grid showing the temperature of the water during the experiment.

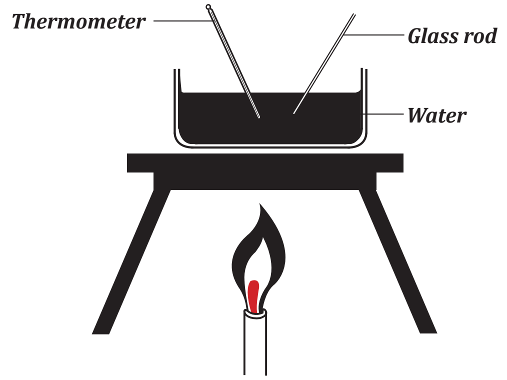


Figure 3.5; Apparatus used to demonstrate the heating ability of luminous and non-luminous flames.

SUMMARY

(a). Heat is energy transferred from one body to another by thermal interactions.
(b). Common types of burners used in laboratories are Bunsen burner, spirit burners, kerosene burners, gas stove, charcoal burner and candle.
(c). The Bunsen burner is the best source of heat in the laboratory.
(d). A Bunsen burner is a small laboratory heat source consisting of a vertical metal tube connected to a gas source and producing a very hot flame from the mixture of gas and air let in through adjustable holes.
(e). The Bunsen burner is named after a German chemist and physicist, Robert Wilhelm Bunsen.
(f). Flames can be classified as luminous and non-luminous flames.
(g). Luminous flame has four zones whereas non-luminous flame has three zones.

REVIEW QUESTIONS

1. Match the items in LIST A with their corresponding statements in LIST B.

List A	List B
i. Spirit lamp	A. Keep a reaction vessel at the temperature of boiling water.
ii. Kerosene burner	B. Container filled with heated sand.
iii. Charcoal burner	C. Small kerosene burner which produces lot of soot.
iv. Gas stove	D. Can burn for a long time when lit and produce high temperature.
v. Tin lamp	E. Uses natural gas as a source of fuel.
vi. Water bath	F. Does not heat to a high temperature of 1000˚C or more.
vii. Sand bath	G. Does not heat to a high temperature.
viii. Candle	H. Uses methylated spirit as fuel.
	I. The best source heat in the laboratory.

2. Write ***TRUE*** or ***FALSE*** for the following statements
   i. Heat is very useful to our day to day life,.................
   ii. Heat increases the speed of chemical reactions.,.................
   iii. The Bunsen burner is named after a German chemist and physicist Robert Wilhelm Bunsen.,.................
   iv. Methane is the good example of natural gas.,.................
   v. Air enters the Bunsen burner through air holes.,.................
   vi. For safety reasons a lit Bunsen burner should always be left on a luminous flame.,.................
3. Choose the most correct answer for the following questions.
   i. .... is a zone of burning gases that produces heat and light.
      A. Fuel B. Heat C. Flame D. Combustion.
   ii. The colour and temperature of flame depend on:
      A. Type of fuel burning and the source of flame
      B. Type of fuel burning and heat involved
      C. Bunsen burner capacity.
      D. Quantity of heat.
   iii. A black substance resulted after incomplete combustion.
      A. Soot B. Rust C. Smoke D. Flame
   iv. When there is an adequate supply of oxygen:
      A. The fuel burns slowly
      B. The fuel burns smoothly
      C. The fuel burns efficiently.
      D. The fuel produces soot.

v. ....and .... produces luminous flames

A. Kerosene burner and blue flame of burner

B. Candle and gas stove

C. Candle and tin lamp

D. Candle and Bunsen burner

4. Outline the differences between luminous and non-luminous flames
5. Draw well labeled diagrams of luminous and non-luminous flames respectively.
6. Briefly answer the following questions
   a. Name six sources of heat that can be used in the laboratory.
   b. Name two sources of heat that can produce a luminous flame
   c. Describe the process of lighting of Bunsen burner.
   d. The non-luminous flame is used in different areas. Give four such areas and explain why the flame is preferred for those uses
   e. Give two reasons why a luminous flame not a non-luminous flame, is preferred for lighting.

Chapter 4

Scientific Procedure

The scientific procedure is the organized set of guidelines used for investigating problems, acquiring new knowledge or correcting and integrating previous knowledge. The main characteristic which distinguishes the scientific method from other methods of acquiring knowledge is that scientists seek to let reality speak for itself, support a theory when a theory's predictions are confirmed and challenging a theory when its predictions prove false.

4.1. SIGNIFICANCE OF THE SCIENTIFIC PROCEDURE

We carry out experiments to find answers to scientific questions. The findings from the experiments help us to solve scientific problems. The following are some of significance resulted from application of scientific procedure;

- The scientific procedure allows for one to prove, or disprove a theory or hypothesis. It allows for results from an experiment to be repeated or validated.
- The scientific procedure is a rational method for examining the universe around us, making observations about it, and hypothesizing rules or models for how it operates.
- The scientific procedure attempts to minimize the influence of bias on the side of the experimenter.
- The scientific procedure provides an objective, standardized approach to conducting experiments and improve the results.
- The scientific procedure is used by scientists to answer questions and solve complex problems.
- The scientific procedure help scientists to detect changes in our environment and suggest ways in which the effects may be avoided.

4.2 MAIN STEPS OF THE SCIENTIFIC PROCEDURE.

The scientific method is a logical way to solve a problem. It is a process that is used to find answers to questions about the world around us. The scientific method is used by researchers to support or disprove a theory. Below are steps of the scientific procedure.

1. Identification of the problem.
2. Formulation of Hypothesis.
3. Experimentation and Observation.
4. Data collection and analysis.
5. Data interpretation.
6. Drawing conclusion.

1. IDENTIFICATION OF THE PROBLEM

A problem statement is a question that compares variables. Careful observation leads to this question. This is when scientists raise questions based on what they have observed. When writing the problem statement think what you want to know or explain. Use observation you have made to write a question about the problem or topic you want to investigate. An example of a problem could be ***"does eating of greasy food cause pimples?"***

2. FORMULATION OF HYPOTHESIS

A hypothesis statement is a statement that expresses what the expected answer to the experiment will be. This is what you think the results of the experiment will show. A hypothesis is an intelligent guess which predict the answer to the problem at hand. An example of a hypothesis could be ***"eating greasy food cause pimples"***

3. EXPERIMENTATION AND OBSERVATION

An experiment is a planned way to test a hypothesis and find out the answer to the problem statement. To conduct an experiment a scientist develops and follows a procedure ***(steps to the experiment)***. The procedure also includes a detailed materials list.

In the experiment there are things that might change: These are called ***variables***.

There are three types of variables. These are:

- ***Dependent variable*** - This is the factor in the experiment that changes its value when the values of the other variables change. It is the factor being measured.
- ***Independent variable*** - This is the factor that is manipulated so as to obtain different values for comparison.
- ***Controlled variable*** - This is the factor in the experiment that does not change, or is kept constant. It does not affect the outcome of the experiment.

In the experiment on investigating whether eating greasy food causes pimples or not:

- Pimples are the factor under investigation. It is the dependent variable.
- Eating greasy food is the factor that is varied. It is the independent variable.
- The number of people involved as experimental sample is kept constant. It is the controlled variable.

It is often important to carry out an experiment where all the independent variables are kept constant. This is called a ***control experiment.***

A control experiment is an observation designed to minimize the effects of variables other than the single independent variable.

Example: In the experiment on whether eating greasy food causes pimples or not, there can be two equal groups of people where by one group will be provided with greasy food while the other group won't be provided with greasy food but rather food which is greasy free. The group provided with greasy free food is the control experiment.

4. DATA COLLECTION AND ANALYSIS

Data is the information you get when you test the variable. Confirm the results by retesting and modify the procedure if needed. Did you get the same results each time you retested? If not, why? Is there a change you can make that would make results more consistent?. Include tables, graphs and photographs that show your data.

Consider the table below;

Table. 4.1; number of people with pimples at different food amounts

Amount (grams) of greasy food	Number of people with pimples
50	
150	
200	
250	
300	
350	
400	
450	
500	

5. DATA INTERPRETATION

After recording and analyzing the data, you look for trends or patterns and explain why they occur that way. This is what is known as interpretation of data. These trends will help you to make your conclusion.

Example: in the above experiment, it may be noticed that as you increase the amount of greasy food the number of people with pimples increases. This is because increasing the amount of greasy food increases the effect of having pimples.

6. DRAWING A CONCLUSION

A conclusion statement is a statement that presents the findings of an experiment, what the data shows, and states if the hypothesis was correct (supported) or incorrect (negated). Conclusion statements also make recommendations for further study and possible improvements to the procedure.

Example, in the experiment on whether the greasy food causes pimples or not, you may notice that increasing the amount of greasy food increases the number of people with pimples.

From this experiment you may conclude that ***"eating greasy food cause pimples"*** This supports and approves the formulated hypothesis.

REPORTING RESULTS

Scientists usually communicate by publishing their findings in scientific journals and other publications. They can also present their results to the scientific community at seminars, symposia and meetings.

Consider the below figure that shows a summary of scientific procedure.

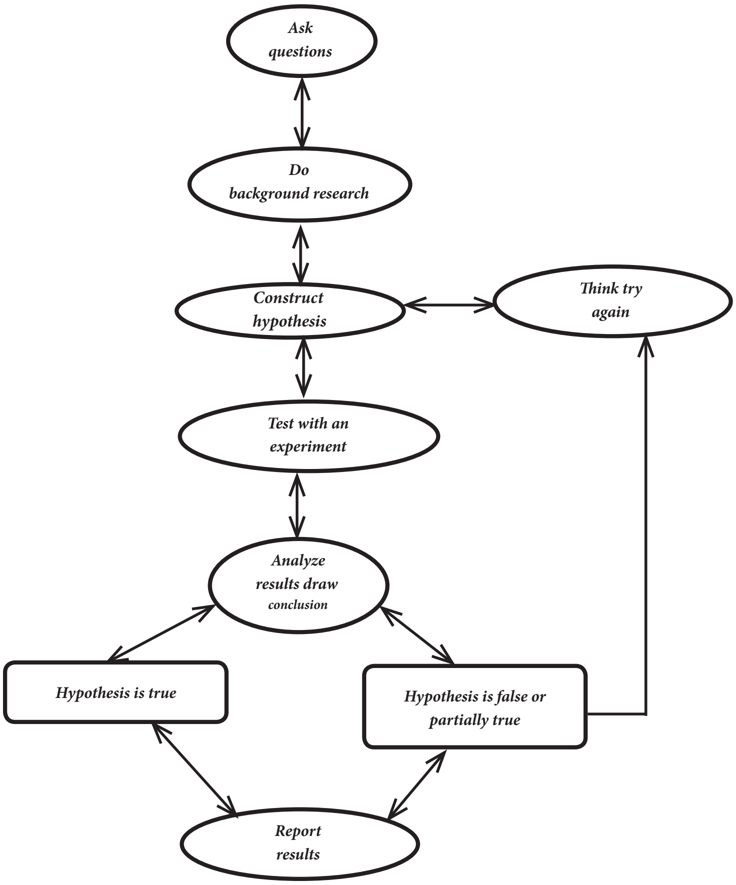


Figure 4.1; steps in scientific procedure.

4.3 APPLICATION OF THE SCIENTIFIC PROCEDURE

The scientific procedure have proven to be so fruitful and it is applied in many fields. The following are the common areas where it is applied.

1. In hospitals and health centres.

Diagnosis and treatment of diseases mainly depend on scientific procedure. Systematic steps undertaken by a medical practitioner leads to identification and subsequent treatment of illness in general.

2. In project work.

Careful and persistent studying of a problem over a period of time leads to planning of a project work thereby solutions to a problem can be found.

3. In field work

Field work is an investigation or search for material or data, made in the field as opposed to the classroom, laboratory or official headquarters. In field work, finding answers to problems and testing of hypothesis are important elements.

4. In research centres

A research centre is an establishment endowed for doing research. Research centres may specialize in basic research or may be oriented to applied research. Scientific research is the term mainly used to activities undertaken in research centres. Scientific research involves scientific methods.

5. In school laboratory

A laboratory is a special room or building designed and applied for scientific experiments. Scientific experiments relies heavily on scientific procedures.

SUMMARY

(a). The scientific procedure is the organized set of guidelines used for investigating problems, acquiring new knowledge or correcting and integrating previous knowledge.

(b). We carry out experiments to find answers to scientific questions.

(c). Main steps of the scientific procedure are: identification of the problem, formulation of hypothesis, experimentation and observation, data collection and analysis, data interpretation and drawing conclusion.

(d). The three types of variables are dependent variable, independent variable and controlled variable.

(e). The scientific procedure has proven to be so fruitful and it is applied in many fields.

REVIEW QUESTIONS

1. Choose the most correct answer for the following questions;

vii. .... is the organized set of guidelines used to answer questions or solve problems.

A. Scientific experiment B. Scientific research
C. Scientific procedure D. Hypothesis

viii. The second step of the scientific procedure is ....

A. Data interpretation. B. Drawing conclusion
C. Indentificaton of the problem. D. Formulation of hypothesis.

ix. ....is a suggested answer to a question or problem.

A. Hypothesis B. Diagnosis C. Prognosis D. Inquiry

x. A test done to find out if the formulated hypothesis is true or false is called

A. Research B. Experiment C. Observation D. Investigation.

xi. ....are factors that affect the problem being investigated.

A. Variance B. Variables C. Objects D. Data.

2. In an experiment to find out what causes rusting of iron you might carry out a control experiment. How would you do this and why would you do it?

3. What are the advantages of following scientific procedure over trial and error methods?

4. Write ***TRUE*** or ***FALSE*** for the following statements.

i. If the results you obtain from an experiment do not support your hypothesis, give ideas for further testing to find a solution.
ii. The sixth step in the scientific procedure is drawing conclusion.
iii. Factors in an experiment that can be manipulated to get desired results are called independent variables.
iv. A theory can be derived without any experimental results.
v. A law can only be claimed when all scientists agree with the experiment results and procedures.

Chapter 5

Matter

5.1 CONCEPT OF MATTER

Matter is anything that occupies space and has mass. Mass is a measure of a quantity of matter in an object. Mass is usually measured in kilograms or grams. Matter is also a general term for the substance of which all observable physical objects consists.

5.2 STATES OF MATTER

Properties describe matter. A block of wood, milk and air all have properties. All the material on earth exist in three states; solid, liquid and gas. The state of matter refers to the group of matter with the same properties.

SOLID

A solid has a definite size and shape. A solid is usually hard and not easily deformed. Examples of solids are wood, computer, desk and floor.

LIQUID

Liquids do not have a definite shape. Liquids have definite volume and take the shape of its container. Liquids can flow, be poured and spilled. Examples of liquids are milk, water and juice.

GAS

A gas is matter that has no specific shape or size of its own. Some one can feel gas when the wind blows. The wind is moving air. Air is many gases mixed together. Gases occupy all the space in the container holding them. Examples of gases are oxygen, hydrogen and chlorine.

CHANGES OF STATES OF MATTER

Matter can change its state from one form to another .Matter exist in three states: Solid, Liquid and Gas. Changes of states of matter are caused by alterations in temperature and pressure.

MELTING AND FREEZING

Melting is the change from solid to liquid state .When a solid is heated the particles gains energy and vibrate faster. Eventually, they break free from their fixed positions and begin to move in clusters. A solid will change to a liquid at the melting point. The temperature at melting point must remain constant until all the solid has changed to liquid.

The melting point of a solid tells us how strongly its particles are held together. Substances with high melting points have strong forces between their particles. Those with low melting points have weak forces between their particles.

Freezing is the opposite of melting. For any particular substance, the melting point is the same as its freezing point. The freezing point of pure water, 0˚C, is the melting point of ice. It is not affected by atmospheric pressure.

BOILING AND EVAPORATION.

Boiling is the change of a liquid to vapour. When a liquid boils, it evaporates. When a liquid is heated, the molecules gain energy and move faster. Some of the molecules at the surface of the liquid attain enough energy to escape into the air. As more of the liquid molecules escape to form a gas, the liquid is said to evaporate. This happen at the boiling point of the liquid.

The temperature at which a liquid boils tells us how strongly the particles are held together in the liquid. Liquids with high boiling points have stronger forces between their particles, than liquids with low boiling points.

Consider the below figure which shows the relationship between the three states of matter.

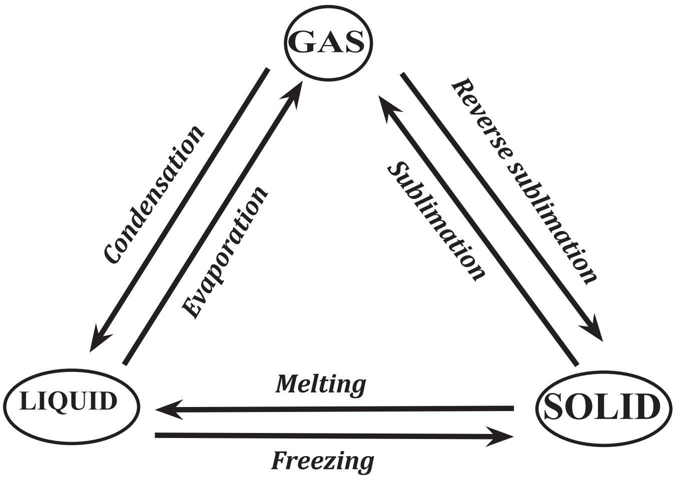


Fig. 5.1; relationship between the three states of matter.

Condensation, reverse sublimation and freezing are ***exothermic processes*** where by evaporation, sublimation and melting are ***endothermic processes***.

Exothermic processes gives off heat energy while endothermic processes requires heat energy.

Experiment 5.1

Aim: To observe the change of matter from one state to another

Apparatus: condenser, source of heat, beaker. Other materials: ice cubes.

Procedure

1. Put some ice in a beaker and heat. Carefully observe the change that take place.
2. Heat the substance further until it changes into vapour.
3. Pass the vapour through a condenser. Record your observations.

Discussion questions

1. How do ice look like before heating?
2. Explain the shape of ice when it melts and when it vapourizes.
3. Explain what happens when vapour is passed through a condenser.

PARTICULATE NATURE OF MATTER.

For many years scientists were not sure whether matter was continuous or particulate. Later, it was discovered that matter is made up of particles. This was proved by a phenomenon known as ***Brownian motion.*** It can be observed in liquid and gaseous substances. Therefore it is now known that matter is not continuous.

Matter is made up of many tiny particles put together. The tiny particles which make up matter are either ions, atoms or molecules. We can not see atoms or molecules by our eyes. They are very small. We can prove that matter is particulate by doing experiments. If you open a bottle containing a strong perfume, the smell of perfume will soon be detected in other parts of the room. If you weigh a bottle of perfume before and after opening a bottle you will not be able to detect any mass change. The particles that were lost must be very small as there is no detectable mass change.

You may dissolve some sugar in water. The sugar seems to disappear completely. When you taste the water, it will be sweet, meaning that the sugar is still there. But in the water, the sugar is divided into such small particles that we can't see them with our eyes.

BROWNIAN MOTION

Brownian motion is an irregular motion of tiny particles suspended in a fluid. In 1827, a Botanist called Robert Brown observed through microscope that pollen grains suspended in water moved short distances in an irregular manner.

Experiment 5.2

Aim: to observe Brownian motion

Materials: Pollen grains, distilled water, microscope, microscope slide, dropper.

Procedure:

1. Put a drop of distilled water on a microscope slide.
2. Put some pollen grains on top of the drop of water.
3. Place the slide under a microscope and observe any movement of the pollen grains.

Discussion questions

1. What kind of movement have you observed?
2. Why do the grains move this way?

Pollen grains move in a random zigzag motion. This is because they are constantly bombarded by water particles which cannot be seen even through a laboratory microscope. This shows that matter is particulate in nature and not continuous in form.

Activity 5.1: To show that a coloured compound is made up of small coloured molecules.

Procedure:

1. Dissolve about 1.0g of purple potassium manganate (VII) crystals in $1dm^3$ of water. Stir until it all dissolves. Label it solution A ***(figure 5.2)***.
2. Measure $1000\ cm^3$ of solution. Label this solution B.
3. Repeat the process four more times, taking $100\ cm^3$ of the subsequent dilute solution at each step. You will finally get solution samples of different colour intensity A,B,C,D,E and F ***(Figure 5.2)***.

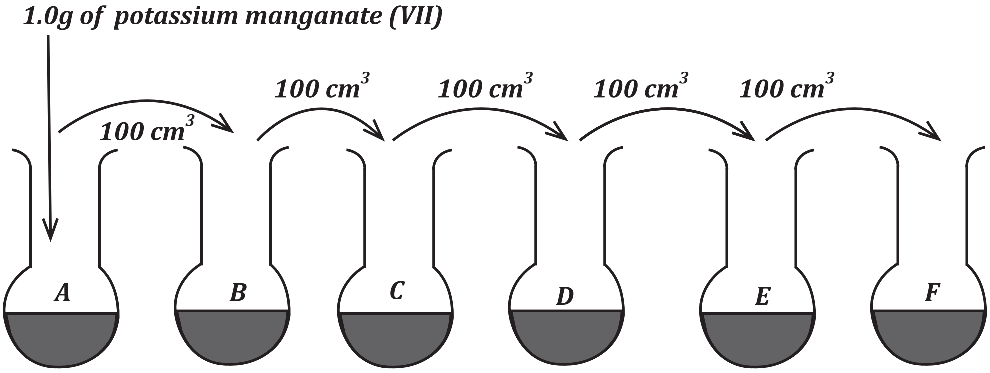


figure 5.2; observing the colour change when a solution of potassium manganate (VII) is diluted several times over.

When potassium manganate (VII) dissolved, its particles spread throughout the water. The resulting solution was dark purple. When the purple solution was diluted, the particles spread further apart, and there were fewer in each subsequent diluted solution. As the solution is diluted the colour gets less intense paler and paler pink. But you will still see a pink colour. Hold piece of white paper behind the solution if you are not sure.

The experiment shows that a small amount of a coloured substance can give colour to a large volume of water. It also suggests that molecules are extremely small.

Activity 5.2: to show that liquids are made of particles

Procedure

1. Place a few drops of ether on a watch glass and observe what happens after a short time.
2. Spray perfume at one corner of the classroom.

Discussion questions

1. Can you detect the smell of ether from a distance?
2. If you stay in any corner of classroom can you detect the smell of perfume?

Spread of the ether and perfume so that they are detected further away from their original positions is called ***diffusion***. Diffusion is the movement of particles from an area of high concentration to the area of low concentration. Therefore liquids are made of particles.

KINETIC NATURE OF MATTER.

Enough evidence shows that matter exists as particles. These particles are atoms, ions or molecules, and are in constant motion. All matter is always in motion. Solids appear to be stationary, but their particles keep on vibrating all the time. At high temperatures, the movement of particles is faster. Temperature changes may alter the state of matter. Matter exists in three states; solids, liquids and gaseous. At very low temperatures, matter tends to exist as solid. As temperature rises, solids change to liquids, and liquids to gases. The reverse is possible when temperature is lowered.

The gaseous state is one in which the particles are moving independently of each other. They move in all directions and at great speeds. Gas particles maintain large distances from each other. They collide as they move. A gas will speed out to fill all the space in which it is contained. A gas cannot have any shape of its own. It is possible to compress a gas because the particles are far apart.

Particles in a liquid are much closer than those in a gas. Forces of attraction exist between the molecules. Molecules in a liquid can move in all directions but not as fast as gas molecules. A liquid may have a definite volume, but does not have a definite shape. A liquid will always take the shape of its container.

Particles in a solid are very close together. The forces of attraction between solid particles are large. Movement of particles from one place to another is not allowed. Rotations and vibrations of solid particles are allowed. In a solid the particles are arranged in a fixed pattern. A solid can therefore have fixed shape. Strong force must be applied to alter the shape. You can't compress a solid but you can hammer it into a different shape.

Consider the below figure which shows particles in solid, liquid and gaseous state

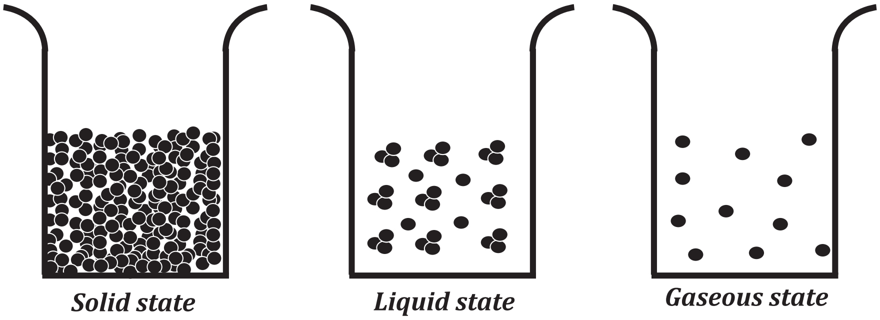


Figure 5.3 particles in solid, liquid and gaseous state

The way in which particles behave in solids, liquids and gases is called ***kinetic molecular behavior.***

Activity 5.3 to demonstrate the kinetic nature of matter.

Procedure:

1. In groups of about twenty, sit closely together so that your shoulders touch.
2. Can you individually move about freely?
3. Move into any empty classroom and walk about freely. How far are you able to move around?
4. Go into the field and move about. How far are you able to move?

Discussion question

Which of your movements in the above steps compares to movement in solids, liquids and gases respectively?

Students' movement is limited when they sit close together. This can be compared to particles of a solid that can only vibrate but remain in their respective positions. When the students move about in a classroom, they have a measure of freedom to move, but cannot go beyond the walls of the room. Their movement is similar to that of particles in a liquid, which can move about freely, but only within the container holding the liquid.

The students in the field can move far apart from each other within the field. If the field is not fenced, they can move beyond its boundaries. This is similar to the movement of gaseous particles which are far apart from each other and fill the container holding the gas. If the container is opened, the gas particles move out of it.

5.3 PHYSICAL AND CHEMICAL CHANGES.

Physical changes are temporary changes in the physical properties of a substance. Chemical changes are permanent changes in the chemical properties of a substance.

CHARACTERISTICS OF PHYSICAL AND CHEMICAL CHANGES.

Physical changes are concerned with energy and states of matter. A physical change does not produce a new substance. Changes in state or phase (melting, freezing, evaporation, condensation, sublimation) are physical changes. Examples of physical changes include crushing a can, melting an ice cube and breaking a bottle.

Chemical changes take place on the molecular level. A chemical change produces a new substance. Examples of chemical changes include combustion (burning), cooking an egg, rusting of iron, mixing hydrochloric acid and sodium hydroxide to make salt and water.

Consider the below table which shows examples of physical and chemical changes.

Table 5.1; examples of physical and chemical changes

Physical changes	Chemical changes
Melting an ice cube.	Burning wood
Crumpling a sheet of paper	Ripening of green tomato
Casting silver in a mold	Milk goes sour
Aluminum foil is cut in half	Jewelry tarnishes
Butter melts on warm toast	Bread becomes toast
Water evaporates from the surface of the ocean	Rust forms on a nail left outside.

Physical changes	Chemical changes
A juice box in the freezer freezes.	Gasoline is ignited.
Dissolving salt in water	Food scraps are turned into compost in a compost pile.
Water boiling	You take an antacid to settle your stomach
Cutting carrots	Digestion of food
Grinding meat to make hamburger	Frying an egg
Frost forming on the window	Creation of a metal alloy
Fog forming at night	Baking a cake
Making tea	A battery turning on flash light
Dissolving sugar in water	Mixing vinegar with baking soda.
Milo dissolves into hot milk	Fireworks explode to form colorful light and loud sounds.
A plate is dropped and shatters	An egg is cooked to become a white and yellow solid.
Grass is mowed	Spoiling food
Metal knife is sharpened	Burning leaves
Breakfast cereal goes soggy	Burning of candle
Heating an iron bar until it glows red hot.	Burning of camphor
Melting of wax	Rancidification of butter
Heating of zinc oxide	Photosynthesis
Heating of camphor	Combustion of fuel
Heating of ammonium chloride	Electrolysis of water.
Dissolution of sulphur in carbon disulphide	Hydrogen burning in oxygen
Cutting of wood	Decomposition of potassium chlorate.

Experiment 5.3

Aim: To demonstrate types of changes in matter.

Apparatus: Beaker, evaporating dish, source of heat, crucible.

Other Materials: Sugar, water, candle, wax, paper.

Procedure;

1. Pour water into a beaker. Dissolve some sugar in it. Pour the solution in an evaporating dish and evaporate the solvent. Let the resultant product cool. What is the resultant product? Does it look like the sugar you dissolved?
2. Heat some candle wax in a crucible until it melt. Let it cool. Is the final product different from the initial one?
3. Burn a piece of paper until it is completely burnt. What is the final product?

Is it the same as the initial substance?

4. Put some sugar in a crucible, heat and let it melt completely until it changes color. Is the final substance the same as the initial product?

Discussion questions

1. Explain which of the changes observed were physical and chemical respectively?
2. What are characteristics of physical and chemical changes?

Consider the below table which shows differences between physical and chemical changes.

Table 5.2 Differences between physical and chemical changes;

Physical change	Chemical change
Alters only the physical properties but the molecular composition remains totally unaltered.	Alters the chemical properties and molecular compositions.
No new substances are formed as a result of physical change	A new substance is always formed as a result of chemical change.
The change is temporary	The change is permanent
The change is reversible	The change is irreversible
No energy change take place as a result of physical change	Energy changes take place as a result of chemical change.
Mass of a substance remains totally unaltered	Mass of substances gets totally altered during chemical change.

SUMMARY

(a). Matter is anything that occupies space and has mass.
(b). Mass is a measure of a quantity of matter in an object.
(c). Matter exists in three physical states of gas, liquid and solid.
(d). Matter can change its state from one form to another.
(e). Matter is particulate and not continuous in form.
(f). Brownian motion is an irregular motion of tiny particles suspended in a fluid.
(g). Diffusion is the movement of particles from an area of high concentration to the area of low concentration.
(h). All matter is always in motion.
(i). Physical changes are temporary changes while chemical changes are permanent changes.
(j). At high temperature, the movement of particles is faster while at low temperatures the movement of particles is slow.

REVIEW QUESTIONS

1. Choose the most correct answer for the following questions;

i. .... Is anything that has mass and occupies space.

A. Weight B. Element C. Matter D. Atom

ii. A measure of the quantity of matter in an object is called.........................

A. Mass B. Weight C. Space D. Gram

iii. Matter exists in...................... different physical states

A. One B. Three C. Two D. Four.

iv. A gas can be .... to form a liquid

A. Condensed B. Evaporated C. Solidified D. Neutralized.

v. Matter is ___________ in form

A. Continuous B. Particulate C. Solidified D. Neutralized.

2. Match the correct item from ***List B*** with the corresponding one in ***list A***.

List A	List B
i. Solid	A. Has no definite shape
ii. Liquid	B. Take the shape of the container
iii. Gas	C. Retains a fixed shape
iv. Melting point	D. The temperature at which a liquid boils to form a gas
v. Boiling point	E. The temperature at which a solid melts to form a liquid
vi. Sublimation	F. A change from solid directly to a gas
vii. Brownian motion	G. Movement of particles from high concentration to low concentration
viii. Diffusion	H. Irregular motion of tiny particles suspended in a fluid
	I. The point of melting substances.
	J. The point of liquefaction.

3. Write ***TRUE*** or ***FALSE*** in the space provided against each question.

i. The kinetic molecular behavior is the way particles behave in solids, liquids and gas

ii. Physical changes are irreversible, they only affect chemical properties of substances

iii. Chemical changes are reversible, they affect physical properties of a substance.....................

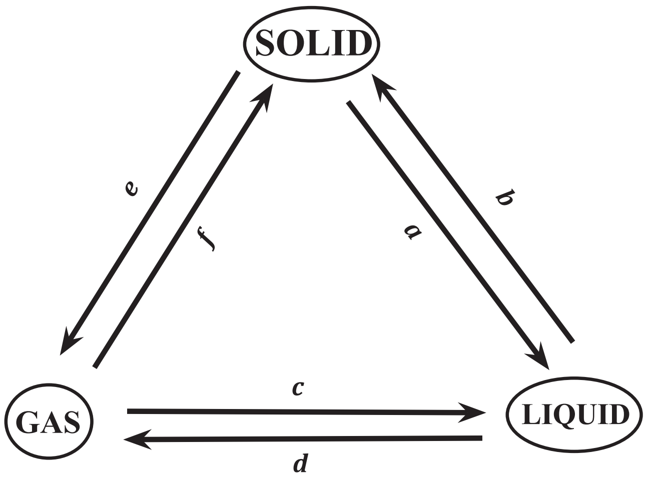
iv. A physical change result in the production of new substances while a chemical change does not............................

v. Matter is made of particles that are able to display some form of movement..................

4. Copy and fill the following table. One of the item has been done for you.

PROPERTY	GAS	LIQUID	SOLID
Shape			
Movement of particles			
Compressibility		Not easily compressible	
Spaces between particles			
Flow			
Volume			

5. Outline the differences between physical and chemical changes.
6. Copy the following diagram



Name the process a-f above.

7. Categorize the following changes as either chemical or physical.
   a. Ammonium carbonate is dissolved in water gradually with stirring. The solution becomes very cold.
   b. Exactly 6g of copper filings are reacted, the mass is found to be 5.55g.
   c. A small piece of sodium was dropped into water. It dated about on the water surface and finally it burst into flames.
   d. Copper(ii) sulphate is heated strongly; it changes from blue to white. On addition of water to the white substance, it changes back to blue.
   e. When ethanol is mixed with water, a homogenous solution is formed. When the solution is distilled, the two liquids are obtained.
   f. If water is kept in a deep freezer it solidifies to ice. If the ice is kept in the

sun, it liquefies to water.

g. Water is added drop wise to dry calcium oxide on a watch glass. Heat is developed, steam is formed, the oxide cracks and puffs up and finally crumble to a powder about three times as bulky.

8. Find out about how a refrigerator works. It contains a liquid called a refrigerant that has a low boiling point. Explain why a refrigerator is cold inside but hot at the back outside.

Chapter 6

Elements, Compounds and Mixtures

6.1 ELEMENTS AND SYMBOLS

An ***element*** is a pure substance which cannot be split into simpler substances by simple chemical means. An element is a substance because it has the same composition throughout. When two or more elements combine they form a compound.

There are about 105 known chemical elements and some of them are extremely rare. Some elements do not exist in the earth at all and they are man made in very small quantities. Examples of elements are; sulphur, iron, copper, aluminium, tin, lead and magnesium.

NAMES AND SYMBOLS OF ELEMENTS.

Each element has a short hand symbol to represent it. These symbols are used internationally. In 1808, Dalton published a list of elements in which the symbols were very difficult to write. In 1811 Swedish chemist ***Johns Jacob Berzelius*** published a system of chemical symbols based on letters and this system is still used today.

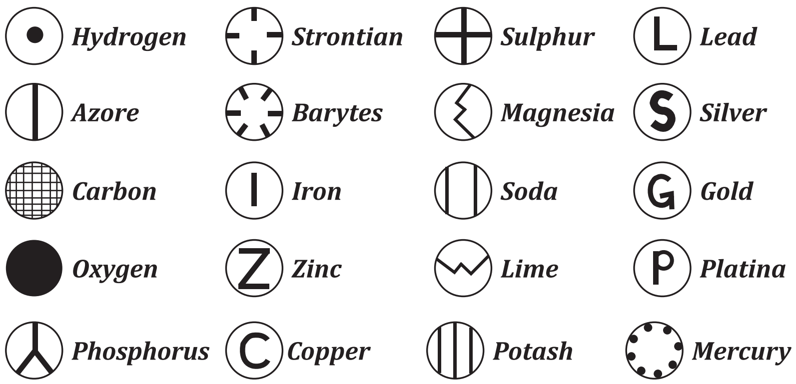


Figure 6.1; Dalton's elements symbols.

The rules by which the symbols were assigned can be explained as follows:

1. An element can be represented by a symbol that is derived from the first letter of its English name. Examples; Iodine(I), Fluorine (F), Hydrogen (H), Nitrogen (N), Phosphorus (F), and Sulphur (S).
2. Names of different elements may have the same first letter, example copper and calcium. It is thus necessary to differentiate the elements. Therefore, another letter, usually the second from the name is used with the first one. The first letter will be capital while the second will be a small letter. Examples of elements whose symbols have two letters are; chlorine (Cl), cobalt (Co), magnesium (Mg), manganese (Mn), and aluminum (Al).
3. In some cases, the symbols can be derived from Latin names instead of the common English names. Examples of elements whose symbols are derived from their Latin names are: sodium is natrium in Latin and it's symbol is Na, potassium is kalium in Latin its symbol is K, copper is cuprum in Latin and its symbol is Cu, iron is ferrum in Latin and its symbol is Fe, mercury is hydrargyrum in Latin and its symbol is Hg, silver is argentum in Latin and its symbol is Ag, gold is aurum in Latin and its symbol is Au, tin is stannum in Latin and its symbol is Sn and lead is plumbum in Latin its symbol is Pb.

Consider the table below which shows symbols of first twenty elements.

Table 6.1 the symbols of first twenty elements.

Element	Symbol
Hydrogen	H
Helium	He
Lithium	Li
Beryllium	Be
Boron	B
Carbon	C
Nitrogen	N
Oxygen	O
Fluorine	F
Neon	Ne
Sodium	Na
Magnesium	Mg
Aluminium	Al
Silicon	Si
Phosphorus	P

Element	Symbol
Sulphur	S
Chlorine	Cl
Argon	Ar
Potassium	K
Calcium	Ca

IMPORTANCE OF CHEMICAL SYMBOLS

The following are importance of chemical symbols in the field of chemistry.

1. It is easy to understand elements quickly instead of having to memorize full names.
2. It is now possible to write equations in short form instead of having to write each element with its full name.
3. The quantity of the element being referred to can be known readily. Example, the symbol Ag, could refer to one atom or more mole of silver.

Activity 6.1; to test the electrical conductivity of elements.

Procedures;

1. Connect four dry cells and a 2.5v bulb in a series circuit. The dry cells will produce about 6.0 volts.
2. Connect the terminals as shown in figure 6.2
3. Put the element to be tested in a beaker. The following substances will be needed: Sulphur powder, iron filings, copper, aluminium, tin, lead, magnesium.
4. Hold the two terminals of the circuit so that they touch the element to be tested.
5. Observe whether the bulb lights or not. Compare the intensity of the light as different elements are tested.
6. Which of the following elements conduct electricity?

Complete the following table with yes or no.

Element	Electrical conductivity
Sulphur	no
Iron	
Copper	
Aluminium	
Lead	
Tin	
Magnesium	

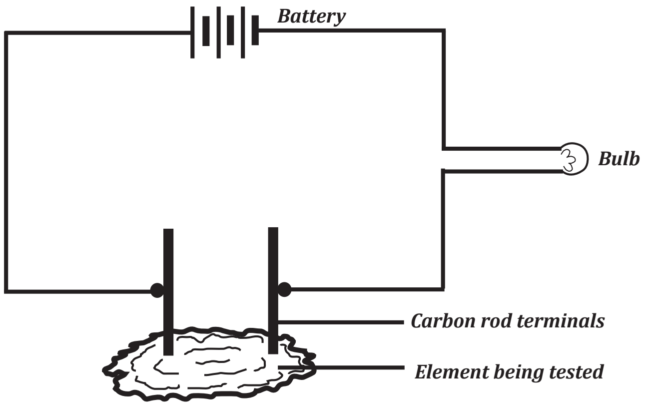


Figure 6.2; electrical circuit for testing elements

DISCUSSION QUESTIONS

1. Between metals and non-metals which are poor electric conductors?
2. Can you explain the reason for the answer you have given above?

MOLECULES

A ***molecule*** is an electrically neutral group of two or given atoms held together by chemical bonds. Atoms of oxygen always exist in pairs. Sulphur exists as rings of eight atoms and phosphorus as tetrahedral of four atoms,

Consider the below diagrams for molecules;

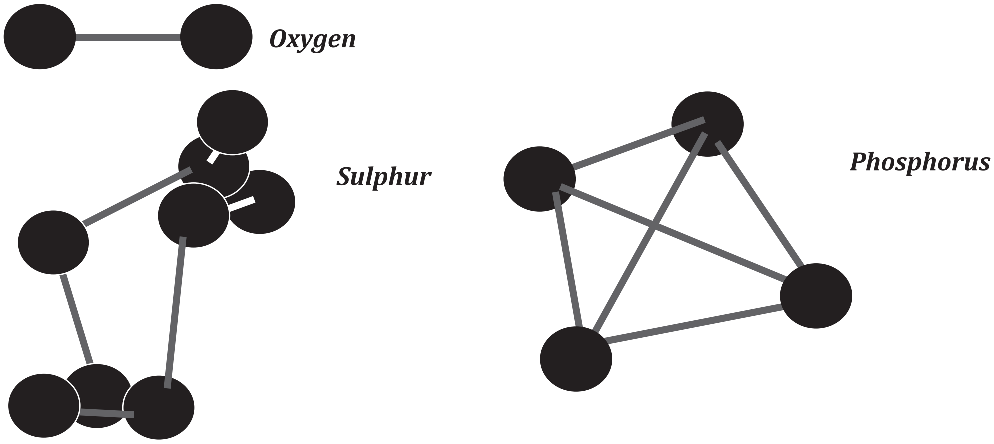


Figure 6.3; oxygen, sulphur and phosphorus molecules.

Table 6.2 examples of molecules

Elements	Atomic symbol	Molecular symbol
Oxygen	O	O_2
Sulphur	S	S_8
Phosphorus	P	P_4
Nitrogen	N	N_2
Fluorine	F	F_2
Chlorine	Cl	Cl_2
Bromine	Br	Br_2
Iodine	I	I_2
Hydrogen	H	H_2

6.2 COMPOUNDS AND MIXTURES

A compound is a pure substance that is made up of two or more elements chemically combined together. The combination is always in a fixed ratio. Examples of compounds are carbon dioxide, water, table salt and limestone.

A mixture is a combination of two or more substances physically combined together. Combination in a mixture is in any ratio. Since mixtures are not chemically combined, they can be separated by physical means. Examples of mixtures are: oil and water, muddy water, sand and salt.

Mixtures can be ***homogeneous*** or ***heterogeneous***. A homogenous mixture has uniform composition, appearance and properties. Example a small amount of sugar dissolved in a cup of water. After stirring, every section of the solution is identical in composition, appearance and physical properties.

A heterogeneous mixture has different compositions, appearance and properties at various points in the mixture. Example in a mixture of ice and water, at one point we find a solid while at another we find a liquid.

Activity 6.2; to prepare a mixture of iron and sulphur.

Procedure;

1. Measure approximately 5g of iron filings and 5g of sulphur powder. Mix them in a mortar and grind them with a pestle to mix well.
2. Observe the mixture with a hand lens. Can you distinguish the specks of iron from the specks of sulphur? Does the mixture retain the properties of iron and those of sulphur?
3. Put some of the mixture on a piece of paper. Move a strong magnet below

the paper. What do you observe?

4. Put ¼ of spatulaful of the mixture in a test tube. Add dilute hydrochloric acid. If the gas is evolved ask your teacher how to identify it. Enter your observations in the table.

		Sulphur	Mixture of iron and sulphur	Iron filings
1	Appearance			
2	Effects of a magnet			
3	Action of dilute hydrochloric acid			

You will notice that the properties of the mixture of iron and sulphur are similar to the properties of the separate elements.

Activity 6.3; to prepare a compound of sulphur and iron.
Procedures;

1. Prepare about 6g of a mixture of sulphur powder and iron filings by measuring out 3g of each. Grind them separately using a pestle and mortar then mix them together thoroughly.
2. Put this mixture in a hard glass test tube.
3. Heat the mixture strongly until you are convinced that the reaction is over.
4. Cool the product. Empty the product into a mortar and break it into small pieces.
5. Put one of the pieces of the product on a watch glass and observe it with a hand lens. Is it different from the original mixture?
6. Put another piece on a piece of paper and bring a magnet near it. Is it affected by the magnet?
7. Put a third piece of the product into a test tube and add dilute hydrochloric acid to it. If a gas is evolved, ask your teacher to show you how to identify it.
8. Enter your observations in a copy of the table below:

		Compound of iron and sulphur	Mixture of iron and sulphur
1	Appearance		
2	Effects of a magnet		
3	Action of dilute hydrochloric acid		

Note;

When hydrochloric acid is added to iron sulphide, hydrogen sulphide gas is produced. This is a very toxic gas and teachers must ensure that this is done on a very, very small scale preferably in a fume cupboard, a very well ventilated room or even outside. All remaining iron sulphide should be collected and disposed off. If it is left in a sink it can react later if acid is put down the sink.

Discussion question

Have you noted any difference in properties between iron sulphide and a mixture of iron and sulphur? Explain the differences you have noted.

Table 6.3; Differences between compounds and mixtures.

Compounds	Mixtures
Components cannot be seen separately	Components may be seen separately.
The constituent elements can not be separated by physical methods	Components can be separated by physical methods.
Compounds have definite compositions by mass of the elements. The proportions can not be changed.	Mixtures may vary widely in composition. The elements are mixed in any proportion.
A chemical change is involved when new substances are formed.	No chemical change takes place when a mixture is formed.
The properties of compounds are very different from those of the individual elements	The properties of a mixture are those of the individual elements.

6.3 SOLUTIONS, SUSPENSIONS AND EMULSIONS

SOLUTIONS

A ***solution*** is a homogenous mixture of two or more substances in which one of the substances is a solute and the other a solvent. A ***solvent***; Is the component of a solution that dissolves a solute. A ***solute***; is the component in a solution that is dissolved in a solvent. Example; table salt dissolves in water to form a solution. The salt is a solute and water is a solvent.

When common salt (sodium chloride) is added to water and stirred, it can be seen to disappear from view. The salt is still there but it is ***dissolved*** in water. What can only be seen is the ***solution***. The mixture is said to be ***homogenous*** because it is the same throughout.

When sand is mixed with distilled water, neither of the two disappears. The sand can be seen to spread throughout the water. The mixture is said to be ***heterogeneous***, because it is not uniformly mixed. Both components of the mixture can be seen. Sand is not dissolved in the water. It is said to be suspended in the water.

TYPES OF SOLUTIONS

A solution can be unsaturated, saturated or supersaturated.

- ***Unsaturated solution*** is the solution that can still dissolve more solute, at a given temperature.
- ***Saturated solution***; is the solution that can dissolve no more solute at a given temperature.
- ***Supersaturated solution;*** is the solution that temporarily holds more solute than the saturated solution at a given temperature.

Experiment 6.1

Aim: To distinguish between saturated, unsaturated and super saturated solutions,

Materials;

Common salt, beaker, Bunsen burner, tripod stand, spatula, glass rod, trough.

Procedure;

1. Measure about 100 cm^3 of water and pour it in a beaker.
2. Add about a spatulaful of the common salt and stir. Does all the salt dissolve?
3. Continue to add more salt to the solution, stirring until no more salt can dissolve.
4. Place the solution on the tripod stand and heat gently, as you continue stirring.
5. Stop heating when the salt dissolves.
6. Place the beaker in a trough that is half filled with cold water and allow cooling for 5 minutes. Record your observation.
7. Note that saturation depends on temperature.

DISCUSSION QUESTIONS.

1. What type of solution do you get when a spoonful of table salt dissolves in 100 ml of water?
2. What solution do you have, at room temperature when no more salt can dissolve?
3. What name do you give to the final solution?

On adding a spoonful of salt to a beaker of water, unsaturated solution is formed. As you dissolve more salt in the unsaturated solution, a point is reached when the water loses its ability to dissolve any more salt at room temperature. It becomes a saturated solution. If you heat the water, it increases its solubility and more salt can be added. If this solution is cooled, the salt particles reappear in the solution. This solution is a supersaturated solution.

CLASSIFICATION OF SOLUTIONS INTO THREE STATES OF MATTER

Solutions may be solids, liquids or gases, consider the below table which expresses examples of types of solutions.

Table 6.4; examples of types of solutions.

		SOLUTES		
		SOLID	LIQUID	GAS
Solvents	Gas	Ammonium chloride sublimes in air to form a solution	Water vapour in air	Nitrogen and other gases in air
	Liquid	Glucose in water and table salt in water	Ethanol in water and various hydrocarbons in each other.	Carbon dioxide in water
	Solid	Steel and other metal alloys	Mercury in gold and hexane in paraffin wax	Hydrogen in metals.

SUSPENSIONS

A ***suspension*** is a heterogeneous mixture of liquid and fine particles of a solid in which the solid does not dissolve but is suspended in the liquid. The suspended particles are slightly visible and will settle over time if left undisturbed. Examples are muddy water and chalk powder in water.

When suspended particles stay spread out in the water and do not settle, a jelly like substance is formed and it is called a ***colloid solution***. Liquid droplets or fine solid particles floating in a gas are called ***Aerosols***.

Experiment 6.2;

To prepare suspensions. Materials; mud, chalk powder, water, copper carbonate, beakers.

Procedure;

1. Put each of the solid substances in a separate beaker.
2. Add enough water to cover and stir until the solid mixes with the water.
3. Leave the mixture to settle. What do you notice?

Discussion questions

1. What place in real life situations can we find mixtures?
2. Can you differentiate mixtures from solutions?

Examples of suspensions are:

Syrup, some paints, body sprays, insecticides and blood in our bodies.

DIFFERENCES BETWEEN SUSPENSIONS AND SOLUTIONS

Consider the table below:

Table 6.5; differences between suspensions and solutions.

Solutions	Suspensions
Solutions are homogenous	Suspensions are heterogeneous
Solutions are transparent	Suspensions are opaque or not clear
Particles completely dissolve in solutions	Particles separate on standing
Components in solutions can be separated by evaporation	Components in suspension can be separated by filtration.
The particles of a solution are at the ion or molecular level and can not be seen by naked eyes	The particles of a suspension can be seen by the naked eye
Light can pass through a solution	Light cannot pass through a suspension

EMULSIONS

An emulsion is a mixture of liquids that does not dissolve each other well. When two liquids are mixed together, they may dissolve in one another, or remain as separate layers. Liquids which mix together completely are said to be ***miscible***. Miscible liquids form homogenous mixtures. Ethanol and water mix together completely to form a single homogenous solution.

Other liquids will never mix even if they are shaken together. They will separate out into two layers, where the less dense is on top. Water and kerosene are such liquids. They are said to be ***immiscible.***

Examples of emulsions are ethanol and cooking oil mixture, milk, paint.

Activity 6.4; to prepare a simple emulsion

Procedures;

1. Add warm water to the jar until it is a quarter full.
2. Add a few drops of olive oil or cooking oil.
3. Replace the lid and shake the jar thoroughly. What happens when you leave the jar to stand?
4. Now add a couple of drops of soap solution and shake again. What happens now when you leave the jar to stand?

6.4 METHODS OF SEPARATING MIXTURES

Mixtures contain useful substances mixed with unwanted materials. In order to obtain these useful substances, chemists often has to separate them from impurities.

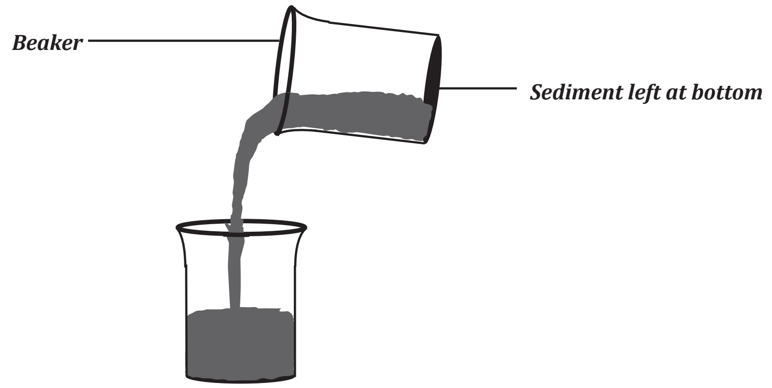
The formation of a mixture tends to occur easily and is said to be ***spontaneous***. Separating mixtures is more difficult to do. One of the jobs done in chemistry is to separate mixtures into pure substances. Some of the methods of separating mixtures are; decantation, filtration, evaporation, distillation, sublimation, chromatography and centrifugation.

DECANTATION

Decantation is the process of separating a heterogeneous mixture of liquid and solid by pouring out the liquid only and leaving the solid at the bottom of the container.

The method of decantation works well when the solid component is made up of large particles.

Consider the diagram below;



Decantation is usually used in blood tests which require the clearer part of the blood separated from its solid components.

Experiment 6.3

Aim; to separate substances by decantation

Materials; muddy water and beakers

Procedures;

1. Put some muddy water in a beaker and let it stand.
2. Carefully pour out the clear water from the top into another beaker.

DISCUSSION QUESTIONS

1. What is observed when the muddy water is left to stand for some time?
2. Explain how you can apply decantation in real life situation.

FILTRATION

Filtration is the method used to separate a heterogeneous mixture of a solid and a liquid. The solid is separated from the liquid using a porous filter, such as filter paper. The solid obtained is the ***residue*** and the liquid is the ***filtrate***

Activity 6.5;

To separate chalk and salt from their mixture by filtration

1. Grind some chalk in mortar using a pestle.
2. Mix one spatulaful of the chalk powder with two spatulafuls of salt in a 250cm^3 beaker.
3. Add water and stir until all the salt has dissolved.
4. Follow the steps as outlined in figure 6.5 to fold a filter paper.
5. Put the filter paper in a funnel, and add some drops of water to fix the filter paper on the walls of the funnel.
6. Set up the apparatus as in figure 6.6.
7. Pour the contents of the beaker onto the filter paper carefully. Don't touch the filter paper. Don't stir.

Consider the below diagrams;

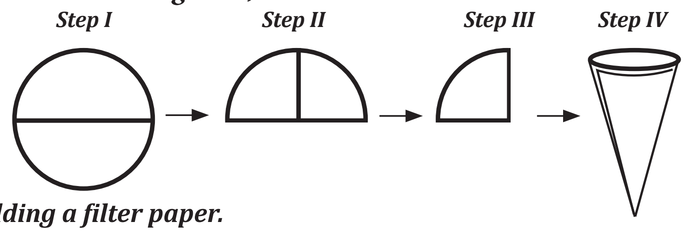


Figure 6.5; folding a filter paper.

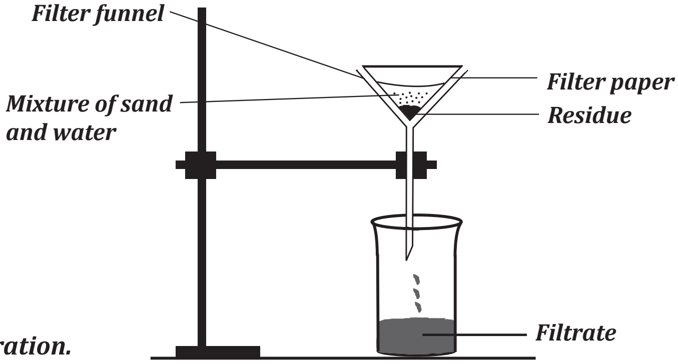


Figure 6.6 filtration.

Discussion questions

1. What is the nature of the filtrate obtained?
2. What is the residue?

During the filtration, the solid chalk remains on the filter paper and the salt solution passes through the small holes in the filter paper.

EVAPORATION

Evaporation is the process of separating a solute from a liquid solution. During evaporation the solvent is converted from liquid to gas through heating and the solute remains as residue. Evaporation is the opposite process of condensation.

Experiment 6.4

Aim; to separate substances by evaporation.

Apparatus; evaporating dish, beaker, tripod stand, Bunsen burner, wire gauze.

Chemicals; salt solution.

Procedures;

3. Pour some salt solution into the evaporating dish and place on a wire gauze on the tripod stand.
4. Heat the solution until almost all the water evaporates.
5. Let the mixture cool completely.

Discussion questions

1. What is the purpose of heating the solution?
2. What do you observe when the solution cools?

When a solution of common salt in water is heated until the water evaporates, white solid crystals of the salt are obtained.

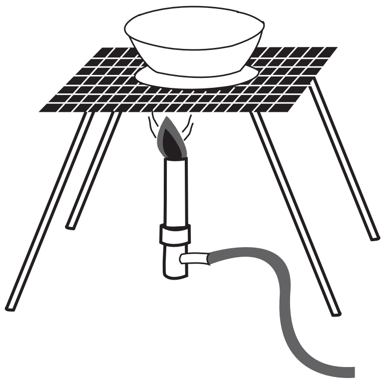


Figure 6.7 evaporation.

DISTILLATION

Distillation is the process of separating out a pure liquid from a mixture by boiling in which a liquid vapourizes and cooled into a distillate. There are two types of distillation. These are simple distillation and fractional distillation.

SIMPLE DISTILLATION

Simple distillation involves heating the solution to its boiling point and cooling the vapour to form a liquid thereby the solid solute remains behind the distilling flask. An example is distilling water from copper (II) sulphate Solution.

Activity 6.6; to obtain pure water from copper (ii) sulphate solution.

1. Put a spatulaful of copper (ii) sulphate crystals in a beaker.
2. Add water and stir until all the crystals have dissolved.
3. Transfer the blue solution into a side arm 250 cm^3 distilling flask.
4. Add a few china fragments to the flask; bubbles of vapour form on their rough surface and ensure smooth boiling.
5. Insert a rubber bung carrying a thermometer as shown in figure 6.8.
6. Connect the side arm of the flask to a Liebig condenser.
7. Set up the apparatus as seen in figure 6.8
8. Allow cold water into the condenser through B and out of it through C.
9. Heat the distilling flask and note the temperature at which the solution starts to boil.
10. Collect the distillate in a conical flask or a beaker.

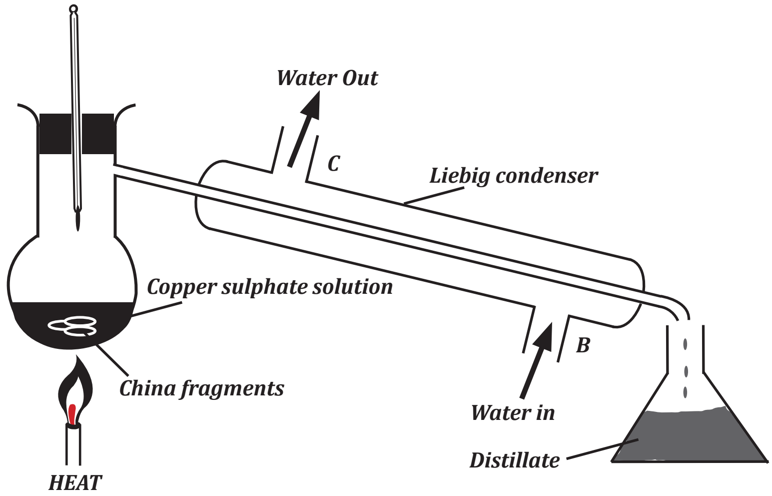


Figure 6.8 simple distillation apparatus.

During the distillation the copper (ii) sulphate remained in the flask. The liquid collected, called the distillate, is distilled water. The temperature recorded by the thermometer during the experiment should be 100˚C. The Liebig condenser has cold water running through an outer tube to cool the steam down and cause it to condense.

Fractional distillation

Fractional distillation is the method of separating a mixture of two or more liquids that form a homogenous solution by means of fractionating column. The method works well if the boiling points of the liquids are not too close.

A mixture of ethanol and water is a good example for demonstrating fractional distillation. These liquids mix completely to form a homogenous solution. The boiling point of pure ethanol is 78˚C and that of pure water is 100˚C. In simple distillation only one liquid is involved while in fractional distillation, several liquids may be involved. The liquid with the lower boiling point is the first to be collected as a distillate.

Experiment 6.5

Aim: To separate two miscible liquids

Apparatus: Liebig condenser, fractionating column, beakers, distillation flasks, retort stands and clamps, thermometer, Bunsen burner, tripod stand, wire gauze, conical flask.

Chemicals: ethanol, water.

Procedure;

1. Set up the experiment as shown in figure 6.9
2. Heat the ethanol and water mixture. Note the temperature at which the first liquid is distilled.
3. Record all observation.

Discussion questions

1. Which liquid distills first and at what temperature?
2. What role does the fractionating column play?

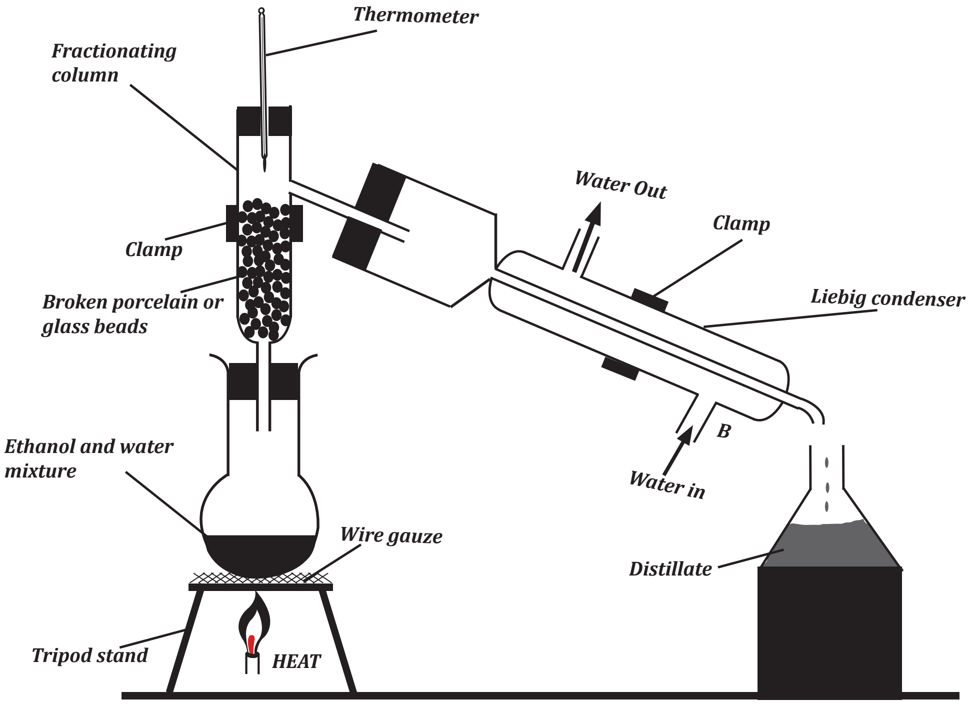


Figure 6.9; fractional distillation.

SUBLIMATION

Sublimation is a direct change from the solid state to the gaseous state usually on heating. The reverse process of direct change from the gaseous state to the solid state is called ***reverse sublimation*** or ***deposition***. The solid that forms after the vapor cools is called a ***sublimate***.

The process of sublimation can be used to separate mixtures where one of the substances sublimes. Iodine and ammonium chloride are the commonest examples of such substances.

Activity 6.7; To separate a mixture of sodium chloride and ammonium chloride by sublimation.

Procedure;

1. Put approximately two spatulafuls each of ammonium chloride and powdered table salt in a 250 cm^3 beaker (hard glass pyrex type). Use a glass rod to mix the powder.
2. Place the beaker over a gauze mat on a tripod stand.
3. Cover the beaker with a round –bottomed flask containing cold water.
4. Heat the beaker gently. You should avoid forming lots of smoky fumes by too strong heating.

You will notice that a white solid forms on the cold flask. This is ammonium chloride. The sodium chloride does not change when it is heated.

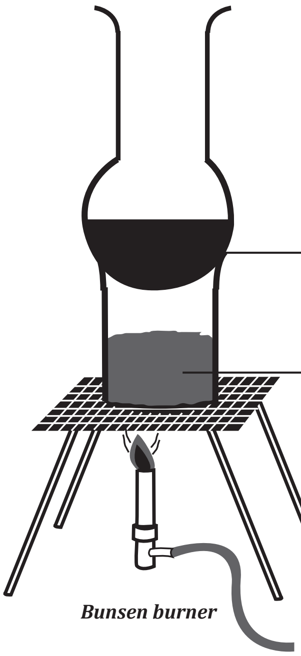


Figure 6.10; sublimation of ammonium chloride

FUNNEL SEPARATION

This is a method used for separating immiscible liquids using a separating funnel. Immiscible liquids are those that do not mix at all when they are kept in the same container. If they have different densities the most dense liquid settles at the bottom while the least dense remains at the top of the separating funnel. When shaken together they make an emulsion but on standing the liquids will separate out into the original layers. By opening the tap, the liquid at the bottom is let out first.

Activity 6.8 to separate a mixture of kerosene and water by the funnel separation method.

Materials: separating funnel, kerosene, water, and beaker.

Procedure

1. Pour water into a beaker
2. Add kerosene and stir
3. Pour the mixture into a separating funnel.
4. What do you observe?
5. Open the tap slowly and drain away the liquid at the bottom first, then the second.

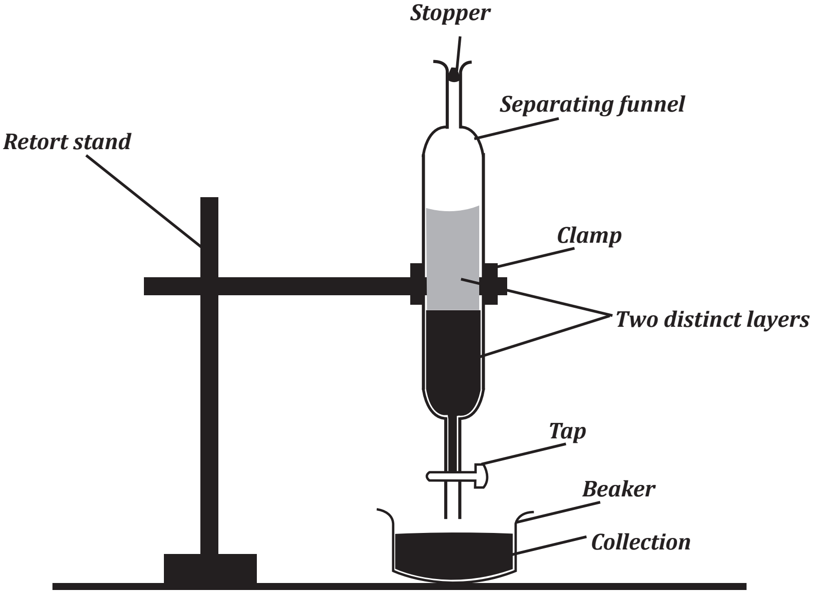


Figure 6.11; separating immiscible liquids.

When kerosene is mixed with water and poured into a separating funnel, it will move to the top of the funnel. The water is drained out first.

SOLVENT EXTRACTION

Solvent extraction is a method used to separate compounds based on their relative solubility. This process can be used for extracting essential oil from plant materials, especially seed, using a liquid that dissolves the oil. The oil is then distilled and the solvent let to evaporate. This process is also known as liquid-liquid extraction.

Experiment 6.6

Aim: to extract oil from groundnuts

Materials: ethanol, groundnuts, pestle and mortar, fractional distillation equipment, beaker.

Procedure

1. Crush a handful of groundnuts in the mortar.
2. Pour in about 200ml of the ethanol and mix well.
3. Allow the mixture to settle for a while then decant the liquid into a beaker.
4. Use fractional distillation to separate the ethanol from the oil.

Discussion question

Explain the importance of using fractional distillation method in the separation of ethanol from oil.

CENTRIFUGATION

Centrifugation is a method of separating a suspension by spinning a sample around very quickly in a machine called ***a centrifuge***, which causes sediment to settle and separate quickly from the liquid it is mixed. In this process the sediments go to the bottom with the liquid above it. The liquid can then be separated by decantation. A centrifuge is also used to separate serum from a blood sample.

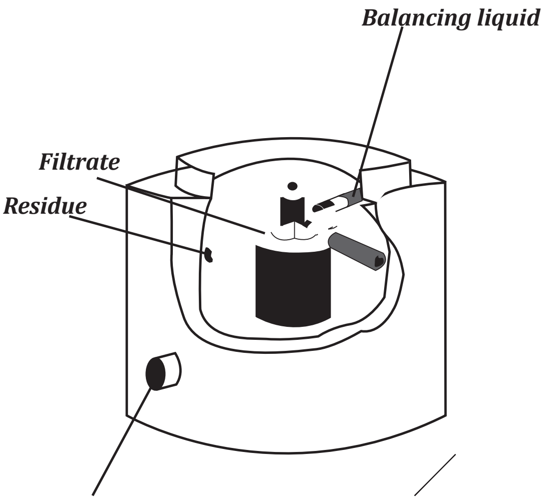


Figure 6.12 Centrifuge

CHROMATOGRAPHY

Chromatography is a method by which different solutes in a solution can be separated. Frequently they are coloured substances in a mixture. This method enables the different solutes, to be separated and identified. This technique was discovered when scientists were extracting coloured dyes from plants. The coloured substances are separated by letting them spread across a piece of filter paper. The substances separate out because they move across the paper at different speeds.

Chromatography works better when a solvent is used. Solvents like water, ethanol or ether may be used for substances which do not dissolve in water. There are several techniques of chromatography in which the simplest of them is called ***paper chromatography.*** Chromatography can also be used to separate colorless substances in which chromatograms can be seen.

A ***chromatogram*** is a pattern of separated colours formed during the process of chromatography. Each coloured ring on the filter paper corresponds to a different solute in the mixture.

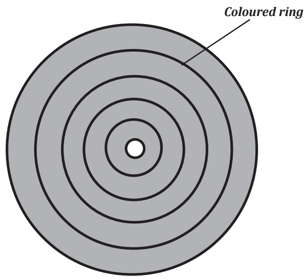


Figure 6.13 chromatogram

Activity 6.9, to separate the coloured substances in variegated leaves by using paper chromatography.

1. Chop small pieces of variegated leaves into a mortar.
2. Add a spatulaful of clean sand (which has been washed by water).
3. Grind with a pestle and add some drops of ethanol as you grind. This process should break down plant cells and release the colour.
4. Squeeze out the green juice as it forms into a small beaker.
5. Continue with the process until you have at least $20cm^3$ of the green juice from the leaves.
6. Place a filter paper over the top of a petri dish, as shown in figure 6.14.
7. Using a teat pipette put one drop of the juice at the centre of the filter paper. Let it dry completely. When it is dry, add another drop of the juice and again, let it dry completely. Continue adding a drop at a time at the same spot, and each time letting it dry.
8. Use another dropper to add ethanol onto the coloured spot a drop at a time. Allow the first drop of ethanol to spread out before you add the next one. Continue adding ethanol drop wise at the same spot until the colours have spread out far enough. Let the chromatogram dry.

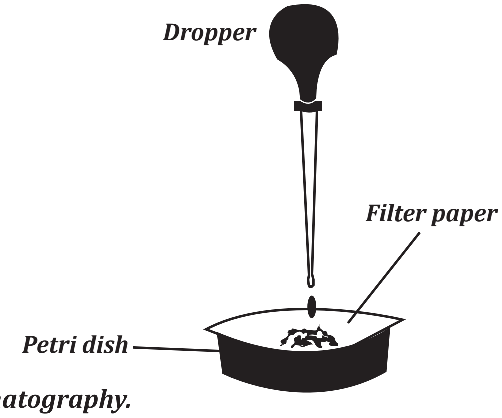


Figure 6.14; chromatography.

In chromatography the moving solvent is called a ***mobile phase*** and may be a liquid or a gas. The material that absorbs the solvent is called a stationary phase and is usually a solid, or a liquid supported on a solid. The mobile phase flows through the ***stationary phase*** and carries the components of the mixture with it. The substance that is separated during chromatography is called the ***analyte***. In paper chromatography, the stationary phase is a very uniform absorbent paper. The mobile phase is a suitable liquid solvent or mixture of solvents.

SUMMARY

(a). An element is a pure substance which can not be split into simpler substances by a simple chemical process.
(b). Chemical symbols are often derived from the Latin or Greek name of the element and may not have much similarity to the common English name.
(c). A compound is a pure substance that is made up of more than one element in a chemical combination.
(d). A mixture is a physical combination of two or more substances in any ratio.
(e). A homogenous mixture has uniform composition, appearance and properties.
(f). A heterogeneous mixture has different composition, appearance and properties at various points in the mixture.
(g). A solution is a homogenous mixture of solute and solvent.
(h). A solution can be unsaturated, saturated or supersaturated.
(i). An emulsion is a mixture of liquids that do not mix well
(j). A suspension is a heterogeneous mixture of liquid and fine particles of a solid in which the solid does not dissolve but is suspended in the liquid.

REVIEW QUESTIONS

1. Write ***TRUE*** or ***FALSE*** for the following statements.
   i. All known elements have names ....
   ii. An element may be represented by a symbol that is derived from the first letter of its English name.............................
   iii. Hydrogen, oxygen and nitrogen are examples of compounds......................
   iv. A heterogeneous mixture has uniform composition...........................
   v. A saturated solution is one that can still dissolve more solute at a given temperature..............................
2. Match items in ***LIST A*** with their corresponding statements in ***LIST B***.

List A	LIst B
i. Decantation	A. Used to separate a heterogenous mixture of solid and a liquid
ii. Evaporation	B. separate mixtures using a moving solvent on material that absorbs the solvents
iii. Distillation	C. used to separate mixtures by heating a liquid to very high temperature until it vaporizes
iv. Funnel separation	D. separates a solute from a liquid solution
v. Sublimation	E. Separates immiscible liquids using a separating funnel.
	F. Separates a heterogenous mixture of a liquid and a solid or solids by pouring out of the liquid only and leaving the solid at the bottom of the container.
	G. A solid change state directly to gas, usually on heating.

3. Choose the most correct answer for the following questions.

i. The substance that is separated during chromatography is called;
   A. Analysis   B. Precipitate   C. Analyte   D. Stationary phase.

ii. Chromatography can be used in hospitals to:
   A. Detect types of drugs in blood   B. Analyze blood in crime scenes
   C. Testing purity of machines   D. Inject sick people.

iii. One of the following is NOT involved in extracting oil from groundnuts
   A. Petroleum   B. Ethanol
   C. Groundnuts   D. Fractional distillation apparatus.

iv. The solid obtained during filtration is called_____ and the liquid is called _____ respectively.
   A. Filtrate, residue   B. Residue, analyte
   C. Analyte, residue   D. Residue, filtrate

4. Butter is water in oil emulsion. What is meant by water in oil emulsion?
5. If you were stranded in a desert with no water, how could you collect a supply of drinkable water? You only have a spade, a plastic sheet, a few heavy stones and a plastic cup.
6. Mention five elements that derive their symbols from their Latin names.
7. Name the process that can be used to separate the following substances:
   a. Kerosene and water
   b. Iodine and sand
   c. A mixture of petrol and diesel
   d. A mixture of sulphur and iron filings.
   e. Oil from groundnuts.
8. Here are some different mixtures:
   Shaving cream, hair lotion
   Instant coffee, margarine.
   Copy and complete the table below:

Type of mixture	Contains	Example
Solution	One substance dissolved in another	
Foam	Gas mixed with a liquid	
Emulsion	Two liquids which do not mix	
Gel	Liquid mixed with a solid.	

9. Briefly explain how you can tell the difference between a solution and a suspension just by their appearance.

Chapter 7

Air, Combustion, Rusting and Fire fighting

7.1 COMPOSITION OF AIR.

Air is a colourless mixture of gases. It is supposed to be odourless and tasteless. The atmosphere which covers the earth is made of air. More air is near the surface of the earth than away from it. Near the surface of the earth, the air is denser than it is higher up. Air is composed of several gases like nitrogen, oxygen, carbondioxide, water vapor, noble gases and methane.

Air may sometimes contain traces of impurities such as carbon monoxide (CO), sulphur dioxide (SO_2), hydrogen sulphide (H_2S) and other gases. Consider the below table which shows compostion of air by volume.

Table 7.1 Composition of air by volume.

Gas	Symbol	Percentage by volume	Boiling point (˚C)
Nitrogen	N_2	78.0	-196
Oxygen	O_2	21.0	-183
Argon	Ar	0.9	-186
Carbon dioxide	CO_2	0.03	-57
Neon	Ne	0.002	-246
Methane	CH_4	0.0002	-164
Helium	He	0.0005	-269
Krypton	Kr	0.0001	-153
Hydrogen	H_2	0.00005	-252.9
Xenon	Xe	0.00001	-108

TESTS FOR GASES IN AIR.

The test for the presence of different gases in air can be carried out using several procedures.

Activity 7.1: To determine the presence and percentage of oxygen in air.
Procedure.

1. Clean two syringes of more than 150 cm^3
2. Make very small cuttings of copper foil using a pair of scissors. The copper foil should be cleaned first with sand paper until it shines.
3. Place the copper cuttings at the centre of a hard glass silica tube.
4. Join the silica tube ends to two syringes as shown in figure 7.1
5. At the beginning, one syringe A should be closed while the other B is filled with 100 cm^3 of air.
6. Heat the silica tube strongly while you push air from one syringe to another. The air passes over the copper, which uses up the oxygen in air to make copper oxide. Copper + Oxygen → copper oxide.
7. Allow the tube to cool and read the volume on one syringe while the other is closed.
8. Repeat the heating and cooling until the volume of air does not change any more.
9. Summarize the results as illustrated in table 7.2.

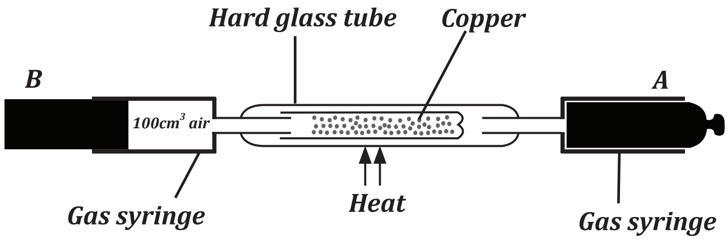


Fig 7.1 finding percentage of oxygen in air.

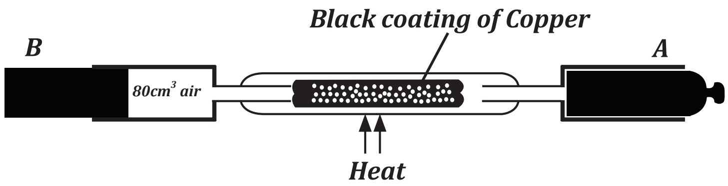


Fig 7.2, Reaction of copper with oxygen in air sample.

Table 7.2 Summary of results of activity 7.1

Volume of air in the syringe	Actual reading.
Before heating	100cm^3
After first heating and cooling	
After second heating and cooling	
After third heating and cooling	

Discussion questions.

1. What volume of air was used up by copper?
2. Why was the heating and cooling repeated three times?
3. What is the percentage of oxygen in air?
4. Why was it necessary to cool the apparatus before reading the volume?
5. Why was it necessary to push the air from one syringe to the other as heating of the copper continued?

Activity 7.2 To show that air contains carbon dioxide and water vapour.

Procedure.

1. Grind some blue crystals of copper (II) Sulphate with a pestle in a mortar to a very fine powder.
2. Heat the powder strongly while stirring until the entire blue colour is lost. You should remain with a white powder. Put this powder in test tube A.
3. Put lime water in test tube B.
4. Set the apparatus as seen in figure 7.3
5. Put ice cubes in a beaker around test tube A. Add common salt on the ice cubes. Add some water in the beaker.
6. Pass air through the system by using a filter pump or an aspirator.
7. Observe any color changes in test tube A and test tube B.
8. Stop the reaction when you think there is no more color change.

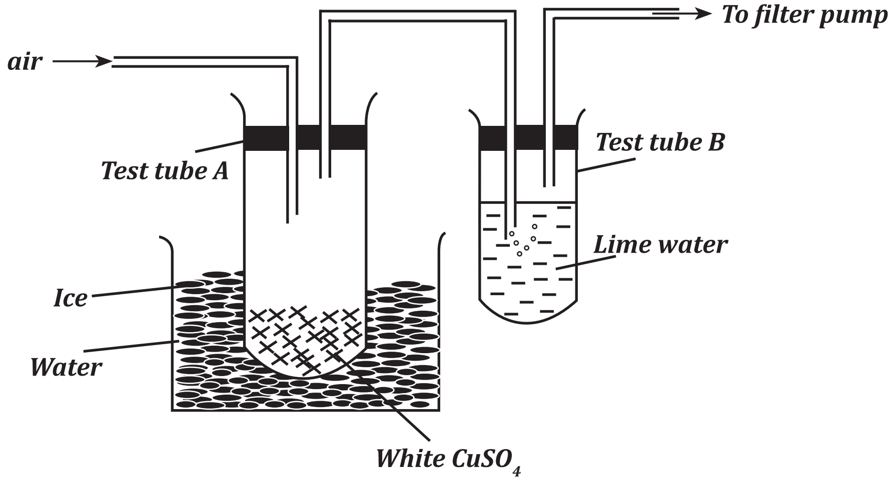


Fig. 7.3 showing that air contains carbon dioxide and water vapor.

DISCUSSION QUESTIONS.

1. What colour change did you observe on white powder in test tube A?
2. What caused the color change in test tube A?
3. Where did the dew in test tube A come from?
4. Why was ice water mixed with salt?
5. What colour change did you observe in test tube B?
6. What caused the colour change in test tube B?

SEPARATION OF AIR INTO ITS CONSTITUENT GASES.

The task of separating air into its constituent elements is not simple. It cannot be done in laboratory. It is done in the industry. The chemical industry needs the gases from the air in their pure form. The separation is carried out by fractional distillation of liquid air. Air is liquefied by subjecting it to ***high pressure*** and ***low temperature.***

There are three main stages in the process of liquefaction of air.

a. Carbon dioxide and water vapour must be removed first. By cooling the air to very low temperatures, ice and solid carbon dioxide are removed.
b. The remaining air is then compressed to about 150 times the atmospheric pressure. As the compressed air gets very hot, it has to be cooled.
c. The compressed cooled air is allowed to expand rapidly. The rapid expansion cools the air to very low temperature, and liquid-air drops out. At -200˚C, only helium and Neon remain as gases. These cold gases are used to cool the compressed air.

The liquid – air is then subjected to fractional distillation. The liquid air is allowed to warm up. Nitrogen boils off first because it has a lower boiling point, -196˚C. Oxygen follows by boiling at -183˚C.

AIR AS A MIXTURE.

Air is a mixture and not a compound due to the following reasons.

1. The composition of air does not correspond to any simple chemical formula. Any compound must have a specific formula.
2. The composition of the air varies slightly in different localities. In the same locality the composition is different at different times. If it were a compound, it would always have a definite composition by mass everywhere and all the time.
3. When the constituents of the air are mixed in their right proportions, there is no energy (heat or light) change. A compound is always made by a chemical change involving energy changes.
4. The gases in air can be separated by purely physical means. (i.e. fractional distillation.) Elements of a compound can never be separated by physical means.

USES OF THE GASES IN AIR.

1. ***Nitrogen.***

This gas is present in air in the largest proportion. It is not directly available to plants, or animals. The gas is not essential for breathing, although its presence in the air helps to dilute the oxygen. Nitrogen in plants and animals is obtained by indirect methods e.g. nitrogen fixation. Many of the fertilizers are products of reaction between nitrogen and other elements.

2. ***Oxygen.***

All plants and animals need oxygen for breathing and respiration. Aquatic animals use the oxygen dissolved in water. Ocean divers and mountain climbers must carry their own oxygen in tanks. Oxygen is necessary for burning of fuels and rusting. Green plants supply oxygen to the air during photosynthesis.

3. ***Carbon dioxide.***

Respiration in plants and animals produce carbon dioxide. All fires from burning fuel produce the gas. Automobiles and all other fuel engines produce carbon dioxide in their exhaust. Green plants take carbon dioxide from the atmosphere during photosynthesis. They use it to make food (sugars and starch). The gas is used extensively in fire extinguishers.

4. ***Noble gases.***

Argon, helium, krypton, neon and xenon comprise a "family" or group of elements present in air known as noble gases.

Noble gases are never prepared in the laboratory. They are obtained by fractional distillation of liquid – air. Helium is obtained from natural gases. It is used to fill certain photographic flash bulbs.

Argon is used for filling electric light bulbs. It is the cheapest and most abundant of the noble gases. If air is used instead, the filament may burn out.

Neon is extensively used for filling electric discharge tubes for illumination and for advertising signs.

Helium is second to hydrogen as the lightest gas known. As the gas for filling balloons, it has replaced hydrogen.

7.2 COMBUSTION.

Combustion is a chemical reaction that involves the burning of a substance in the presence of oxygen. Combustion is accompanied by liberation of heat and light. In this sense, combustion has the same meaning as burning. In a complete combustion reaction, a compound reacts with an oxidizing element such as oxygen or fluorine, and the products are compounds of each element in the fuel with the oxidizing element.

Example.

$$\underset{\text{Methane}}{CH_{4(g)}} + \underset{\text{Oxygen}}{2O_{2(g)}} \longrightarrow \underset{\text{Carbondioxide}}{CO_{2(g)}} + \underset{\text{Water vapour}}{2H_2O_{(g)}} + \text{energy}$$

A simple example can be seen in the combustion of hydrogen and oxyen which is a commonly used reaction in rocket engines whereby the result is water vapour.

$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(g)} + \text{heat}$$

Combustion does not necessarily involve oxygen. Sodium burns in chlorine to form sodium chloride. In combustion, one of the reactants must be a gas. Ignition is the process of starting the combustion of fuel. Ignition is very applicable in the cylinders of an internal – combustion engine.

Experiment 7.1

Aim: To find the products of combustion of a candle in air.

Materials: Mortar and pestle, candle, two test tubes, funnel, delivery tube, rubber, beaker, corks.

Other materials: blue copper (II) Sulphate, ice water, lime water.

Procedure.

(i). Grind blue copper (II) Sulphate with a pestle in a mortar to get a fine powder. Heat this powder very strongly until it turns completely white.

(ii). Put this powder in the test tube A.

(iii). Connect a delivery tube to a funnel using a rubber connection.

(iv). Keep test tube A in a beaker of ice water

(v). Put lime water in test tube B.

(vi). Set the apparatus as seen in figure 7.4

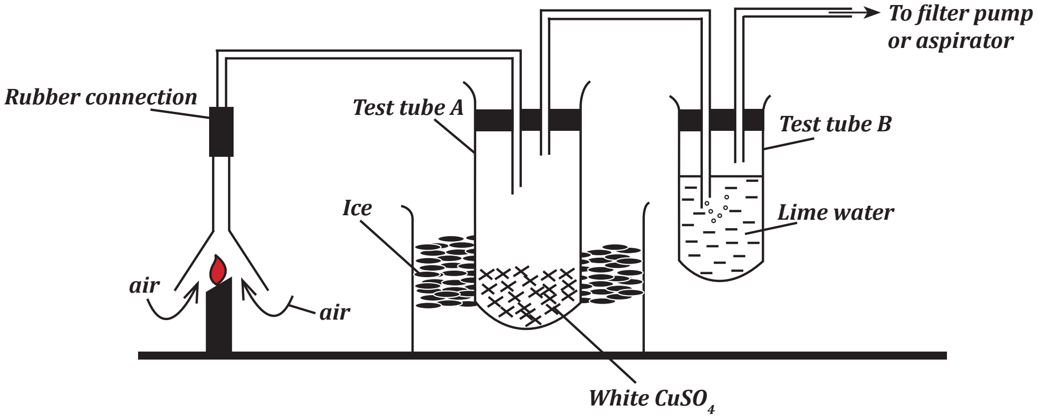


Figure 7.4. Investigating the products of the combustion of a candle

When the air has passed through the apparatus;

- The white anhydrous copper (II) Sulphate turns blue. This shows that water was formed when the candle burned.
- The lime water turns milky. This shows that carbon dioxide was formed when candle burned. When substances burn they actually gain in mass.

APPLICATION OF COMBUSTION IN REAL LIFE.

Combustion is applied in many areas such as ;

1. ***Industries.*** In engines, in large boilers, incinerators for burning wastes, during welding and smelting, Ovens.
2. ***Domestic activities.*** Cooking, heating homes, burning wastes, lighting lamps.
3. ***Laboratory.*** Sterilization, during experiments, lighting burners.

7.3 FIRE FIGHTING

Fire fighting is the extinguishing of harmful fires. Fire is the state of combustion in which ignited material combines with oxygen and gives off light, heat and flame. Fires can be dangerous. The most dangerous fires are those which are caused by explosions. Explosions may shatter the whole building and kill many people.

HOW TO PREVENT FIRES.

1. Extinguish all glowing substances before you dispose them.
2. Use the fuel which is appropriate for a particular burner. Petrol should not be used in a kerosene burner.
3. Cool a burner completely before refilling it with a liquid fuel.
4. Flammable chemicals should be locked out of the working laboratory.
5. After taking a flammable liquid from its container for an experiment, remove the container from the laboratory before you start the experiment.
6. Containers of flammable liquids should be well stoppered always.
7. Never heat flammable liquids with direct flames.
8. Never pour flammable liquids into the sink.

HOW TO CONTROL FIRE.

A fire will start or continue to burn if three factors are satisfied.

1. There should be a combustible material. Examples of materials are solid, liquid and gaseous fuels. These substances can catch fire and burn. You can extinguish the fire by removing the combustible materials.
2. There should be a supply of oxygen. Fuels will only burn if there is enough oxygen. You can extinguish the fire by blocking air or oxygen supply to the fire.

3. The temperature should be at the kindling point of the fuel or above it. Every fuel has its own kindling point. Below the kindling point, the fuel will not catch fire. You can extinguish the fire by lowering the temperature below the kindling point.

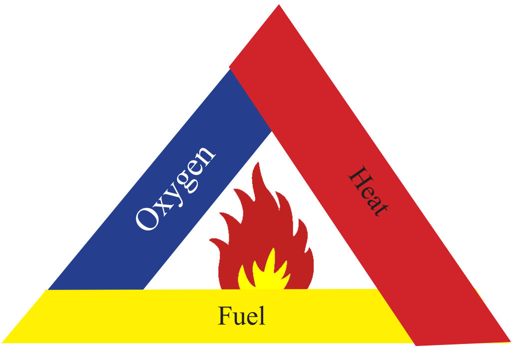


Fig.7.5 fire triangle

CLASSIFICATION OF FIRES.

Classification of fires depends on what is actually burning.

(a) Class A fires. The burning materials are solids such as wood, coal, paper, rubber, plastics, clothes, furniture. Water is the best extinguisher of these fires. If water is not available use any other type of fire extinguisher except carbon dioxide.

(b) Class B fire. The burning material is a flammable liquid e.g. petrol, kerosene, alcohol, ether, varnishes, fats, etc. If the fire is small a fire blanket or sand may be used. If the fire is large, use foam or carbon dioxide extinguishers. You should never use water on class B fires.

(c) Class C fires. The burning material is a liquefied gas. Example acetylene, hydrogen or coal gas. Foam extinguishers and carbon dioxide extinguishers are the best. It is very important to turn off the source of gas. Spray water on the gas tank to cool it.

(d) Class D fires. The burning material is a metal. Alkali metals e.g. sodium or potassium may catch fire when they come in contact with water and oxygen. At high temperatures many metals react with oxygen vigorously. Dry powder extinguisher and foam extinguisher are the best in class D fires.

(e) Class E fires. These are fires on electrical equipment. You must first find the main switch and put it off. You may use carbon dioxide extinguisher to put off the fire. You should never use water on electric fires. Water increases electrical conductivity.

(f) Class F fires. The burning materials are cooking appliances with oils and fats at high temperatures. Use wet chemical extinguishers to handle such fires.

TYPES OF PORTABLE FIRE EXTINGUISHERS.

A portable fire extinguisher is one that can be easily moved from one place to another. It is usually hung in an upright position on a wall. It consists of a metal container that contains the extinguishing substance stored at high pressure. Fire extinguishers are classified according to the type of chemicals they contain.

(a) Liquid carbon dioxide extinguisher.

This is an extinguisher which uses liquid carbon dioxide. The liquid is contained in a metal container. When the safety pin is removed, carbon dioxide evaporates as solid "Snow", (carbon dioxide sublimes). The snow settles on the fire and smothers it.

(b) Soda – Acid fire extinguisher

This is an extinguisher which uses concentrated sulphuric acid and sodium hydrogen carbonate solution. The solution is contained in a metal container with a metal knob at the top of the acid bottle. When it is hot, the knob breaks the bottle, the acid and/soda solution react to form carbondioxide. The metal case is turned upside down. The gas mixes with the liquid and the froth (bubbles of carbon dioxide) comes out of the jet. The froth covers the burning substances to keep air from reaching it.

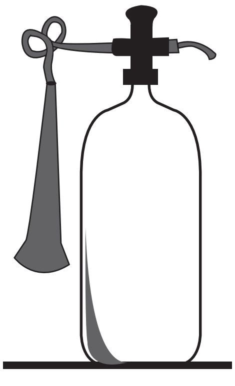
(c) Foam type fire extinguisher.

This extinguisher also contains a solution of sodium hydrogen carbonate mixed with aluminium sulphate and saponin. Saponin is a substance which forms bubbles very readily. When the three substances mix, they form a foam of carbodioxide bubbles. The foam also keeps air away from the burning material.

(d) Powder – type fire extinguisher.

This extinguisher uses powdered sodium hydrogen carbonate, and a gas kept at high pressure. When the gas cartridge is broken using the top cap, the carbon dioxide under pressure propels the powders. The powder forms a layer over the burning materials to keep air away.

Fig 7.6 Liquid carbondioxide extinguisher



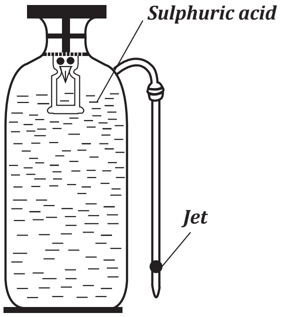


Fig 7.7, soda -acid fire extinguisher

CHEMICAL COMPOSITION OF DIFFERENT AGENTS IN PORTABLE FIRE EXTINGUISHERS

1. Air pressurized water (APW)

This consists of ordinary tap water pressurized with air. It is suitable for class A fires and unsuitable for Class B, C and D fires.

2. Dry chemical (DC).

Consist of fine sodium bicarbonate powder pressurized by nitrogen. It is suitable for class A, B, C and E fires. It is unsuitable for class D fires. When used indoors it can obscure vision.

3. Carbon dioxide (CO_2).

It consist of non – flammable carbon dioxide gas under extreme pressure. Suitable for class B, C and E fires and unsuitable for class A fires.

4. Halon.

Consist of Bromochloro – difluoro – Methane. It is suitable for class A and E fires and unsuitable for class B and C fires.

5. Foam.

Consist of proteins and fluoro – proteins. Suitable for class A and B fires and unsuitable for class E fires.

6. Wet chemical.

It consist of potassium acetate. Suitable for class F fires and unsuitable for class E fires.

7. ABC.

Consists of mono- ammonium phosphate with a nitrogen carrier. Suitable for class A, B and C fires and unsuitable for electronic equipment.

USING A PORTABLE FIRE EXTINGUISHER.

Use the PASS method when using a portable fire extinguisher.

- P = Pull the pin
- A = Aim hose at the base of the fire.
- S= Squeeze the handle
- S= Sweep back and forth with the extinguisher.

Remember these fire fighting tips.

1. Most of the fire extinguishers are emptied in less than a minute
2. Do not attempt to fight a large fire
3. Always call the fire department when a fire breaks out.
4. Always leave yourself a way out – keep your back to an exit.
5. Keep a reasonable distance from the fire as it may suddenly change direction.

6. Never use a portable extinguisher on people instead use a fire blanket.
7. Do not test a portable extinguisher to see if it works. It may leak and afterwards fail to work during an emergency.
8. Do not return a used portable extinguisher to the wall.

Activity. 7.3:
Visit the nearest fire station. (Fire brigade). Ask the officials to tell you:

a. How to establish fire protection at your home
b. What to do in case of a large fire in your school laboratory.
c. How to rescue a friend who is caught in a fire.
d. The best way to fight fires using sand buckets, a water hose and an asbestos blanket.

7.4 RUSTING.

Rusting is a chemical process that occurs in iron or steel after reaction with water and oxygen from the air. Rust is the reddish – brown coating on the surface of iron or steel.

Rusting causes a lot of damage economically. The rusting of bridges, corrugated iron sheets on roofs, containers, rails, etc causes economic problems. If a ship is not protected from rusting, it will not last long in the ocean. All efforts must be made to stop iron or steel items from rusting. This can be done if we know the conditions which causes or favour rusting.

Rusting and burning have something in common; they both require oxygen. During burning, magnesium combines with oxygen of the air, forming magnesium oxide.

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

During rusting, iron combines with oxygen of the air in the presence of water to form brown hydrated iron (III) oxide, ‘rust’

$$4Fe_{(s)} + 3O_{2(g)} + 2xH_2O_{(l)} \longrightarrow 2Fe_2O_3 \bullet xH_2O_{(s)}$$

The only difference between rusting and combustion is in the time required for the two processes. Heat is generated during rusting just as it is during burning, but it is dissipated without attracting notice because of its much slower rate of production.

CONDITIONS FOR RUSTING.

There are several conditions that are necessary for rusting to take place.

Experiment. 7.2

Aim: To find the conditions needed for iron to rust.

Procedure.

1. Clean four test tubes and label them W,X,Y and Z. Put them in a test tube rack.
2. Use sand paper to clean 8 new iron nails. Put two of the nails in each test tube.
3. Add dry anhydrous calcium chloride in test tube W until the nails are completely covered. Stopper the test tube using a rubber bung.
4. Cover the nails in test tube X with a wet cotton wool. Cover the mouth of the test tube with a dry cotton wool.
5. Boil water in a beaker until all air in it is driven off. Pour this water into test tube Y, until the nails are submerged in the water. Pour oil on top of the boiled water.
6. Boil some oil until you are sure it contains no water or air.
7. Pour the boiled oil in the test tube Z until the nails are submerged in the oil.
8. Keep the set up aside for two days, after which you observe what will have happened to the nails in the four test tubes.

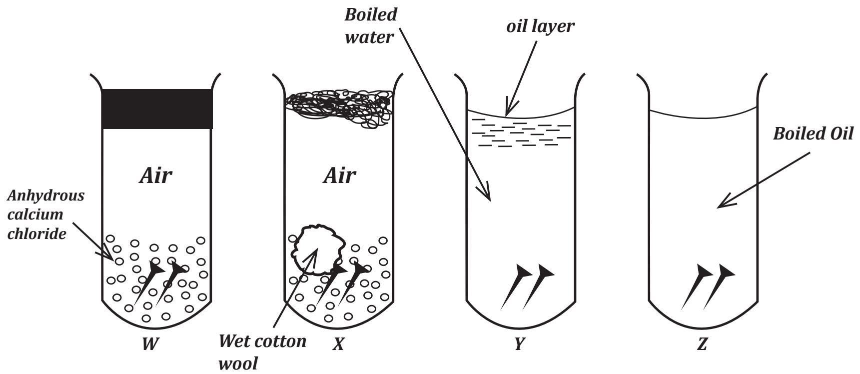


Fig 7.8 Conditions for Iron rusting.

DISCUSSION QUESTIONS.

1. Why was it necessary to clean the iron nails with sand paper?
2. What was the use of the anhydrous calcium chloride in test tube W? What other substance can be used instead of calcium chloride for the same purpose?
3. What was the use of the oil layer in test tube Y?
4. Why was the water in test tube Y boiled first?
5. Why was the oil in test tube Z boiled?

6. In which test tube was air present but water absent?
7. In which test tube was water present but air absent?
8. In which test tube was both air and water absent ?
9. In which test tube was both air and moisture present?
10. In which test tube did rusting occur most?
11. In which test tubes did rusting not occur?

Within two days, the nails in test tube W will not have formed rust. This is due to the fact that anhydrous calcium chloride absorbs moisture and also the test tube was stoppered to prevent entrance of air.

The nails in test tube X will undergo rusting because of the presence of moisture and air.

Nails in test tube Y will not undergo rusting because there is no dissolved oxygen in boiled water. Also the presence of oil prevents oxygen from dissolving in the test tube.

Nails in test tube Z will not undergo rusting because of the absence of both water and oxygen.

METHODS OF PREVENTING RUSTING.

To stop iron and steel from rusting we must protect them from water (moisture) and oxygen (air). Some of the methods used to prevent rusting are mentioned below.

1. Painting.

An oil base paint is applied on the surface of ships, frills, cars, machines, roofs etc. The layer of paint keeps oxygen and water away.

***2. Oiling*.**

Moving parts should not be painted. The layer of paint will be easily scratched off. Instead they are oiled or greased. The greasing also helps lubrications.

3. Alloying.

It is possible to mix iron and steel with other metals to form alloys. Stainless steel is an alloy which does not form rust. The stainless steel contains chromium, nickel and manganese mixed with iron.

4. Galvanization.

This is the process of coating iron or steel with zinc, a metal that does not rust. Zinc is inexpensive and therefore economical to use.

5. Anodizing.

This is a method used to protect large structures such bridges, ships and pipelines from rusting. The method is based on a process known as sacrificial protection. The iron or steel structure to be protected from rusting is joined to a reactive metal (e.g. magnesium) by a wire. The magnesium is corroded instead of the iron or steel, hence the term sacrificial protection. The magnesium metal becomes the anode and the structure becomes the cathode; hence anodizing. The anode is sacrificially corroded while the cathode is protected.

6. Use of silica gel.

Silica gel is a compound that absorbs water and thus renders its surroundings dry. It is a granular form of silicon encased in plastic. Some fragile instruments such as cameras have iron and steel parts. These cannot be protected by painting, oiling, galvanizing etc. Most often, a small bag of silica gel is put inside the bag carrying the instrument to absorb moisture in order to avoid rusting.

7. Tin plating.

Tin is a useful metal for the food processing industry since it in non – toxic, ductile and corrosion resistant. Tin plating is the coating of iron with tin, Tin cans are made of steel, but the inside of the can is coated with a thin layer of tin. This is why they are suitable for canning of foods.

8. Use of plastic.

Plastic can be used to make parts of some machines or instruments. These instruments need not to be made by using iron or steel. By using plastic, rusting is prevented and the resultant product is less expensive.

9. Electroplating.

Electroplating is a process that uses electrical current to reduce dissolved metal cations so they form a coherent metal coating on an electrode. Electroplating can also prevent rusting. Articles like car bumpers, taps and kettles are chromium plated. Often these articles are nickel plated first to give protection against rusting and finally given a very thin coating of chromium. This give a very shiny surface.

SUMMARY

(a). Air is a colourless mixture of gases.
(b). Air is composed of several gases like nitrogen, oxygen, carbon dioxide, water vapour, noble gases and methane.
(c). Separation of air into its constituent gases is carried out by fractional distillation of liquid air.
(d). Air is a mixture and not a compound
(e). Combustion is a chemical reaction that involves the burning of a substance in the presence of oxygen.
(f). Combustion can be applied in industries, domestic activities and in laboratories.
(g). Fire fighting is the extinguishing of harmful fires.
(h). Fire is the state of combustion in which ignited material combines with oxygen and gives off light, heat and flame.
(i). Classification of fire depends on what is actually burning.
(j). Fire triangle consists of oxygen, heat and fuel.
(k). Rusting is a chemical reaction that occurs in iron or steel after reaction with water and oxygen from the air.

REVIEW QUESTIONS.

1. Copy and complete the table below. One of the items has been done for you.

Class	Burning material	Appropriate extinguisher
A	Ordinary solid combustible material such as paper, wood and clothing.	Water, any type of portable extinguisher except carbon dioxide.
B		
C		
D		
E		
F		

2. Choose the best answer to complete each statement or to answer each statement or to answer each question.

(i) …………………… is a colourless mixture of gases.
A Hydrogen B Oxygen C. Air D. Helium

(ii) Air may sometimes contain traces of …………… such as carbon dioxide and sulphur dioxide.
A. Impurities B. Purities
C. Pure substances D. Noble gases.

(iii) The chemical reaction that involves the burning of a substance in the presence of oxygen is called……......................

A. Combination B. Synthesis

C. Combustion D. Reduction.

(iv) The state of combustion in which ignited material combines with oxygen and gives off light, heat and flame is called…......……...…..

A. Fire B. Fire fighting C. Ignition D. Reaction.

(v) ……….. is the extinguishing of harmful fires.

A. Fire prevention B. Fire protection

C. Fire rescuing D. Fire fighting.

3. Write TRUE for statement that is correct and FALSE for statements that are not correct.

i. A fuel is a substance that releases energy when burnt....……………
ii. Hydrogen, heat and fuel are components of fire triangle……..………
iii. A portable fire extinguisher is one that can be easily moved from one place to another…………….
iv. Portable fire extinguishers should not have ratings or coding on them……………..
v. The procedure for how to use a fire extinguisher is easy to remember using the acronym “PASS”…………………….
vi. Rusting causes a lot of damage economically………………..
vii. When using a fire extinguisher, do not return a used portable extinguisher to the wall…………………

4. Match the items in list A with the responses in List B by writing the letter of the correct response beside the item number.

List A	List B
i. Oiling	A. The coating of items using a special pigment paint
ii. Painting	B. The coating of iron with tin
iii. Galvanization	C. Coating iron with oil
iv. Tin plating	D. Used in case some parts of machines do not have to be made of steel or iron.
v. Silica gel	E. A chemical containing silicon found in sand and rocks.
vi. Use of plastic	F. Coating iron or steel with zinc.
vii. Electroplating	G. Sacrificial protection.
viii. Anodizing	H. Nickel plating and finally coating with a thin layer of chromium
ix. Alloying	I. Mixing iron with rust

5. Give the difference between rust and rusting.
6. What do you understand by ignition?
7. Copy and complete the table below. One of the items have been done for you.

Gas	Symbol	Percentage by Volume
Nitrogen		
Oxygen		
Neon		
Argon		
Carbondioxide		
Methane		
Helium		
Krypton		
Hydrogen		
Xenon	Xe	0.00001%

8. With the aid of a chemical equation explain the process of rusting.
9. Explain the similarities and differences between burning and rusting.
10. Mention three components of a fire triangle.

Chapter 8

Oxygen

8.1 CONCEPT OF OXYGEN.

Oxygen is a gas that forms about 21% by volume of the air. Oxygen derives its name from two Greek words; ***"oxys"*** which means ***acid*** and ***"gonos"*** which means ***producer***. At the time of naming, it was mistakenly thought that all acids required oxygen in their composition.

Oxygen combines with non – metals to form non – metallic oxides and with metals to form metallic oxides.

PREPARATION OF OXYGEN.

Oxygen can be prepared in a laboratory or at industrial scale. Some of the methods used to prepare oxygen are;

1. Fractional distillation of air.
2. Heating potassium chlorate
3. The action of manganese (IV) oxide on hydrogen peroxide.
4. Heating mercuric oxide.
5. Heating potassium permanganate.

8.2 LABORATORY PREPARATION OF OXYGEN.

The common methods of preparation of oxygen in the laboratory include the following.

1. Decomposition of hydrogen peroxide.
2. Heating of potassium chlorate in the presence of a catalyst
3. Heating compounds rich in oxygen.

The most commonly used methods for laboratory preparation of oxygen is by using hydrogen peroxide and potassium chlorate. In each case manganese (IV) oxide is used as a catalyst. With potassium chlorate, heat is needed to activate the decomposition. No heat is required when hydrogen peroxide is used to prepare oxygen.

Experiment 8.1

Aim: laboratory preparation of oxygen by the action of manganese (IV) oxide on hydrogen peroxide.

Apparatus: beehive shelf, flat – bottomed flask, delivery tube, trough, thistle funnel, gas jar, two holed rubber bung.

Chemicals: Hydrogen peroxide solution, manganese (IV) oxide and water.

Procedure

1. Arrange the apparatus as shown in figure 8.1. Make sure that the beehive shelf is completely immersed in water.
2. Put a small quantity of manganese (IV) oxide into a flat bottomed flask.
3. Fill the gas jar with water and invert it over a beehive shelf in a trough of water. Make sure the water in the gas jar has no bubbles.
4. Slowly run in hydrogen peroxide solution from the thistle funnel. The reaction is normally vigorous.
5. Allow the first bubbles of the gas to escape then collect several gas jars of oxygen and keep them well covered.
6. Open one of the jars and smell some of the gas.

Discussion questions

1. Why is oxygen collected over water?
2. Why are the first few bubbles of gas not collected?

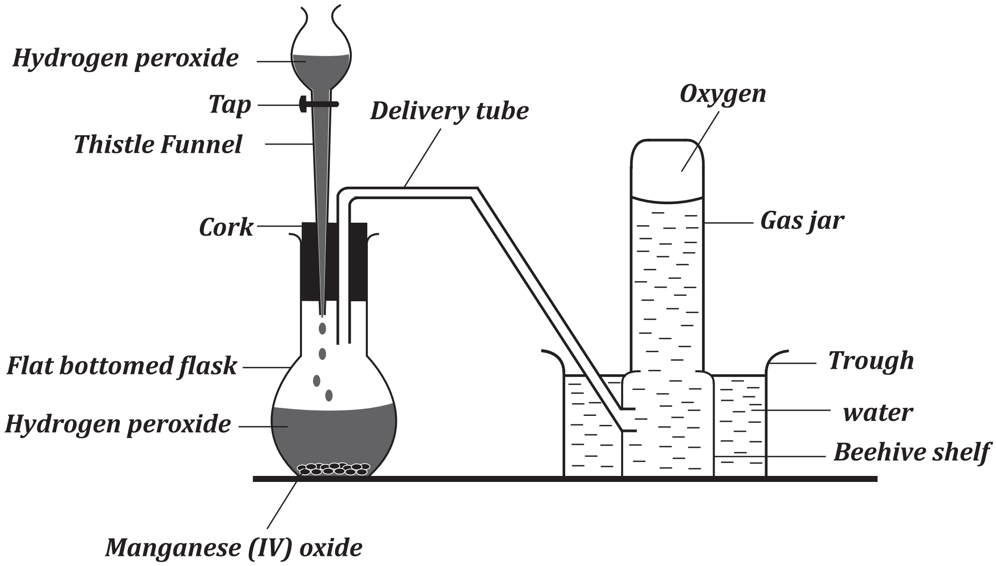


Fig 8.1 Laboratory preparation of oxygen

Hydrogen peroxide decomposes slowly at room temperature but the reaction is speeded up by a catalyst known as manganese (IV) oxide.

Word equation

Hydrogen peroxide $\xrightarrow{MnO_2}$ water + Oxygen

Chemical Equation

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

A catalyst alters the rate of a chemical reaction without itself undergoing any chemical change. It is possible to collect oxygen by downward displacement of water because it is only slightly soluble in water.

Experiment 8.2

Aim: Laboratory preparation of oxygen from potassium chlorate.

Apparatus;

Trough, beehive shelf, gas jar, delivery tube, retort stand, pyrex test tube and Bunsen burner.

Chemicals;

Water, potassium chlorate and manganese (IV) oxide.

Procedure

1. Crush some potassium chlorate crystals in a mortar and grind them by pestle to get fine powders.
2. Place the mixture of potassium chlorate and manganese (IV) oxide in a pyrex test tube.
3. Fit the apparatus as shown in figure 8.2 and make sure that the beehive shelf is completely immersed in water.
4. Heat the mixture strongly, moving the flame to and fro.
5. A gas is readily given off and can be collected over water. Potassium chloride is left in the pyrex test tube.

Manganese (IV) oxide (MnO_2) is used as a catalyst as in previous preparation. The catalyst makes potassium chlorate break down quickly at lower temperature and thus speed up the rate of a chemical reaction.

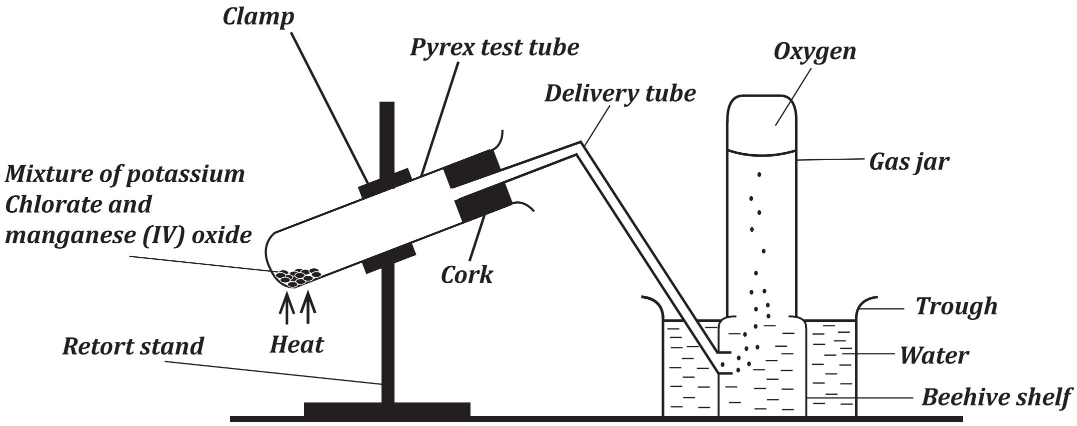


Fig.8.2 Laboratory preparation of oxygen from potassium chlorate

Word equation

$$\text{Potassium chlorate} \xrightarrow{MnO_2} \text{Potassium chloride} + \text{Oxygen}$$

Chemical Equation

$$2KClO_{3(s)} \xrightarrow[\text{Heat}]{MnO_2} 2KCl_{(s)} + 3O_{2(g)}$$

TEST FOR OXYGEN.

Oxygen gas rekindles a wooden glowing splint. The glowing splint bursts into flame and burns vigorously when lowered in a gas jar with oxygen. Oxygen itself does not burn.

8.3 PROPERTIES OF OXYGEN.

Physical properties.

1. Oxygen gas is colourless, tasteless and odourless.
2. Oxygen is slightly soluble in water.
3. Oxygen is about 1.5 times denser than air.
4. The boiling point of oxygen is -183˚C.
5. The freezing point of oxygen is -218˚C.

Chemical properties.

1. Oxygen supports combustion. This is why oxygen relights a glowing splint.
2. Oxygen is a very strong oxidizing agent.

$$\underset{\textit{Methane}}{CH_{4(g)}} + \underset{\textit{Oxygen}}{2O_{2(g)}} \longrightarrow \underset{\textit{Carbondioxide}}{CO_{2(g)}} + \underset{\textit{water}}{2H_2O_{(l)}}$$

$$2H_2S_{(g)} + O_{2(g)} \longrightarrow 2H_2O_{(l)} + 2S_{(s)}$$

Hydrogen Sulphide ***Oxygen*** ***water*** ***Sulphur***

3. Oxygen reacts with metals to form basic oxides.

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

Magnesium ***Oxygen*** ***Magnesium oxide***

$$2Ca_{(s)} + O_{2(g)} \longrightarrow 2CaO_{(s)}$$

calcium ***Oxygen*** ***Calcium oxide***

4. Oxygen reacts with non - metals to form non metal oxides (acidic oxides)

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

Carbon ***Oxygen*** ***Carbon dioxide***

$$S_{(s)} + O_{2(g)} \longrightarrow SO_{2(g)}$$

Sulphur ***Oxygen*** ***Sulphur dioxide***

Experiment 8.3 To investigate properties of oxygen.

1. Label the gas jars of oxygen you have collected as H,I, J,K and L.
2. Examine the gas in jar for colour and smell. Waft the gas towards your nose and inhale.
3. Put a small candle on a combustion spoon. Light the candle and lower it into jar I. Observe the candle flame. See figure 8.3
4. Light a wooden splint. Let it burn for some time and then blow out the flame. It should keep glowing.
5. Lower the glowing splint into jar J (figure 8.3 (b)) does the glowing splint relight?
6. Clean a piece of magnesium ribbon using sand paper. Fix the ribbon on a combustion spoon as shown in figure 8.3 (c)
7. Pour a little distilled water into Jar K.
8. Light the free end of the magnesium ribbon and put it in jar K immediately. When the reaction is over, shake the gas jar well. Put red litmus paper in the jar and leave it there for five minutes. Record any colour change.
9. Put a little Sulphur into a combustion spoon.
10. Pour a little distilled water into jar L.
11. Set the Sulphur alight in the combustion spoon and put in jar L immediately. When the reaction is over, shake the gas jar well. Put blue litmus paper in the jar and leave it there for five minutes. Record any colour change.

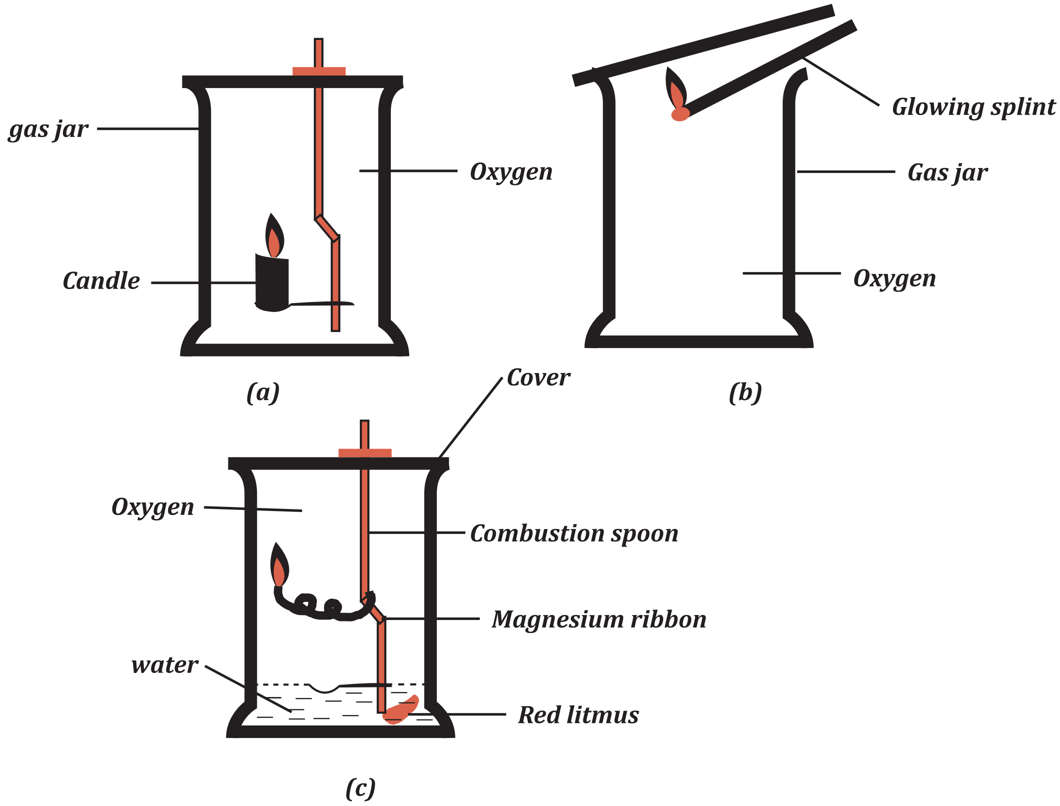


Fig. 8.3 Investigating some of properties of oxygen

Discussion questions

1. What is the colour of oxygen gas?
2. Does oxygen have any smell?
3. Was the candle flame brighter in the oxygen jar?
4. What happened to the glowing splint in the oxygen Jar J.
5. What was the color change to the litmus paper in Jar K? What does this show about the product formed when magnesium burns in oxygen?
6. What was the colour change to the litmus paper in Jar L? What does this show about the product formed when sulphur burns in oxygen?
7. What is a word equation for the reaction of oxygen with magnesium and the reaction of oxygen with sulphur respectively?

8.4 INDUSTRIAL MANUFACTURE OF OXYGEN.

The most commonly used method is the fractional distillation of liquefied air.

FRACTIONAL DISTILLATION.

Liequefaction of air; Air is liquefied by, first of all, being filtered to remove dust, and then cooled to -200˚C. At this temperature it is a liquid. As the air liquefies;

- The water vapour condenses and is removed using special filters.
- The carbon dioxide freezes at – 79˚C and is removed.
- The oxygen liquefies at - 183˚C
- The nitrogen liquefies at - 196˚C

Distillation:

Liquid nitrogen and oxygen mixture is then separated into the two components by fractional distillation. The liquid mixture is passed into a fractionating column from the bottom. Since the column is warmer at the bottom than at the top, the liquid nitrogen boils at the bottom of the column. The Gaseous nitrogen rises to the top where it is piped off and stored. The liquid oxygen collects at the bottom of the column.

8.5 USES OF OXYGEN.

Life on earth depends very much on the presence of oxygen. Animals and plants, both on land and in water need oxygen gas. Some important industrial processes need oxygen.

1. Oxygen in breathing and respiration. Oxygen is the only respiratory gas used by both plants and animals. In respiration, oxygen is used to burn food to produce energy, carbon dioxide and water vapour.

Oxygen is used in hospitals to people who are sick. Some diseases damage the lungs and stop them from taking enough oxygen from the air. Oxygen is given to people under anaesthetics and having operations. People who have almost drowned or have breathed poisonous gases are often given oxygen. Premature babies are kept in oxygen tents. Mountain climbers and airmen who fly very high use oxygen in tanks. Ocean divers use oxygen in tanks. Ocean divers use stored pure oxygen.

2. Oxygen in welding. Much oxygen is used in oxy-acetylene and oxy – hydrogen blow pipes. The oxy – acetylene flame is intensely hot (over 2000˚C)

3. Oxygen in purification of iron. Oxygen is blown into molten iron to remove impurities e.g. carbon or phosphorus. Such impurities are removed as gases.

4. Oxygen in space rockets. In outer space there is no oxygen. Rockets must carry their own supplies of liquid oxygen. Liquid oxygen is used to burn fuels in rockets. An oxygen hydrogen liquid mixture is the most known powerful propellant.

5. Oxygen in burning. Oxygen is essential for burning and enables vital combustion processes to take place. Without oxygen, we could not have fires or we could not be able to burn fuel in order to drive machines and cars.

Burning is a similar process to respiration. Their main difference is that burning take place faster while respiration is a slow process.

Respiration does not produce light energy. Also, burning can occur in any material while respiration occurs only in living cells organisms.

SUMMARY

(a). Oxygen is a chemical element with symbol O and atomic number 8 and it makes up about 21% by volume of air.

(b). The melting point of oxygen is -218˚C and boiling point is – 183˚C

(c). At standard temperature and pressure, oxygen is a colorless, odorless gas with the molecular formula O_2.

(d). The common allotrope of elemental oxygen on earth is called dioxygen, O_2.

(e). Trioxygen, O_3 is usually known as Ozone and is a very reactive allotrope of oxygen that is damaging to lung tissue.

(f). Naturally occurring Oxygen is composed of three stable Isotopes, ^{16}O ^{17}O and ^{18}O with ^{16}O being the most abundant (99.762% natural abundance).

(g). Industrial manufacture of Oxygen is mainly by fractional distillation of liquid air.

(h). Oxygen is prepared in the laboratory by mainly two methods; decomposition of hydrogen peroxide with Manganese (IV) oxide as a catalyst and heating potassium chlorate with manganese (IV) oxide as a catalyst.

REVIEW QUESTIONS.

1. Choose the most correct answer for the following questions.
   i. …………………….. is the most abundant element on the earth's outer crust.
   A. Hydrogen B. Oxygen
   C. Nitrogen D. Carbon dioxide.
   ii. ……………… alters the rate of a chemical reaction without itself undergoing any chemical change.
   A. Matter B. Element C. A catalyst D. Compound.
   iii. The chemical test for oxygen is…………………
   A. Oxygen does not support combustion
   B. Oxygen burns in air with a pop sound
   C. Oxygen reacts with hydrogen to form water.
   D. Oxygen gas rekindles a glowing splint.
   iv. …………… is the breaking down of a chemical compound into elements or smaller compounds.
   A. Composition B. Decomposition C. Dissolution D. Degradation.
   v. ………….. burn in oxygen to produce basic oxides while …………. burns in oxygen to produce acidic oxides.
   A. Metals, non metals. B. Non metals, metals.
   C. Metals, metals D. Non metals, non metals.

2. Match the items in list A with their correct statements in List B.

List A	List B
i. Physical propertis of oxygen.	A. ^{16}O, ^{17}O and ^{18}O
ii. chemical properties of oxygen.	B. ^{16}O and ^{15}O
iii. industrial manufacture of oxygen.	C. -183˚C
iv. the boiling point of oxygen.	D.180˚C
v. Isotopes of oxygen.	E. Oxygen support combustion and it is a very strong oxiding agent
	F. Oxygen gas is colourless, tasteless and odourless
	G. Fractional distillation of liquid air
	H. 100˚C

3. Write ***TRUE*** for a correct statement and ***FALSE*** for an incorrect statement.
   vi. Oxygen is an essential element for chemical processes like combustion, rusting and respiration………….

vii. It is not possible to collect oxygen by downward displacement of water because it is only slightly soluble in water…………..

viii. Basicoxidesreactwithwatertoformbasicsolutions………………..

ix. Acidic oxides react with water to form acidic solutions……………

x. Industrial manufacture of oxygen is mainly by heating compounds rich in oxygen……………..

4. Hydrogen peroxide can be a dangerous chemical. A bottle of hydrogen peroxide was left standing on a sunny window sill. The bottle had a screw top. Suddenly one day the bottle exploded.
   (a) Why did the bottle explode?
   (b) How would you suggest the hydrogen peroxide is stored safely?

5. Describe briefly four uses of oxygen.
6. The non – metal phosphorus burns in oxygen to form phosphorus (V) oxide.
   (a) Write a word equation for the reaction
   (b) What colour could you expect litmus paper to turn when phosphorus (V) oxide is tested? Explain your answer.
7. With the aid of well labeled diagram and a balanced chemical equation, outline the laboratory preparation of oxygen from hydrogen peroxide.
8. How does oxygen react with sulphur, charcoal, magnesium, calcium and sodium? Name specifically the products obtained in each reaction.

Chapter 9

Hydrogen

9.1 THE CONCEPT OF HYDROGEN

Hydrogen is the lightest and most abundant element in the universe and make up about 90% of the universe by weight.

Hydrogen is derived from two Greek words ***"hydro"*** which means ***water*** and ***"genes"*** which means ***generator***. Hydrogen was discovered by Henry Cavendish in 1766 in London. Cavendish described it as "inflammable air from metals" because he collected it over mercury.

PREPARATION OF HYDROGEN

Hydrogen gas can be prepared industrially or in a laboratory. Among others, methods used to prepare hydrogen gas are:

1. Electrolysis of water.
2. Thermal decomposition of ammonia
3. The reaction of water with certain metals
4. The reaction of water with hot carbon.
5. The reaction of dilute acids with moderately reactive metals

9.2 LABORATORY PREPARATION OF HYDROGEN

The common method for preparing hydrogen in the laboratory is by the action of moderately reactive metal on a dilute acid. The common moderately reactive metals are zinc, iron and magnesium. The common choice is zinc metal and dilute hydrochloric acid or dilute sulphuric acid. Zinc is preferred as it reacts smoothly with either hydrochloric acid or sulphuric acid.

Experiment 9.1

Aim; laboratory preparation of hydrogen.

Apparatus; Thistle funnel, gas jar, trough, beehive shelf, two-holed rubber bang, flat-bottomed flask.

Chemicals; Dilute hydrochloric acid, Zinc granules, water.

Procedure;

1. Arrange the apparatus as shown in figure 9.1
2. Put some pieces of zinc granules into a flat-bottomed flask.
3. Fill a gas jar with water and invert it over the beehive stand in the trough.
4. Add dilute hydrochloric acid through the thistle funnel until the zinc is covered with the acid and the end of the thistle funnel is immersed in the acid.
5. Collect the gas over water, ensuring you only remove the gas jar when it is full.
6. Collect several jars of the hydrogen gas.

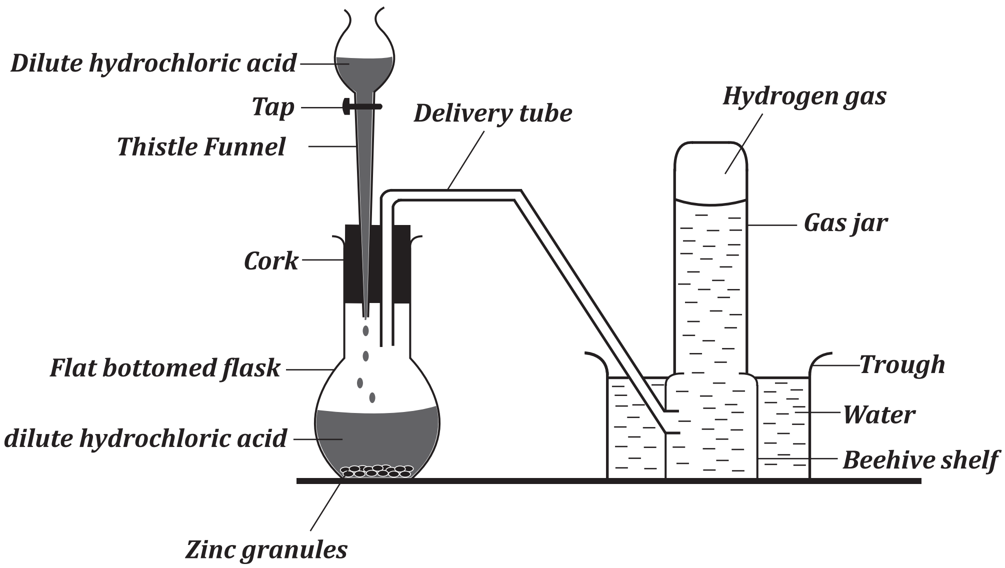


Fig 9.1 laboratory preparation of hydrogen gas

Discussion questions

1. Why were the few bubbles of gas not collected?
2. What is the colour of the gas?
3. Why is it possible to collect the gas using this method?
4. Suggest one problem with collecting hydrogen this way.

Word equation

Zinc + Hydrochloric acid $\longrightarrow$ Zinc Chloride + Hydrogen gas

Balanced chemical equation

$$Zn_{(s)} + 2HCl_{(aq)} \longrightarrow ZnCl_{2(aq)} + H_{2(g)}$$

Hydrogen can be collected over water because it does not dissolve in water. It can also be collected by upward delivery (downward displacement of air) because hydrogen is much less dense than air.

If a sample of dry hydrogen is required the gas has to be passed through a drying agent before collection. Then the gas could either be collected in a gas syringe or by upward delivery.

Experiment 9.2; To demonstrate that hydrogen is less dense than air.

1. Prepare hydrogen as shown in figure 9.2
2. Collect some amounts of hydrogen in a test tube. Light a glowing splint and hold it near the mouth of the test tube. You will hear a pop sound as hydrogen burns explosively with the oxygen in the air to form water vapour. This is the test to identify hydrogen.
3. Collect the hydrogen in an inverted test tube labeled A.
4. Allow the gas to escape from test tube A to another test tube B as shown in figure 9.2. The test tube B should be inverted directly over the mouth of test tube A. B is filled with air.
5. After few minutes test the contents of the test tube A and B using a burning splint. By the end of experiment you will find that; when testing A with lighted splint there was no pop sound while on testing B with lighted splint there was a pop sound.

DISCUSSION QUESTION

What do these result show? And why does this happen?

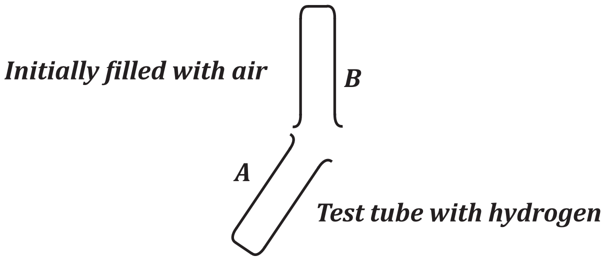


Figure 9.2; "Pouring" hydrogen upwards.

Experiment 9.3; demonstration on filling the balloon with hydrogen.

1. Add dilute sulphuric acid to a reagent bottle containing about 20g of zinc granules.
2. Tie a balloon to the mouth of the bottle using a cotton thread at point A; see figure 9.3
3. Allow the balloon to be fully inflated with hydrogen.
4. Tie another long cotton thread at point B around the mouth of the balloon.

5. Unite the thread A, and allow the balloon to rise in the air. Hold it in the air with the long thread.

You will see the balloon rise upward toward the ceiling. Leave the balloon inflated for few days. You will see the balloon eventually shrink as the hydrogen diffuses into the air.

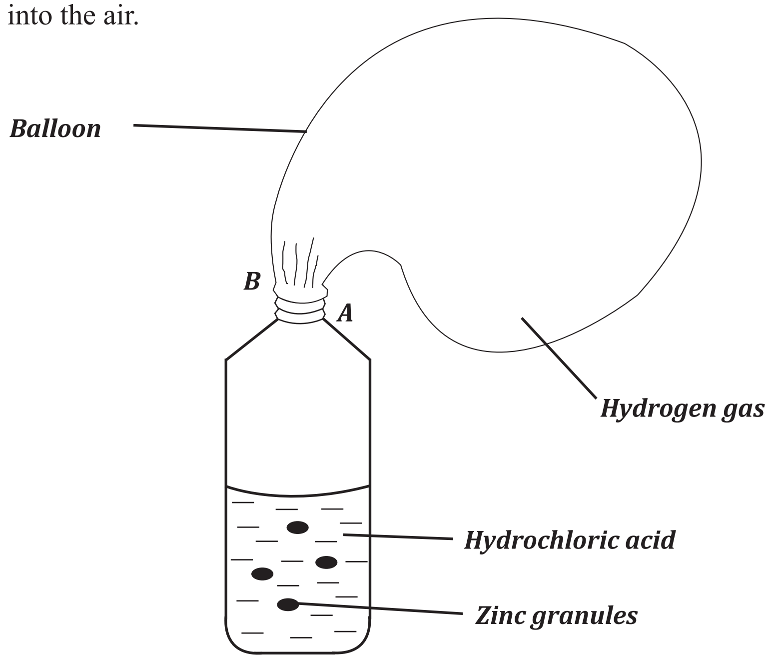


Figure 9.3; filling a balloon with hydrogen

9.3 PROPERTIES OF HYDROGEN

Physical properties

1. Hydrogen is a colourless gas
2. Pure hydrogen is odorless
3. Hydrogen is neutral to litmus paper or solution.
4. Hydrogen is the lightest gas
5. Hydrogen is insoluble in water
6. Hydrogen does not support combustion

Chemical properties of hydrogen

1. Hydrogen combines easily with other chemical substances at high temperatures
2. Hydrogen does not usually react with other elements at room temperature.
3. Hydrogen is highly flammable and burns with a blue flame.
4. A mixture of hydrogen and oxygen explodes when lit.
5. Hydrogen reacts with the oxides and chlorides of many metals to produce

free metals.

6. Hydrogen reacts slowly with oxygen to produce water. The reaction can be speeded up by catalyst
7. Hydrogen is neither acidic nor basic.
8. Hydrogen reduces metal oxides to metals.

Experiment 9.4; investigation of the effect of hydrogen on copper (ii) oxide.

1. Set up the apparatus as in figure 9.4 in which the test tube has a hole in the end.
2. Turn on the supply of hydrogen and let the gas pass all through the apparatus.
3. Light the jet to burn excess hydrogen.
4. Heat the tube close to the black copper (ii) oxide and observe the change from black to brown as the reaction occurs.
5. Leave the apparatus to cool before dismantling.

The reaction which takes place can be described as;

Word equation

Copper (II) oxide + Hydrogen ⟶ Copper + hydrogen oxide (water)

Balanced chemical equation

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

The copper (II) oxide is ***reduced*** as it loses oxygen. ***Reduction*** of copper (ii) oxide has taken place.

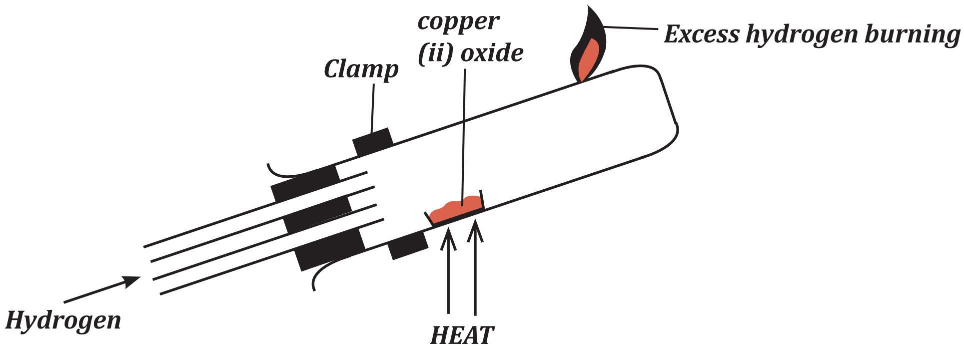


Fig9.4 Reduction of copper (ii) oxide with hydrogen

9.4 INDUSTRIAL PREPARATION OF HYDROGEN

Hydrogen is a very important industrial substance and has to be manufactured on a large scale. Some of the methods used to produce hydrogen industrially are explained as follows;

1. Electrolysis of water

Electrolysis of water is a process which decomposes water into oxygen and hydrogen gas by means of an electric current. The current is passed through the water. The electrical power sources are connected to two electrodes placed in the water. Hydrogen will collect at the cathode and oxygen will collect at the anode.

Reduction at cathode;

$2H^{+}_{(aq)} + 2\bar{e} \longrightarrow H_{2\,(g)}$

Oxidation at anode;

$2H_2O\ (l) \longrightarrow O_{2\,(g)} + 4H^{+}_{(aq)} + 4\bar{e}$

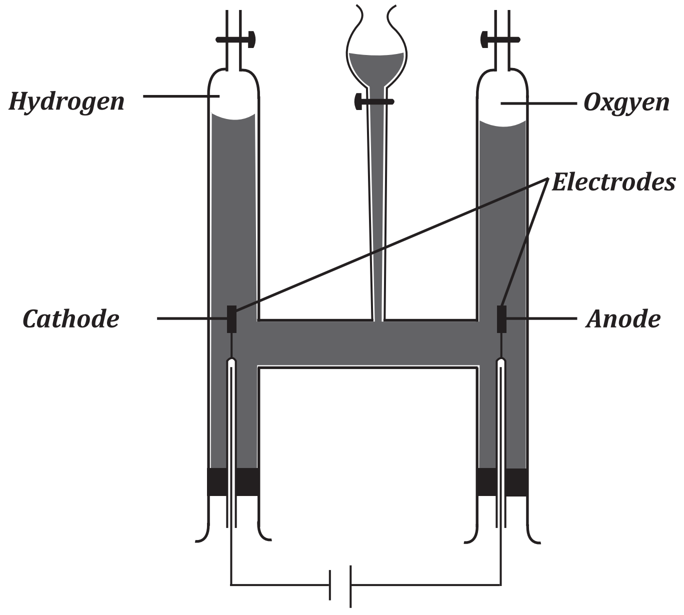


Fig 9.5 electrolysis of water using Hoffman voltameter.

2. The reaction of water with hot carbon (Bosch process).

Water gas is produced when steam is passed over white hot coke.

$$\underset{\textit{Coke}}{C_{(s)}} + \underset{\textit{steam}}{H_2O\ (g)} \longrightarrow \underbrace{CO\ (g) + H_2\ (g)}_{\textit{water gas}}$$

Hydrogen can be obtained from water gas by mixing the water gas with steam and passing the mixture over catalyst of Iron (III) Oxide at 450˚C.

$CO_{(g)}$	+	H_2O (g)	$\longrightarrow$	$CO_{2(g)}$	+	$H_{2(g)}$
carbon monoxide		***steam***		***Carbon dioxide***		***Hydrogen***

3. Steam reforming

The most common method of producing commercial large amount of hydrogen is steam reforming of natural gas, also known as ***steam methane reforming (SMR).*** It is a method of producing hydrogen from organic compounds like methane. At high temperatures (700 to 1100˚C), steam reacts with methane to yield carbon monoxide and hydrogen. The reaction take place in the presence of a metal - based catalyst.

Word equation;

Steam + methane $\longrightarrow$ carbon monoxide + hydrogen

Chemical equation;

$$H_2O_{(g)} + CH_{4(g)} \longrightarrow CO_{(g)} + 3H_{2(g)}$$

9.5 USES OF HYDROGEN

The following are main uses of hydrogen;

1.***Manufacture of ammonia (Haber process)***

Ammonia is a compound of nitrogen and hydrogen. Ammonia can be converted into a large number of nitrogen fertilizers e.g. Ammonium sulphate. Ammonia can also be used in production of nitric acid and synthetic (man made) fabrics, such as polyesters and polyamide fabrics.

$3H_{2(g)}$	+	$N_{2(g)}$	$\longrightarrow$	$2NH_{3(g)}$
Hydrogen		***Nitrogen***		***ammonia***

2.***Making margarine***

Margarine can be made from natural fats and oils such as sunflower oil or palm oil. These oils are very runny and have to be hardened before they can be used. The process is known as ***Hydrogenation*** or hardening of oils. Hardening involves reacting the heated oil with hydrogen using a nickel catalyst to speed up the reaction. Margarine is used as an alternative to butter.

3.***Oxy-hydrogen flame***

The oxy-hydrogen flame uses oxygen and hydrogen to form a hot flame and can get to temperatures of up to 3000˚C. This flame can be used for welding and cutting metals.

4.***Manufacture of hydrochloric acid***

Hydrogen is used in the manufacture of hydrochloric acid. It reacts with chlorine to form hydrogen chloride gas which dissolves in water to form hydrochloric acid.

$$\underset{\textbf{\textit{Hydrogen}}}{H_{2(g)}} + \underset{\textbf{\textit{Chlorine}}}{Cl_{2(g)}} \longrightarrow \underset{\textbf{\textit{Hydrogen chloride}}}{2HCl_{(g)}}$$

5. ***Fuel***

Hydrogen is used as a rocket fuel. Hydrogen is used to prepare water gas that can be burnt to propel rockets.

6. ***Filling the balloons***

Hydrogen is the lightest known gas and thus it is used for filling weather balloons and airships. However, since hydrogen is inflammable, it has been replaced by helium gas which is inert.

7. ***Making petrol***

Petrol contains about 14 percent of hydrogen and 5 percent of coal. Coal and hydrogen reacts to form liquid which contains petrol.

8.***Manufacture of methanol***

Hydrogen can combine directly with carbon monoxide to form methanol.

$$\underset{\textbf{\textit{Hydrogen}}}{2H_{2(g)}} + \underset{\textbf{\textit{Carbonmonoxide}}}{CO_{(g)}} \longrightarrow \underset{\textbf{\textit{Methanol}}}{CH_3OH_{(aq)}}$$

9. ***As a reducing agent***

Hydrogen is used to reduce metallic oxides to their metals. The oxide of metals such as iron, copper and lead are reduced to the metals when heated in a steam of gas.

$$\underset{\textbf{\textit{Copper (II) oxide}}}{CuO_{(s)}} + \underset{\textbf{\textit{Hydrogen}}}{H_{2(g)}} \longrightarrow \underset{\textbf{\textit{Copper}}}{Cu_{(s)}} + \underset{\textbf{\textit{water}}}{H_2O_{(l)}}$$

SUMMARY

(a). Hydrogen is a chemical element with symbol H atomic number 1, and Atomic weight 1.00794.

(b). At standard temperature and pressure, hydrogen is colourless, odourless and tasteless.

(c). The most common isotope of hydrogen is protium with symbol ^{1}H.

(d). Industrial production of hydrogen is mainly from the steam reforming of natural gas, and less often from more energy - intensive hydrogen production methods like the electrolysis of water.

(e). Other isotopes of hydrogen are deuterium (^{2}H) and tritium ^{3}H.

(f). Hydrogen is used to manufacture ammonia because it readily combines with elements e.g. nitrogen.

(g). Hydrogen is used in production of oxy-hydrogen flame because it is highly flammable.

(h). Hydrogen is used to manufacture hydrochloric acid due to the fact that it readily reacts with other chemical substances.

(i). Hydrogen is used to prepare water because it is highly flammable.

(j). Hydrogen was once used to inflate weather balloons because it is lighter than air.

(k). Hydrogen is used to manufacture margarine because it is a reducing agent.

(l). The boiling point of hydrogen is -252.9˚C and the melting point is -259.34˚C

(m). The density of hydrogen is $8.988 \times 10^{-5} g/cm^3$.

(n). Elemental classification of hydrogen is non-metal.

REVIEW QUESTIONS

1. Choose the most correct answer for the following questions;

   i. While oxygen is the most abundant element on earth's outer crust .... is the most abundant element in the universe.

      A. Nitrogen B. Hydrogen C. Chlorine D. Carbon dioxide.

   ii. The common method for preparing hydrogen in the laboratory is;.............

      A. Action of moderately reactive metal on dilute acid.

      B. The action of moderately reactive metal on concentrated acid.

      C. Water by electrolysis.

      D. Reaction of water with hot carbon.

   iii. The test for hydrogen;__________________

      A. Ignites with a "pop" sound when a lighted splint is inserted into a test tube with a gas.

B. Relights a wooden glowing splint.
C. Changes litmus paper red
D. Changes litmus paper blue.

iv. Physical properties of hydrogen gas include;
A. A mixture of hydrogen and oxygen explodes when lit.
B. It is neither acidic nor basic.
C. It is tasteless, colourless and odourless.
D. Combines easily with other chemical substances.

v. Among others, uses of hydrogen include;
A. In the oxy-carbon dioxide flame
B. Manufacture of carbon dioxide
C. As an oxidizing agent
D. Fuel, manufacture of methanol and synthesis of hydrochloric acid

2. Match the items in LIST A with their corresponding statements in LIST B.

List A	List B
x. The addition of oxygen to a substance or the removal of hydrogen from the substance..............	A. Protium (^1H) and tritium (^3H)
xi. The removal of oxygen from a substance or the addition of hydrogen to a substance..............	B. Electrolysis of water
xii. Is the process of passing hydrogen through liquid oil to harden it..............	C. Liquid hydrogen and gaseous hydrogen
xiii. .... is a process which decomposes water into oxygen and hydrogen gas, with the aid of an electric current.	D. Hydrogenation
xiv. Isotopes of hydrogen......................	E. Hydro oxidation
	F. Reduction
	G. Oxidation
	H. Combined reduction and oxidation.

3. Write ***TRUE*** for a correct statement and ***FALSE*** for an incorrect statement;
i. Hydrogen can be collected over water because it does no dissolve in water........
ii. Hydrogen can also be collected by upward delivery because hydrogen is much less dense than air.................................
iii. Hydrogen is nowadays used for weather balloons and airships....................
iv. Hydrogen is widely used as an oxiding agent................................
v. Ammonia is a compound of nitrogen and hydrogen...........................

4. (a) List any three physical properties and any three chemical properties of hydrogen

   (b) Mention three properties of hydrogen that make it different from oxygen

5. Describe three methods for the industrial manufacture of hydrogen.
6. Hydrogen is a very attractive source of energy but its use as a major source of energy is not practical. Explain this in terms of its storage, safety and production.
7. What is a chemical test for hydrogen gas?
8. Why is zinc a better choice than sodium and potassium as a metal for reacting with acid to give hydrogen?
9. Explain the laboratory preparation of hydrogen gas from zinc and dilute sulphuric acid.

Chapter 10

Water

10.1 OCCURRENCE AND NATURE OF WATER.

Water is a compound of Hydrogen and Oxygen. It is a very important commodity because it is essential for the sustenance of all living things. Water is also a home for some animals and plants.

Development depends on availability of water. In areas where there is plenty of water, plants and livestock flourish. In dry desert areas, plants, animals and people are scarce.

Three quarters of the earth's surface is covered by oceans and seas. On the average, oceans and seas water contains 3. 6 percentages by mass of dissolved solids as follows; common salts 2.8 %, calcium salts 0.1 %, magnesium salts 0.6%, potassium salts 0.1%. Bromine, Iodine, Copper, Silver and even Gold and Radium are often found in the sea.

THE WATER CYCLE

Water is continually moving above and below the earth, as water vapour, liquid water and ice. The water is never lost but it continually being recycled all around the globe in a system called the water cycle (hydrological cycle).

The water cycle is made up of main five parts.

1. Evaporation.
2. Transpiration
3. Condensation
4. Precipitation
5. Collection.

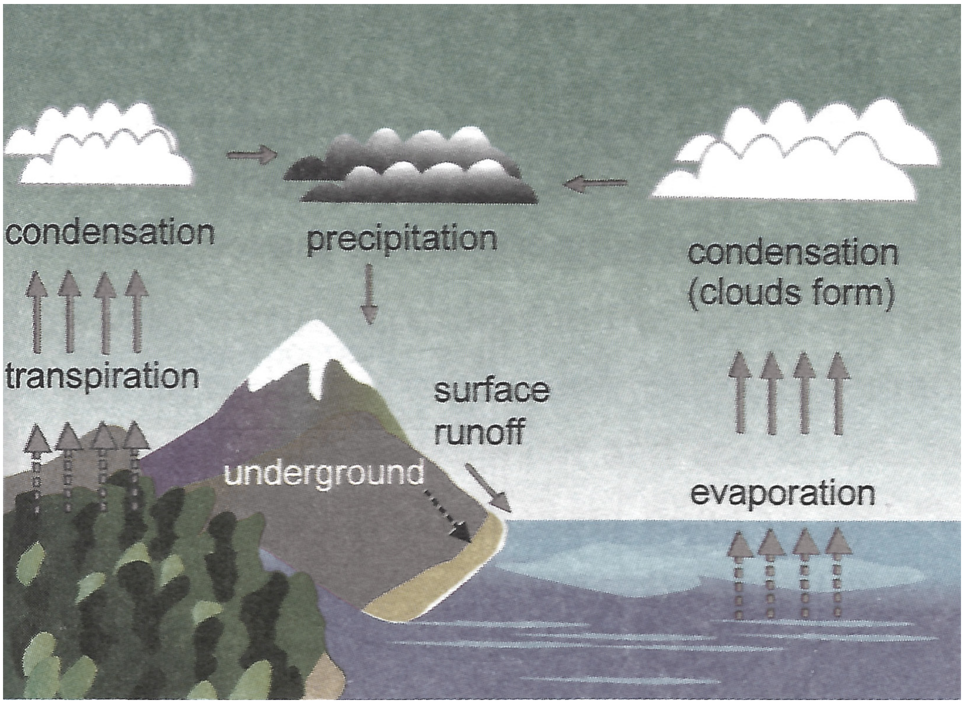


Fig 10.1 the water cycle.

1. ***Evaporation***

   Evaporation is the process in which water vapour leaves the streams, rivers, lakes or oceans and goes into the atmosphere. The sun provides the driving force by heating up the water bodies and turns the water into vapour.

2. ***Transpiration***

   This is the process in which plants lose water through their leaves. This is another way in which water gets back in the atmosphere. This vapour undergoes condensation to form clouds.

3. ***Condensation***

   This is the process in which water vapour in the air cools and changes back into liquid, thereby forming clouds.

4. ***Precipitation***

   This is the process in which clouds get heavy and water falls in form of rainfall, hail or snow.

5. ***Collection***

   Collection can take place when water falls back to earth as precipitation. The water may go into oceans, lakes or rivers, or on land. When water falls on land, It normally soak into the soil and becomes part of the ground water which plants and animals use.

RELATIONSHIP BETWEEN WATER CYCLE AND ENVIRONMENTAL CONSERVATION.

People must observe principles and rules so that the environment can be conserved. Conservation of Natural resources like plants, water and land refers to environmental conservation. Land clearing especially near sources of water should be avoided. Drying of ponds and rivers would in turn affect the normal circulation of water. Tree planting and land reclamation programs should be encouraged in order to conserve the natural environment. The use of natural gas as the source of fuel should be encouraged.

10.2 PROPERTIES OF WATER

The physical properties of water involve aspects such as colour, taste and smell. They also involve melting, freezing and boiling points. Chemical properties of water involve the results that occur when water reacts with other substances.

Experiment 10.1

Aim: to measure the melting and boiling points of water.

Apparatus; Bunsen burner, tripod stand, thermometer, retort stand and clamp, two test tubes.

Procedure;

1. Put some ice cubes in one test tube and water in another test tube.
2. Clamp the thermometer in a vertical position
3. Set the apparatus as illustrated in fig 10.2. Adjust the thermometer so that it dips into the ice cubes, but does not touch the test tube. Record the temperature of the ice cubes.
4. Heat the water in another test tube. Record the temperature after water forms steam.

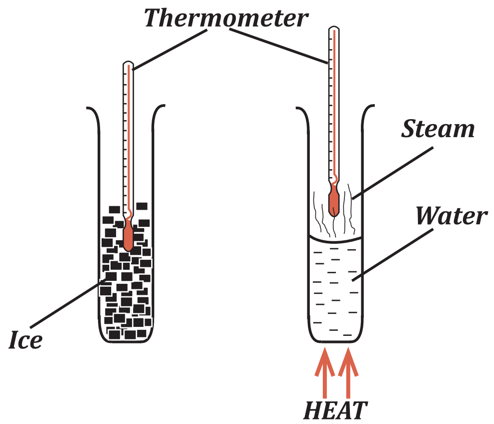


Fig 10.2 finding the melting point and boiling point of water

Discussion questions.

1. Why was the thermometer not allowed to touch the test tube?
2. What was the temperature of the ice cubes?
3. What was the temperature of the boiling water?

The thermometer should not touch the test tube in order to avoid altering the reading on the thermometer.

At atmospheric pressure, pure water boils at 100˚C and freezes at O˚C. Changes in altitude or pressure alter the boiling point of water.

Experiment 10.2

Aim: To show the presence of water.

Materials; Distilled water, blue and red litmus papers, cobalt chloride paper, watch glasses, anhydrous Copper (II) Sulphate.

Procedure;

1. Pour some amount of distilled water onto a watch glass.
2. Dip a strip of blue litmus paper into the water on the watch glass.
3. Repeat the steps 1 and 2 using red litmus paper and cobalt chloride paper respectively.
4. Put some amount of anhydrous Copper (II) Sulphate on a watch glass and add some little distilled water.

Discussion questions.

1. What colour changes were obtained on the blue litmus, red litmus and cobalt chloride papers?
2. What was the colour change observed on anhydrous Copper (II) Sulphate when some distilled water was added?

Water is neutral i.e. neither acidic nor basic and therefore it has no effect on either red or blue litmus papers.

Water reacts with cobalt chloride paper and the paper changes from blue to pink. The cobalt chloride paper is used to test for the presence of water in substances.

White anhydrous Copper (II) Sulphate turns blue when water is added. The blue colour indicates the hydrated Copper (II) Sulphate is formed.

Experiment 10.3

Aim: the synthesis of water.

Materials: U-tube apparatus, filter pump, thistle funnel, test tube.

Procedure;

1. Pass a stream of Hydrogen through a U – tube containing anhydrous calcium chloride to ensure that no water is present.
2. Allow all the air to be displaced from the apparatus.
3. Ignite the gas coming out of the apparatus. When the sample burns quietly with blue flame, it indicates that all the air has been displaced and the gas coming off is pure Hydrogen.
4. Assemble the rest of the apparatus as shown in figure 10.3 below.
5. Ignite the Hydrogen so that the flame is about $3cm^3$ high.
6. Draw the products of the combustion of Hydrogen through the apparatus by means of filter pump for about 10 minutes.
7. Test any liquid which collect in the test tube.

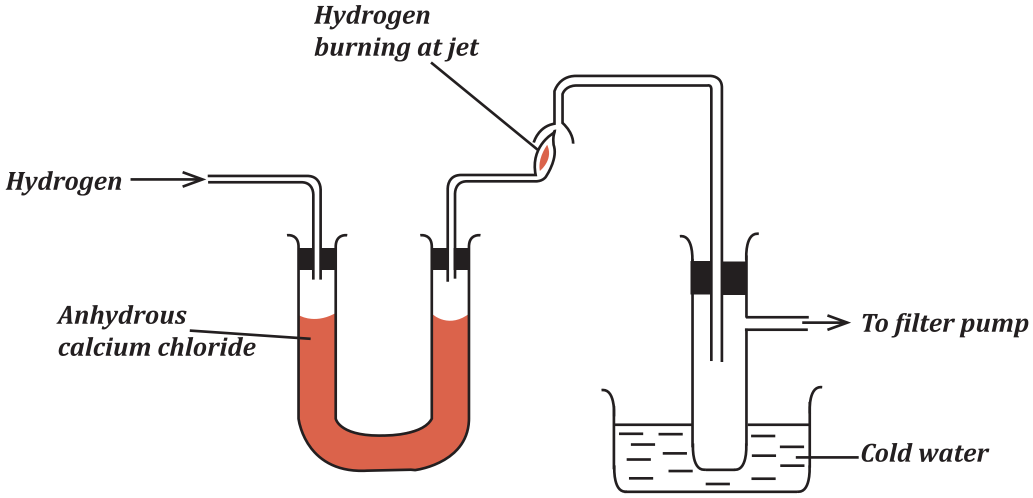


Fig. 10.3 Synthesis of water

By the end of experiment, a colourless liquid collects in the test tube which can be shown to be water by testing it with anhydrous Copper (II) Sulphate. Water turns anhydrous white Copper (II) Sulphate blue.

When few drops of the liquid formed are added to the blue cobalt chloride paper the paper will turn pink and this shows that the liquid formed is water.

Experiment 10.4: to study the reactions of sodium and potassium with cold water.

1. Pour cold water into a trough until it is half filled. Take sodium out of a bottle with dry tongs. Use a dry knife and a dry white tile to cut four small pieces each about the size of a rice grain. Return the rest of the sodium to the bottle and replace the lid. What colour is the metal when freshly cut, and how does it change if left in air? Scrape off any white or brown crust that is on the metal. ***Never touch sodium or potassium with your fingers.***
2. Drop a single piece of sodium into the water in the basin, keeping your face away in case an explosion occurs. Does the sodium sink or float? Does it move around in the water? What is its shape and appearance? Does it burn at any time and if so, what is the colour of the flame? What is the colour of and what happens to the last bit of sodium?
3. Add two more pieces of sodium to the water, one at a time.
4. Feel the solution between your fingers. How does it feel?
5. Add some universal indicator to the solution in the trough. What colour change occurs?
6. Repeat the whole test with a fresh sample of cold water and tiny pieces of potassium, which reacts more vigorously and dangerously than sodium. What other differences are there between the reactions of sodium and potassium on water?

Potassium is a bright silvery metal. It turns dull quickly in air as it reacts with Oxygen. It is therefore kept under paraffin oil. It is so soft that it can be cut easily with a knife. Its density is low so it floats on water.

The reaction between potassium and water is like that between sodium and water.

Potassium is more reactive and so reacts more quickly and the Hydrogen formed at once burns with a purple flame.

Sodium is a silvery, shiny metal. It is malleable and soft and can be cut with a knife. The cut surface at once turns dull because sodium reacts with Oxygen in the air. The density of sodium is low so it floats on water.

When sodium is added to water, it first melts to a silvery ball and moves about on the surface of the water; and eventually gives off Hydrogen and gradually becomes smaller as it reacts.

The solution left feels soapy and greasy when rubbed between the fingers, and it turns universal indicator purple.

Experiment 10.5 Collecting the gas formed when sodium reacts with water.

1. Cut three pieces of sodium as described in the previous experiment. Wrap them in wire gauze so that it forms a cage around them. Leave room for the sodium to move in the cages, if it cannot move there is more danger of an explosion during the experiment.
2. Fill a boiling tube with water and invert it in a trough of water.
3. Drop the wire cage into the beaker and move the boiling tube over it to collect any gas formed.
4. Remove the tube and test the gas with a burning splint. What happens?

Caution;
Do not attempt this experiment with potassium

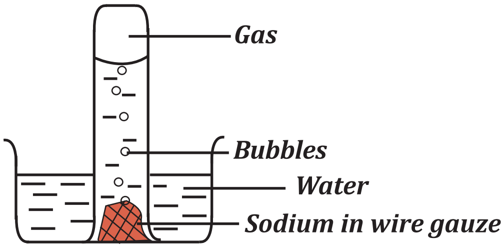


Fig.10.4; Collecting the gas formed when sodium reacts with water

Experiment 10.6 To study the actions of calcium and magnesium on water and collect any gas formed.

(a) Calcium

1. Fill a beaker with water. Use tongs to drop three small pieces of calcium into the water. What happens?
2. Pour some of the solution into a test tube and test it with universal indicator. What change occurs?
3. Filter some of the solution into a test tube. Blow through glass tubing into the filtrate. What change occurs? What do you think the solution is?
4. Collecting the gas.
   a. Fill a boiling tube with water and invert it in the beaker.
   b. Add pieces of calcium to the beaker and cover them with filter funnel. Move the boiling tube over the funnel to collect any gas formed. The stem of the funnel must be short.
   c. Remove the tube and test the gas with a burning splint. What happens?

(b) Magnesium

1. Rub the magnesium ribbon with sand paper until it is bright and shiny, roll the ribbon into a ball and drop it into water. Leave for about 10 minutes. What do you observe?
2. Place a boiling tube filled with water over the magnesium and leave it for 2 – 3 days.

Does any gas collect in the tube? if so, test it with a burning splint.

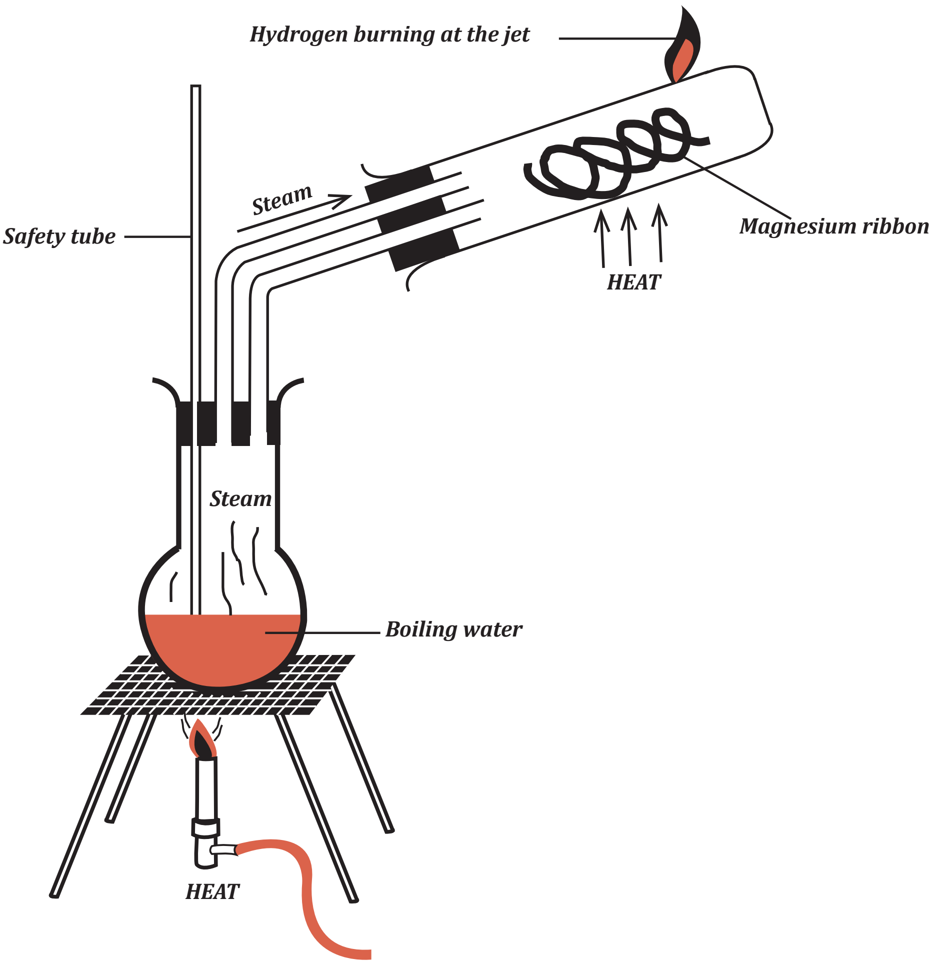


Fig. 10.5 the reaction of magnesium and steam.

Experiment10.7 to study the reaction of magnesium with steam.

1. Rub magnesium ribbon with sandpaper until it is bright and shiny.
2. Set up the apparatus in figure 10.5 and heat the flask containing water until there is a stream of steam going through the apparatus.
3. Then heat the test tube until the magnesium starts to burn.
4. Light the jet of Hydrogen gas leaving the test tube.

Calcium is a shiny, silvery metal. Usually it looks white because calcium oxide forms on its surface. A piece of calcium sinks in water and reacts quietly to form bubbles of Hydrogen. The calcium gradually disappears and forms calcium hydroxide solution. The water becomes cloudy because the calcium hydroxide is not very soluble and it forms a suspension.

Magnesium reacts very slowly with cold water but faster with hot water. Heated magnesium catches fire in steam and burns brightly. The reaction releases a lot of energy and the Hydrogen formed catches fire. White magnesium oxide is also formed. Magnesium hydroxide is not formed because the temperature is too high, but it is formed when the metal reacts with hot or cold water.

Experiment 10.8 to examine the action of steam on heated iron.

1. Place iron powder in the middle part of a combustion tube.
2. Arrange the apparatus as shown in figure 10.6 (but without the gas jar in position)
3. Boil the water gently. Warm the end of the combustion tube where the steam enters so that it does not condense. When all the air is out of the apparatus, place the gas jar in position. What causes the noise in the trough?
4. Heat the iron and keep the steam passing slowly. What happens to the iron? What collects in the gas jar? The safety tube prevents "sucking back" at the end of the experiment when the apparatus is cooling. How does it do this? It also allows water and steam to escape if the apparatus becomes blocked.
5. Test the gas with a burning splint. What happens? What is the gas? Describe the solid residue.

   Steam reacts slowly with heated iron to form Hydrogen and a blue – black Iron oxide.

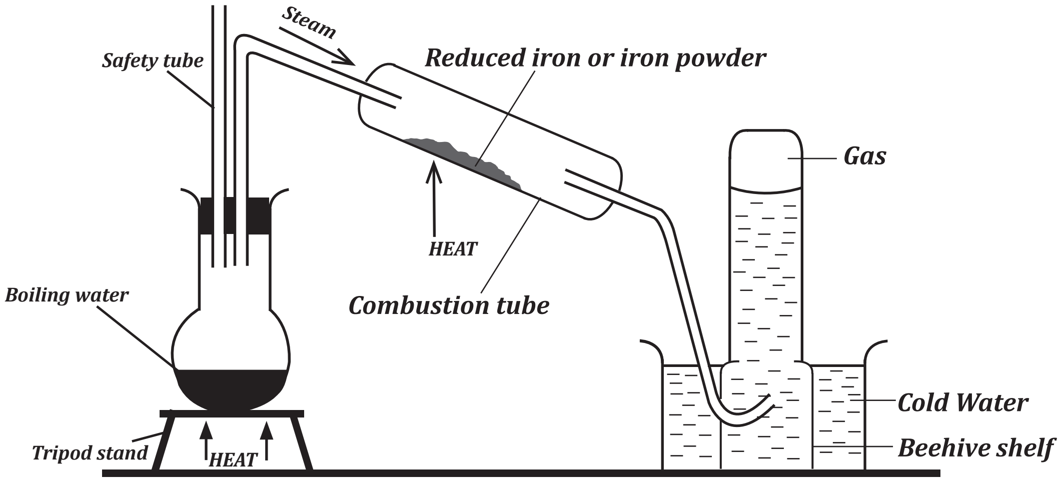


Fig. 10.6 The action of steam on heated iron

According to the experiments performed previously. The properties of water can be described as follows;

Physical properties.

1. Water is colourless, odourless, and tasteless.
2. Water is the only substance that occurs naturally in all three states of matter.
3. The freezing point of water is 0˚C and the boiling point is 100˚C
4. Water is the universal solvent because it dissolve more substances than any other liquid.
5. Water has a high surface tension. It is sticky and elastic and tends to clump together in drops rather than spread out in a thin film.
6. Water has a high specific heat index. It can absorb a lot of heat before it begins to get hot.
7. Water expands when it freeze. Ice is therefore less dense than liquid water.
8. Water is miscible with many liquids example ethanol.

Chemical properties.

1. Pure water is neutral that is it is neither acidic nor basic
2. Cold water reacts with some metals to form metal hydroxides and liberate Hydrogen.
3. Some metals react with steam to give the respective metal oxide and Hydrogen gas.

TEST FOR WATER

There are two main methods which can be used to test water in laboratory.

i. Water changes cobalt chloride paper from blue to pink.

$$\underset{\textit{Cobalt (ii) chloride (Blue)}}{CoCl_{2(s)}} + 6H_2O_{(l)} \longrightarrow \underbrace{CoCl_2 \bullet 6H_2O_{(s)}}_{\textit{(pink)}}$$

ii. Water changes anhydrous Copper (ii) sulphate from white to blue.

$$\underset{\textit{Copper (ii) sulphate (White)}}{CuSO_{4(s)}} + 5H_2O_{(l)} \longrightarrow \underbrace{CuSO_4 \bullet 5H_2O_{(s)}}_{\textit{(blue)}}$$

REACTION OF WATER WITH METALS

(a) Potassium.

Potassium is bright silvery metal. The reaction between potassium and water is like that between sodium and water. Potassium is more reactive and so reacts more quickly and the Hydrogen formed at once burns with a purple flame.

Word equation:

Potassium + water $\longrightarrow$ potassium hydroxide + Hydrogen

Chemical equation:

$$2K_{(s)} + 2H_2O_{(1)} \longrightarrow 2KOH_{(aq)} + H_{2(g)}$$

Potassium is kept under paraffin oil because it quickly react with Oxygen in air and turns dull.

(b) Sodium.

Sodium is a silvery, shiny metal. It is malleable and soft and can be cut with a knife. When added to water sodium;

i. First melt to a silvery ball.
ii. Move about on the surface of the water.
iii. Gives off Hydrogen and gradually become smaller as it reacts.

The solution left feels soapy or greasy when rubbed between the fingers, and it turns universal indicator purple.

Word equation:

Sodium + water $\longrightarrow$ Sodium hydroxide + Hydrogen

Chemical equation:

$$2Na_{(s)} + 2H_2O_{(1)} \longrightarrow 2NaOH_{(aq)} + H_{2(g)}$$

(c) Calcium.

Calcium is a shiny, silvery metal. Usually it looks white because calcium oxide form on its surface. A piece of calcium sinks in water and reacts quietly to form bubbles of Hydrogen

Word equation:

Calcium + water $\longrightarrow$ Calcium hydroxide + Hydrogen

Chemical equation:

$$Ca_{(s)} + 2H_2O_{(1)} \longrightarrow Ca(OH)_{(2)(s)} + H_{2(g)}$$

(d) Magnesium.

Magnesium reacts very slowly with cold water but faster with hot water. Heated magnesium catches fire in steam and burns brightly. The reaction release a lot of energy and the Hydrogen formed catches fire. White magnesium oxide is also formed.

Word equation:

Magnesium + Steam $\longrightarrow$ Magnesium oxide + Hydrogen

Chemical equation:

$$Mg_{(s)} + H_2O_{(1)} \longrightarrow MgO_{(s)} + H_{2(g)}$$

Magnesium hydroxide is not formed as the temperature is too high, but it is formed when the metal reacts with hot or cold water.

Word equation:

Magnesium + Water $\longrightarrow$ Magnesium hydroxide + Hydrogen

Chemical equation:

$$Mg_{(s)} + 2H_2O_{(1)} \longrightarrow Mg(OH)_{2(aq)} + H_{2(g)}$$

(e) Iron.

Steam reacts slowly with heated iron to form Hydrogen and a blue-black iron oxide.

Word equation:

Iron + Steam $\longrightarrow$ Iron oxide + Hydrogen

Chemical equation:

$$Fe_{(s)} + H_2O_{(g)} \longrightarrow FeO_{(s)} + H_{2(g)}$$

(f) Zinc.

Zinc does not react with hot or cold water. If zinc is heated to redness in a currect of steam, Hydrogen is formed.

Word equation:

Zinc + Steam $\longrightarrow$ Zinc oxide + Hydrogen

Chemical equation:

$$Zn_{(s)} + H_2O_{(g)} \longrightarrow \underbrace{ZnO_{(s)}}_{\textbf{\textit{Zinc Oxide}}} + H_{2(g)}$$

10.3 WATER TREATMENT AND PURIFICATION.

Water treatment; is the process of making water usable for domestic, industrial, medical and other purposes. The aim of the treatment process is to remove existing contaminants in the water, thus improve it for safe use. Treatment processes may be physical such as settling, chemical such as disinfection, or biological such as slow sand filtration.

Water purification; is the removal of contaminants from treated water to produce drinking water that is pure enough for human consumption. Substances that are removed include bacteria, algae, and fungi, minerals such as iron and sulphur and human made chemical pollutants.

DOMESTIC WATER PURIFICATION.

Simple and diverse methods used to treat water to make it safe can be explained as follows.

1. Filtering.

Filtering will remove solid particles. The size of the holes of a filter is important. There are two basic types of filters.

Membrane filters use thin sheets with precisely sized pores that prevent objects larger than the pore size from passing through.

Depth filters use thick porous materials such as carbon or ceramic to trap particles as water flows through the materials. Activated carbon filters also remove a range of organic chemicals and heavy metals.

There is a difference between a water filter and a water purifier. Filters do not filter out viruses, but there are water purifiers that pass the water through both a filter and an iodine compound that kills any smaller organisms that have passed through the filter.

2. Boiling.

Boiling is the most efficient way of killing all microorganisms. Water temperatures above 70˚C kill all pathogens within 30 minutes and above 85˚C within a few minutes. So in the time it takes for the water to reach the boiling point (100˚C) from 70˚C, all pathogens will be killed, even at high altitude.The boiled water is then allowed to cool before being filtered using a clean cloth.

3. Adding water treatment chemicals.

There are two types of chemicals; those using iodine and those using chlorine. There is a variety of products on the market, so follow the directions on the bottle. Be advised that many of the tablets have an expiry date and become ineffective after that point. Remember that chemical purification methods may only be partially effective, depending on the water temperature.

Iodine treatment.

Iodine is light sensitive and must always be stored in a dark bottle. It works best if the water temperature is over 21˚C. Iodine has been shown to be more effective than chlorine based treatments in inactivating Giardia Cysts.

Chlorine treatment.

Chlorine can be used to treat water by killing microorganisms in the water. Halazone is an example of a chlorine tablet product. To use, follow the manufacturer's instructions.

Experiment 10.9

Aim; to filtrate river water in the laboratory.

1. Using a filter paper collect some water with a bucket from a nearby river, pour some water into flask and filter this water by means of a filter paper.
2. Using sand filter, instead of a filter paper, carefully carry some large pieces of gravel to the bottom of the funnel. Arrange a layer of smaller pieces of gravel on top of large pieces of gravel, followed by a layer of fine sand. Pour a sample of river water in a beaker.

It can be observed that in both cases the filtrate may appear relatively clear. The sand filter bed method is also used to purify some public water supplies. These filter beds are filled with fine layers of gravel and then sand. The water is allowed to percolate slowly through these layers and both organic and inorganic debris are retained by them.

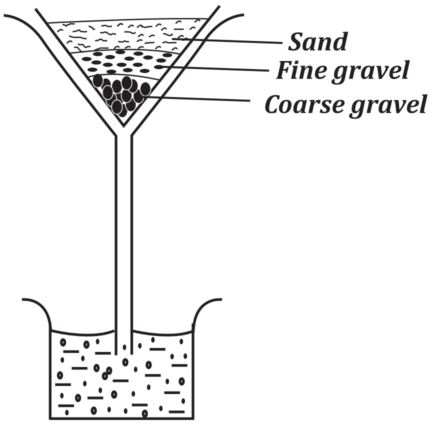


Fig.10.7; sand filter

URBAN WATER TREATMENT.

River water passes through a series of treatment stage before it is fit for urban consumption. The stages can be explained as follows;

1. Chlorination and Aeration. After filtering, the water may be aerated. Water is pumped through fountains and spouts up into the air. Aeration kills many dangerous aquatic bacteria. Another chemical added in water to kill harmful bacteria is liquid sodium hypochlorite, which releases chlorine in water. Chlorine is such a useful disinfectant that is used in swimming pools to kill bacteria.

2. Sedimentation. Dirty water from a river is pumped and stored in large reservoirs where solid particles are allowed to settle to the bottom. The process takes long time for most of the particles to settle.

3. Chemical treatment. Even after a long time colloidal materials will remain suspended. The deposition of colloidal materials such as clay is affected by the addition of a chemical precipitant.

Aluminum sulphate, in the form of potash alum, and slaked lime (calcium hydroxide) are added to precipitate the suspended clay matter. The two chemicals react to form aluminium hydroxide and calcium sulphate.

The aluminium hydroxide is bulky and sticky, the calcium sulphate is far denser than water. Both the solid products settle to the bottom.

Aluminium sulphate + calcium hydroxide → aluminium hydroxide + calcium sulphate.

4. Filtration. Water from the reservoirs is passed through a sand filter. The filter is made up of layers of sand, charcoal and gravel. The charcoal is to remove colored matter and noxious smells from the water. Filtration beds are expensive to install and require considerable labour to maintain. The above processes aim at making water suitable for drinking and hence the water is said to be ***potable***.

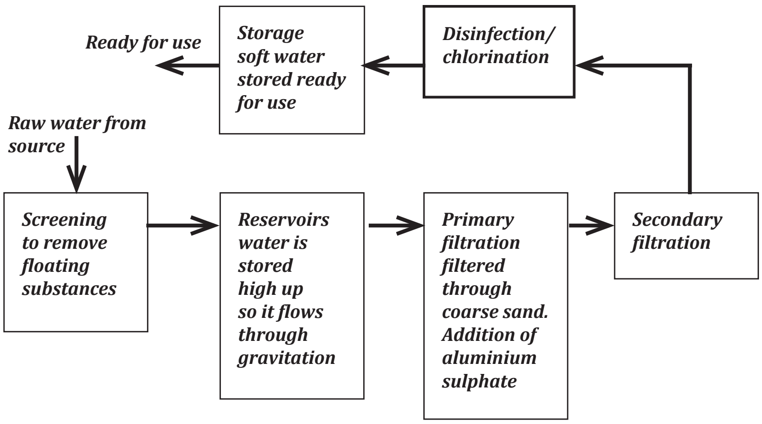


Fig.10.8 Process of urban water treatment

10.4 IMPORTANCE OF WATER TREATMENT

The following are some of the reasons for why water needs to be treated.

1. Untreated water may contain harmful bacteria and other parasites that can cause diarrhoea, typhoid, cholera and other diseases.
2. Treated water is the best for use in laboratories and medical facilities to ensure accurate results from experiments and effective treatment.
3. Treated water is suitable to use in factories in order for manufactured products to be safe for consumption.
4. Treated water is more efficient to use for cleaning in industries and in domestic settings. Untreated water will usually lead to usage of greater amounts of certain substances such as soaps for cleaning.

10.5 USES OF WATER.

There are many ways in which water is useful and that is why it is very important to conserve water. Among them the following are uses of water.

1. Drinking, washing, cleaning and cooking.
2. Irrigation
3. Transportation of people and goods.
4. Generation of hydroelectric power (electricity)
5. Filling swimming pools for recreation activities.
6. Fire fighting.
7. In industries as steam for running machines
8. Cooling engines and parts of machines.

SUMMARY

(a). Water is important on day to day basis for various purposes such as drinking, cooking, cleaning, washing our bodies and clothes.
(b). Water is also used for transportation and recreation activities.
(c). Water is the most abundant compound on Earth's surface, covering about 75% (percent) of the planet.
(d). In nature water exists in liquid, solid and gaseous states.
(e). The density of water is 1,000.00 kg/m^3
(f). Water is a chemical compound with the chemical formula H_2O.
(g). A water molecule contains one oxygen atom and two hydrogen atoms connected by covalent bonds.
(h). Pure water has a low electrical conductivity, but this increases with the dissolution of a small amount of ionic material such as sodium chloride.

REVIEW QUESTIONS

1. Choose the best answer to complete each statement.

   i. A compound of Hydrogen and Oxygen is……………………….
      A. Carbon dioxide B. Water
      C. Sulphuric acid D. Liquid nitrogen

   ii. The water cycle is also referred to as
      A. Hydrological cycle B. Hydrogen cycle.
      C. Carbon cycle D. Nitrogen cycle.

   iii. The water on earth occurs in ………. main states.
      A. Two B. One C. Four D. Three

   iv. The boiling point of water is…………….
      A. 0˚C B. 90˚C C. 100˚C D. 95˚C

   v. Most of the Earth's water is…………………
      A. Pure B. Not pure C. More pure D. Clean.

2. Match the following items by writing a letter of their correct statement in the spaces provided.

List A	List B
i. Evaporation	A. The process in which water vapour cools and changes back into liquid.
ii. Transpiration	B. The process in which plants lose water through their leaves.
iii. Condensation	C. Take place when water falls back to earth and go into oceans, lakes, rivers and on land.
iv. Precipitation	D. The process in which clouds get heavy and water falls in form of rainfall, hail or snow.
v. Collection	E. The process in which water vapour leaves the streams, rivers, lakes and oceans and goes into the atmosphere.
	F. The process in which clouds get heavy.

3. Write ***TRUE*** for a correct statement and ***FALSE*** for an incorrect statement.

i. Water is used for irrigation, in animal dips and for watering flowers……….
ii. In the extraction of certain minerals, water is used as a solvent, or as a means of carrying away impurities.
iii. Large water bodies, especially rivers and artificial lakes are used to generate electrical energy…………..
iv. Water bodies such as oceans, lakes and rivers are not used for fishing…………
v. Water treatment is the removal of contaminants from treated water to produce drinking water that is pure enough for human consumption.

4. (a) Define water.
   (b) Name four different kinds of natural water.

5. Unknown liquid boils at 100˚C and freezes at 0˚C. It has a density of $1g/cm^3$ at 4˚C
   (a) What is the name of that liquid
   (b) Give one test of that liquid.

6. Define potable water.
7. (a) Draw the water cycle and describe the stages in the process
   (b) Explain the importance of water cycle.

8. You are given a sample of colourless liquid and you carry out some tests. Here are the results;

Test	Result
Add anhydrous Copper(ii)Sulphate	Turns from white to blue
pH	7
Evaporate to dryness	White residue
Boiling temperature	105˚C

   (a) Which is the best test to tell you water is in the liquid?
   (b) What does the test suggest about the liquid?

9. (a) Explain the meaning of these terms as far as water treatment is concerned; filtration, sedimentation, chlorination.
   (b)List down four methods of purifying water for cities and towns.

Chapter 11

Fuels and Energy

Fuel is any substance that stores and releases usable energy when its chemical or physical structure is converted.

Fuel releases its energy either through chemical means such as burning or nuclear means such as nuclear fission and nuclear fusion. An important property of fuel is that its energy can be stored to be released only when needed and the release of energy is controlled in such a way that the energy can be harnessed to produce work.

Energy is defined as the capacity to perform work. Energy exists in several forms such as heat, kinetic or mechanical energy, light, potential energy, electrical energy or other forms.

11.1 SOURCES OF FUEL

The energy we use every day comes in many different forms from many different sources, but they can be categorized as two basic types: non-renewable (those that cannot be replenished in a short length of time) and renewable (those that can be replenished in a short length of time).

NON - RENEWABLE FUELS

These are fuels that cannot be replenished in a short length of time. They are considered non-renewable due to the many years it took to create them and mankind's inability to synthesize similar fuels readily. Examples of non-renewable energy sources are: crude oil, liquefied petroleum gas (LPG), natural gas, coal and uranium. All except uranium are called fossil fuels because they originate from decaying plants and animal matter.

RENEWABLE FUELS

These are fuels that can be replenished in a short length of time. The renewable fuels are abundant and constantly replenished.

FOSSIL FUELS

Fossil fuels are fuels formed by natural processes such as anaerobic decomposition of buried dead organisms.

- The age of the organisms and their resulting fossil fuels is typically millions of years.
- Fossil fuels contain high percentages of carbon.
- Fossil fuels include coal, petroleum and natural gas.
- The energy that is released when fossil fuels are burned originates from the sun. It was captured by the plants through photosynthesis. The same principle also applies to charcoal and wood.

EFFICIENCY OF A FUEL

During releasing energy a fuel usually does not release all of the stored energy. Much of the energy is wasted. In a normal fire used for cooking by employing fire wood or charcoal, only about one fifth of the energy stored is actually released in a useful form to the surroundings. The efficiency of the energy source is the ratio of the useful energy released to the total amount of energy stored. It is usually expressed as a percentage, so the efficiency of the fire is about 20%. The efficiency of a car is 40%, a lamp is 10% and that of a microwave cooker is 80%.

CHARACTERISTICS OF GOOD FUELS

There are characteristics which determine good fuel because fuels differ in quality. Good fuels have the following characteristics:

1. They ignite easily
2. They burn well without explosion
3. They give out a lot of heat
4. They are not expensive
5. They have low smoke and ash content
6. They can be stored and transported easily.
7. They do not give off dangerous by-products like poisonous fumes.
8. They are easily controlled.

Activity 11.1; to compare the energy values of foods

Collect small samples of food e.g. bread and biscuit.

1. Clamp the boiling tube as shown in figure 11.1
2. Use a measuring cylinder to transfer $20cm^3$ of water to the boiling tube.
3. Put a thermometer into the water and take the temperature of the water.

4. Fasten the first food sample to a long needle and set the food alight with a Bunsen burner flame.
5. Immediately hold the flame under the boiling tube. Stir the water with a thermometer.
6. Record the highest temperature reached.
7. Repeat the procedure with each of the foods.
8. The food that produces the largest temperature rise will have released the most energy.

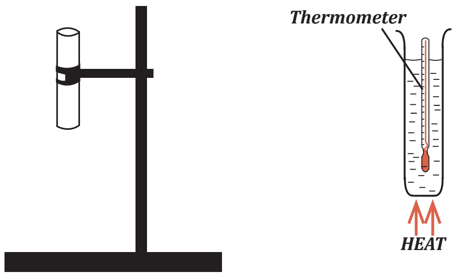


Figure 11.1; comparing the energy values of food.

Consider the table below which shows energy content of some foods.
Table 11.1: energy content of some food stuffs

Food	Energy content (Kj/100g)
Fish	300
Butter	3000
Fresh fruit	200
Bread	1000

It can be seen from the above table that foods that contain carbohydrates (e.g. bread), and fats (e.g. butter) are the ones which contain greatest amounts of stored energy.

Activity 11.2; To carry out destructive distillation of wood in the laboratory.
Procedures;

1. Set up the apparatus as shown in figure 11.2
2. Fill the beaker with cold water and insert in it a test tube containing a mixture of acid liquor and wood tar as shown in figure 11.2. Do not insert the trough in the position at this stage.
3. Heat the saw dust gently and light the wood. The gas escapes from the jet.
4. If no gas is escaping try to heat more strongly.

5. Fill in the trough and collect the wood gas over water as shown in figure 11.2. Collect several gas jars if possible.
6. Go on heating strongly; when there is no further change, remove the trough of water.
7. Remove the source of heat and allow the apparatus to cool.
8. Examine the products, and test them in second test tube with blue litmus paper.

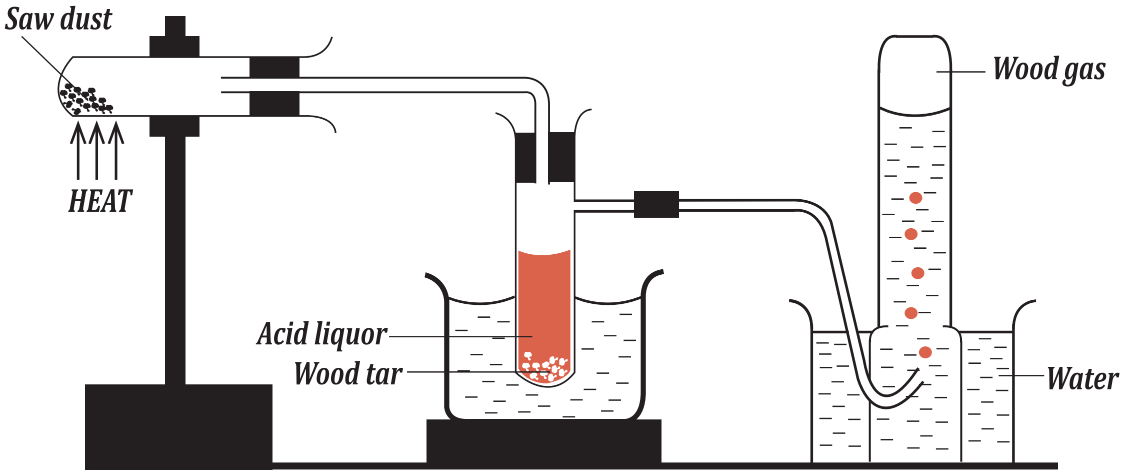


Figure 11.2; destructive distillation of wood.

Destructive distillation of wood can be carried out locally. Consider the following diagram

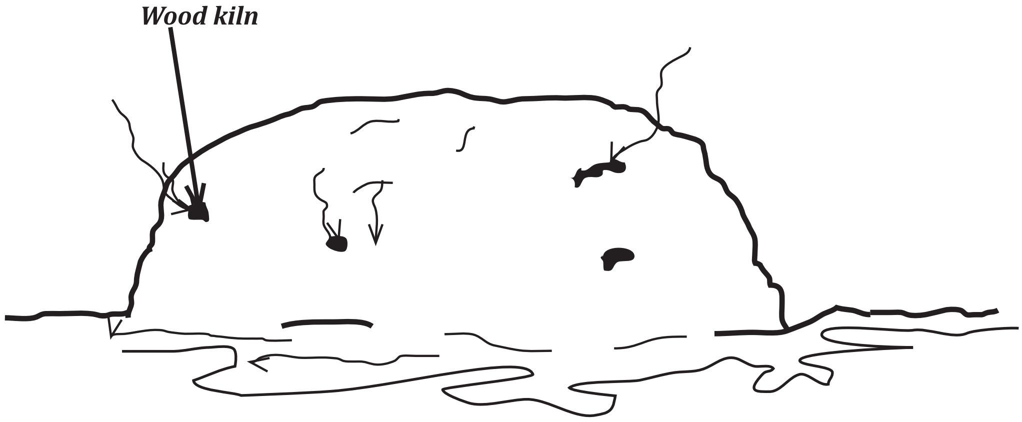


Figure 11.3; local method of destructive distillation of wood.

Discussion questions

1. What fuels do you use at home? Can you find out the different costs of fuels?
2. What fuel do you think is most efficient for use at home?

11.2 CATEGORIES OF FUELS

Fuels can be categorized into three states of matter as gaseous fuel, liquid fuel and solid fuel. The efficiency of a fuel can be explained in terms of the amount of heat energy that can be produced when that fuel is completely burned.

An efficient fuel produces a lot of heat energy when a small amount is used. It has also a smooth combustion which does not produce a lot of gaseous wastes into the environment. Example kerosene produces very hot flame and less soot when used at home. Fire woods burn slowly and produce smoky flame. The heat produced is not high. Charcoal is more efficient than firewood and produces reliable heat energy.

In industries, coke and carbon monoxide are important and efficient fuels as they produce much heat energy. Gas fuels like biogas, coal gas and ethylene are important and efficient fuels for large production of goods in industries. For classification, gaseous fuels are more efficient, followed by liquid fuels and lastly the least efficient are solid fuels.

Fuels can also be categorized on the basis of their occurrence, fuels can be classified into natural (primary) and artificial (secondary) fuels. Natural fuels occur in nature. Examples are; wood, coal, peat, petroleum and natural gas.

Artificial fuels are either manufactured in industries or derived from the primary fuels by refining. Examples are coke, kerosene, petrol, coal gas and producer gas.

Consider the following table.

Table 11.2 categories of fuels according to the physical states.

Solid fuels	Liquid fuels	Gas fuels
Coal, charcoal and fire wood	Diesel, kerosene and petrol	Acetylene for welding, water gas and producer gas for industries, natural gas and town gas for domestic uses.

11.3 USES OF FUELS

The following is the list of uses of fuels by human being in several activities:

1. Fuels are used to run machines in industries and motor vehicles. Example petrol, diesel and coke.
2. Fuels are used for cooking, boiling, and warming homes. Example charcoal, kerosene, fire wood, coke and coal gas.
3. Fuels are used for lighting at home especially in rural areas where electricity is unavailable. Example kerosene.
4. Fuels are used for drying. Example, tobacco leaves are dried in kilns by burning woods.
5. Fuels are used for rocket propulsion and space ships travels. Example hydrogen and uranium or plutonium.

EFFECT OF FUELS ON ENVIRONMENTAL POLLUTION.

The fuels used mostly nowadays release harmful gases into the air that cause pollution. Living things are being affected by polluted air. Acidic rains, green house effect and health effects are all that can result due to animals come in contact with polluted air. When burning coal and petrol, gaseous pollutants like sulphur dioxide, and carbon monoxide can be released into the atmosphere. These gases lead to acidic rain when dissolved in rain water. Green house effect can result after releasing of several gases into the atmosphere. Example carbon dioxide and chlorofluoro carbons. In view of that, we must consider choosing a fuel to use in order to avoid environmental pollution. In addition to that, extensive harvest of plants for firewood and charcoal production would lead to effects like; desertification, disturbance on water cycle, destruction of homes for wild life and eventually the ecosystem.

In order to avoid environmental hazards due to deforestation and pollution an alternative source of fuel must be encouraged and this means that renewable energy sources emphasized.

11.4 CONSERVATION OF ENERGY.

There are different forms of energy. Energy can be potential or kinetic. Potential energy is energy at rest while kinetic energy is energy in motion.

Energy in motion is illustrated by winds, running water, ocean waves, moving machines or falling bodies. Potential energy is stored in different forms. E.g. Petroleum, coal, natural gas, muscles and batteries. Such energy does not do work when it is stored.

The potential energy is capable of doing work when it is changed to other forms of energy. Example heat and light. All energy changes which occurs during chemical and physical changes must conform to the law of conservation of energy. The law of conservation of energy states that.

> ***" Energy can neither be created nor destroyed but it can be converted from one form to another".***

Example; the energy in a fuel is not destroyed when it burns but it simply changes to other forms like heat and light.

CONVERSION OF ENERGY.

Energy can be converted from one form to another as it can be listed below.

1. Conversion of solar energy into light energy by using solar panel
2. Conversion of sound energy into electrical energy. Example using a microphone.
3. Conversion of electric energy into sound energy. Example radio speaker.
4. Conversion of electric energy into light energy. Example electric bulb.

5. Conversion of mechanical energy into sound energy. Example drum beating.
6. Conversion of chemical energy into mechanical energy. Example muscles.
7. Conversion of chemical energy into electrical energy. Example battery.

Activity 11.3 to convert chemical energy into electrical energy.
Procedure.

1. Clean pieces of copper and magnesium ribbon with sand paper.
2. Connect the battery and bulb with the magnesium and copper using connecting wires with crocodile clips.
3. Dip the two pieces of metal into a beaker containing dilute sulphuric acid. You should see the bulb light up for a short time, showing that chemical energy is transferred into electrical energy.

11.5 BIOGAS AS RENEWABLE ENERGY

Biogas refers to a gas produced by breakdown of organic matter in the absence of oxygen. Organic wastes such as dead plant and animal material, animal faeces and kitchen waste can be converted into a gaseous fuel called biogas. Biogas is produced by the anaerobic digestion of biodegradable materials such as biomass, manure, sewage, municipal waste, green waste, plant material and crops.

Biogas comprise primarily methane (CH_4) and carbon dioxide (CO_2) and may have small amount of hydrogen sulphide (H_2S), moisture and siloxanes.

Biogas is a renewable energy which is environmentally friendly. In creating biogas the air should be absent and water present. The organic matter being cattle dung, poultry wastes, night soil, crop residues, food processing and paper wastes. Residue from the agricultural sector, such as spent straw, hay, and cane trash need to be cut into small pieces in order to facilitate their flow into the digester reactor as well as to increase the efficiency of bacterial action. Succulent plant material yields more gas than dry matter does.

THE WORKING MECHANISM OF A BIOGAS PLANT.

The biogas plant consists of two components; A digester or fermentation tank and a gas holder.

The digester is a cube-shaped or cylindrical water proof container with an inlet into which the fermentable mixture is introduced in the form of a semi-liquid mixture of fine particles of manure.

The gas holder is normally an air proof steel container that, by floating like a ball on the fermentation mix, cuts off air to the digester and collect the gas generated.

The organic waste matter is generally animal or cattle dung. These waste products contain carbohydrates, proteins and fat material that are broken down by bacteria. These materials are soaked in water to give the bacteria a proper medium to grow. Absence of air or oxygen is important because bacteria then take oxygen from the waste material itself and in the process break them down.

The fermentation materials should not be too diluted or too concentrated, because this may lead to low biogas production and insufficient fermentation activity respectively. The ratio of raw material to water should be 1:1 i.e. 1500 kg of excrete to 1500 kg of water. The loading of the raw material in the plant is determined by the influent solid contents, retention time, and the digester temperature. High loading is usually used in high ambient temperature. The plant is seeded with adequate population of both the acid forming and methanogenic bacteria. These accelerate the process of digesting sludge. Sometimes addition of lime or other alkali is done in case the overloading of the seeding material is done. For the successive action of the bacteria on the sludge, the pH in the plant must be maintained between 6.0 – 8.0. This is because efficient digestion occurs at a pH near neutrality. The temperature should be maintained at a range that usually favours the digestion process. At high temperature biodigestion occurs faster, reducing the time requirement. A normal period for the digestion of dung would be two to four weeks.

Consider the below diagram which shows the biogas plant.

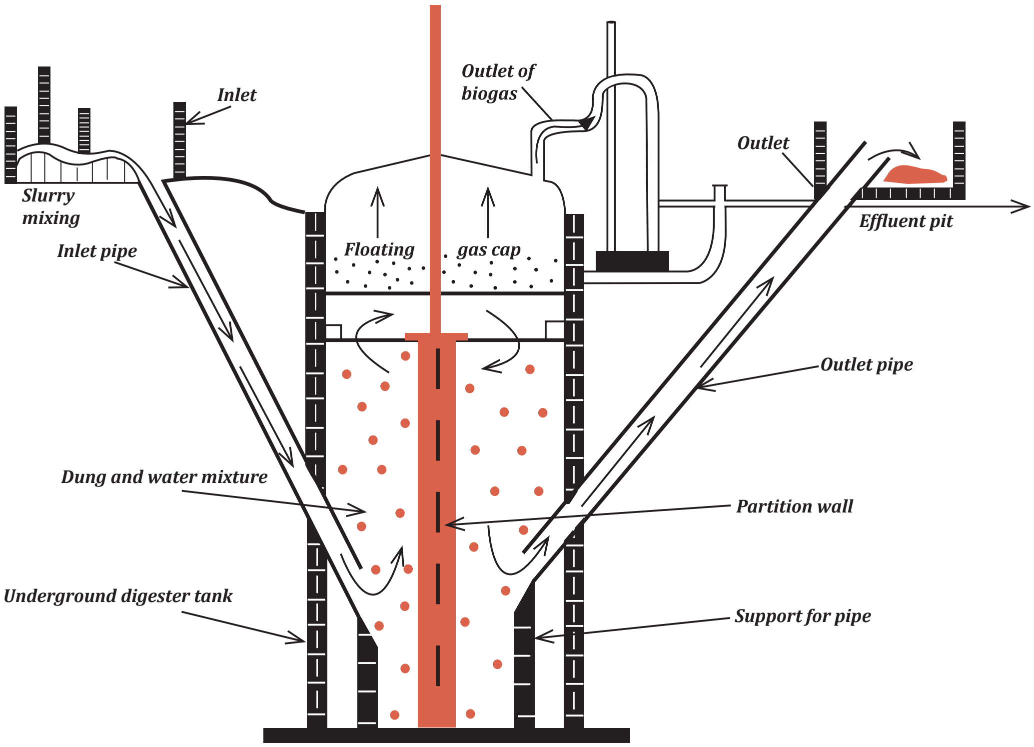


Fig. 11.4 the biogas plant

Activity 11.4 A class to construct a small scale model of biogas plant.

Materials needed:

Biogas plant model, water, pipes, concrete, sewage, cow dung, animal and plant waste.

Procedures

1. Select a place outside the classroom to construct the biogas plant
2. Using the materials you have collected construct a model of biogas plant. Consider figure 11.4
3. The teacher to assist students in each step to construct the biogas plant.
4. Feed the biogas plant with organic residues and let them decompose for 2 – 3 weeks.
5. After three weeks collect the gas and demonstrate its uses.

Discussion questions.

1. What problems did you face during construction of the model?
2. How efficient was the gas generated?

THE USE OF BIOGAS AND ENVIRONMENTAL CONSERVATION.

When biogas is used, many advantages arise. Utilization of biogas would generate enough electricity and also could potentially help reduce global climatic change. Biogas is a combustible gas, that when used produce no air pollutants on burning. Wastes and biodegradable residue may be accompanied by a variety of pollutants, but this can be corrected by remedying the carbon/nitrogen ratio of manure through the addition of shredded bagasse or straw or by dilution.

Consider the below table which explains what 1 cubic meter of biogas can do.

Table 11.3 application of 1 cubic metre of biogas.

Application	1 cubic meter of biogas
Application	One 100w light bulbs for 6 hours
Cooking	Cooks 3 meals for family of 5-6 people
Fuel replacement	0.7 kg of petrol
Power	Can run a 1 horse power motor for 2 hours
Electricity generation	Can generate 1.25 kwh of electricity

SUMMARY

(a). We use energy everyday in homes, industries, schools and in our bodies
(b). A fuel is a store of energy and when a fuel is used energy is released.
(c). Energy is measured in the unit of Joule (J)
(d). Fossil fuels include coal, natural gas and petroleum
(e). When a fuel releases energy it usually does not release all of the stored energy. Much of the energy is wasted.
(f). The efficiency of the energy source is the ratio of the useful energy released to the total amount of energy stored.
(g). Energy can never be created nor destroyed but it can be converted from one form to another.
(h). Biogas is the gas produced by the fermentation of organic matter, including manure and other solid wastes under anaerobic conditions i.e. absence of air.
(i). A fuel can be in any of the three states of matter.
(j). Sources of fuels can be renewable or non – renewable.

REVIEW QUESTIONS.

1. Choose the most correct answer for the following questions.
   i. The process of heating wood in the absence of air is called……………
      A. Destructive distillation B. Destructive condensation
      C. Distillation D. Energy conversion
   ii. A fuel can be in …………….. states of matter
      A. Two B. Four C. Three D. Five
   iii. …………… affects living things in different ways.
      A. Atmospheric air B. Polluted air
      C. Non-poisonous gases D. Fuel
   iv. ………… comes from the heat energy present inside the earth.
      A. Solar energy B. Wind energy
      C. Electric energy D. Geothermal energy.

2. Match the items in list A with their corresponding statements in LIST B.

List A	List B
i. Most efficient category of fuel……	A. Coal, charcoal and fire wood.
ii. An example of renewable resources……	B. Wood
iii. Non renewable energy source…...	C. Natural gas and uranium
iv. Fossil fuel………	D. Remains of plants and animals that lived millions of years ago
v. Solid fuels………………	E. Gas fuel
	F. Liquid fuel.

3. Write ***TRUE*** for a correct statement and ***FALSE*** for an incorrect statement.
   i. Good fuel should have high heat content…………….
   ii. Good fuel should be very expensive. …………….
   iii. Good fuel should not be easily controlled…………….
   iv. Burning of coal and petrol produces gaseous pollutants like sulphur dioxide gas into the air. …………….
   v. Deforestation leaves the land bare to agents of erosion like moving water, animals and wind…………….
   vi. Solar energy is available in plenty and cheap…………….
   vii. Renewable energy is an environmental friendly because it does not pollute the environment…..
   viii. Potential energy is energy in motion while kinetic energy is energy at rest.........
   ix. Biogas is a combustible gas that when used produce no air pollutants on burning……………
   x. Use of biogas is cost effective and environmental friendly…………….

4. (a) What do you understand by renewable energy?
   (b) What is the importance of renewable energy source?
   (c) What do you understand by biogas plant?
   (d) Mention organic materials that can be used in biogas plant.

5. (a) State the law of conservation of energy.
   (b) Mention at least four ways by which energy can be converted from one form to another.

6 (a) State the classes of fuels according to their physical states. Give three examples in each case.
   (b) Classify fuels according to their efficiency.

7. How do you obtain fuel locally from your environment?

8. Describe an experiment that could be used to compare the amount of energy released when wood, charcoal and coal are burned.

9. The table below gives the energy stored in different fuels, A-E and the energy released when these are burned.

(a) Copy and complete the table.

(b) Which fuel has the greatest efficiency?

(c)The same mass of fuel is burned in each case. Which of these statements is true and which is false?

- The efficiency depends upon the amount of fuel burned.
- The more fuel burned the greater the energy released.

Fuel	Mass of sample of fuel	Energy stored in fuel	Energy released by fuel	Efficiency%
A	100g	1250kJ	450 kJ	36
B	100g	1500 kJ	600 kJ	
C	100g	1000 kJ	450 kJ	
D	100g	1800 kJ	450 kJ	
E	100g	2000 kJ	400 kJ	

Chapter 12

Atomic structure

12.1 THE CONCEPT OF ATOM

The name ***atom*** comes from the Greek word ***"Atomos"*** which means ***"Indivisible"***. The two terms of ***"Atomos"*** in Greek are ***"a"*** which means ***"not"*** and ***"Temno"*** which means "cut". As early as 400 BC some people had thought about the structure of matter. Democritus, an ancient Greek thinker, suggested that all matter was made of particles. However, he had no scientific evidence for this. More powerful people like Aristotle disagreed with Democritus and so the ideas were largely forgotten.

By definition, an atom is the smallest particle of an element that can take part in a chemical reaction. The concept of an atom was first studied by John Dalton in his theory named Dalton atomic theory.

DALTON ATOMIC THEORY.

In 1803, Dalton developed his theory about the atom. The five main points of Dalton's atomic theory are;

1. All matter is made up of tiny indivisible particles called atoms.
2. Atoms can neither be created nor destroyed.
3. Atoms of any one element are identical and have the same mass.
4. Atoms of different elements have different masses.
5. Atoms may combine to form compounds, but the combination must involve whole numbers of atoms.

Dalton's atomic theory was the first step towards the formulation of modern atomic theory. Dalton's theory has been subjected to a lot of experimentation. This has led to slight modifications to Dalton's theory. The modifications are;

1. Atoms can be created or destroyed or split by means of nuclear reactions
2. Atoms of the same element may have different masses. These are called ***'Isotopes'***, the existence of which is a challenge to Dalton's ideas. Some atoms of the same element can be quite different in their mass.
3. An atom is made up of even smaller sub atomic particles called protons,

neutrons and electrons.

4. Atoms of different elements may combine in many different ways. Example the formation of carbon dioxide from carbon and oxygen.

12.2 SUB ATOMIC PARTICLES

The early years of the twentieth century saw many scientists working to establish the structure of the atom. Prominent names in this area are; Thomson, Rutherford, Chadwick, Bohr and many others. An atom is now thought to consist of a very small and dense region, called the nucleus, surrounded by orbiting electrons. The main sub atomic particles are protons, neutrons and electrons.

THE ELECTRON.

This is a particle with a unit of negative charge. It has a very small mass, about 1/1800 that of the proton. Its symbol is $^{0}_{-1}e$. At this basic level, we shall assume that the electrons orbit around the nucleus in electron shells at high speeds.

THE PROTON

This is a particle which carries a unit of positive charge. It has a mass approximately the same as that of hydrogen atom, i.e. one atomic mass unit (amu). Protons and neutrons form the nucleus of an atom. The symbol of proton is $^{1}_{+1}P$

THE NEUTRON

Neutron carries no charge at all. It is represented by letter $^{1}_{0}n$. This mass of neutron is the same as that of proton. The table below gives a summary of the properties of sub – atomic particles of an atom.

Table 12.1 Properties of sub – atomic particles.

Sub atomic particle	Symbol	Location	Charge	Real mass (g)	Relative mass
Proton	$^{1}_{+1}P$	In the nucleus	$^{+}1$	1.6726×10^{-24}	1
Neutron	$^{1}_{0}n$	In the nucleus	0	1.6750×10^{-24}	1
Electron	$^{0}_{-1}e$	Outside the nucleus	$^{-}1$	9.109×10^{-28}	$^{1}/_{1840}$

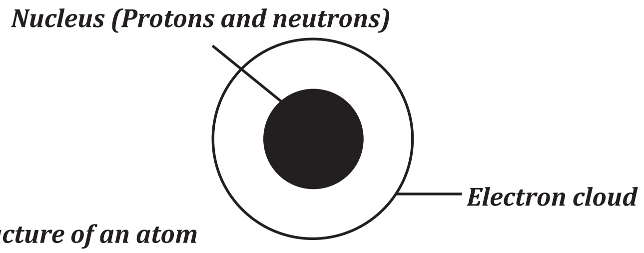


Fig.12.1; structure of an atom

12.3 THE ARRANGEMENT OF ELECTRONS IN AN ATOM

Electrons in an atom move in the space around the nucleus and are at different distance from it. Electrons are arranged in shells called ***energy levels***. An energy level is considered as a region around the nucleus most frequently occupied by electrons of approximately the same energy. Each shell can hold only a certain number of electrons. Each shell can hold a maximum number of electrons given by the formula $2n^2$, where n is the number of the shell starting from that nearest to the nucleus.

According to the formula, the first, second and third shells can contain a maximum of 2,8 and 18 electrons respectively. The arrangement of electrons around the nucleus is also known as the ***electronic configuration*** in which electrons are arranged in different energy levels.

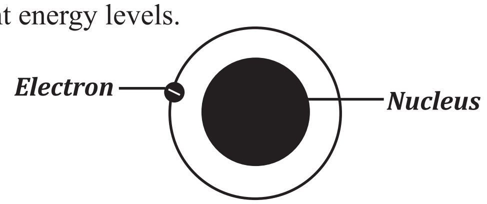


Fig.12.2 Electron diagram of hydrogen

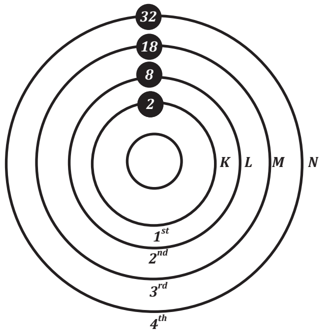


Fig.12.3 Electron shells

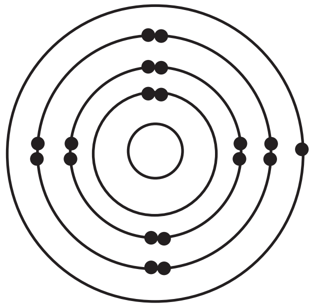


Fig 12.4 Electronic arrangement of an element With atomic number 19

Table 12.2 electron arrangement of the first 20 elements.

Element	Symbol	Number of electrons	Number of shells	Electronic configuration	Period	Group
Hydrogen	H	1	1	1	1	
Helium	He	2	2	2	1	0
Lithium	Li	3	2	2:1	2	1
Beryllium	Be	4	2	2:2	2	2
Boron	B	5	2	2:3	2	3
Carbon	C	6	2	2:4	2	4
Nitrogen	N	7	2	2:5	2	5
Oxygen	O	8	2	2:6	2	6
Fluorine	F	9	2	2:7	2	7
Neon	Ne	10	2	2:8	2	8
Sodium	Na	11	3	2:8:1	3	1
Magnesium	Mg	12	3	2:8:2	3	2
Aluminium	Al	13	3	2:8:3	3	3
Silicon	Si	14	3	2:8:4	3	4
Phosphorus	P	15	3	2:8:5	3	5
Sulphur	S	16	3	2:8:6	3	6
Chlorine	Cl	17	3	2:8:7	3	7
Argon	Ar	18	3	2:8:8	3	0
Potassium	K	19	4	2:8:8:1	4	1
Calcium	Ca	20	4	2:8:8:2	4	2

Activity 12.1 To make the model of electronic diagram.

Materials: Flexible material (wire or thin stick), threads, clay soil or any 3 soft round small objects or fruits with different colours, water and gloves.

Procedures:

1. Make clay balls with a hole, or if you have soft round small objects or fruit make a hole at the centre.
2. Put one on a flexible material (example wire or thin stick) like beads on its thread.
3. Bend that material slowly to make a circle and tie it with a thread or plastic string.
4. Make another clay ball and on it put other small balls one at the top another at the bottom
5. Put a circle you have made on the table or ground and at the centre of it put the clay ball you have prepared in (4) above. Draw that model. Repeat the

exercise but this time making separate models as follows.

a. One circle with two balls on it and clay ball in (4) at the centre.
b. A model with two circles, inner circle (2 balls) and the outer (6 balls) and one clay ball prepared in stage (4) at the centre.
c. A model with three circles, inner circle (2 balls). As normal put a clay ball prepared in stage (4) at the centre. Lay them on the ground and draw them.

Discussion questions.

1. What is the name of the diagram(s) you have drawn?
2. Basing on that diagram, what do
   a. Balls on the string or soft wire represent?
   b. String or wire represent?
   c. Clay ball at the centre represent?

Activity 12.2

Draw the electronic configuration and electronic diagrams of the following atoms

a. Neon with atomic number 10
b. Sodium with atomic number 11
c. Hydrogen with atomic number 1
d. Magnesium with atomic number 12
e. Chlorine with atomic number 17

Activity 12.3

Students to role play on electronic arrangement of an atom. In this activity students will act as electrons and the teacher as the nucleus of an atom.

Procedures:

1. Students are assigned K,L,M and N shells.
2. The teacher should be at the centre and the students should go around in their respective orbits (or shells)
3. Students should touch one another by their hands, so that they form cycles around the teacher.
4. Students now in their shells, extend outwards as possible as they can.
5. Now the teacher should not move. Let students start moving around the teacher in their respective shells. Relate this with electrons in their respective shells.

Discussion questions.

1. What was the aim of the activity?
2. How many electrons are found in the K,L,M and N shells?
3. What was the aim of a teacher to be at the centre?

12.4 ATOMIC NUMBER, MASS NUMBER AND ISOTOPES

ATOMIC NUMBER (Z)

Atomic number is the number of protons in the nucleus of an atom. It is also the number of electrons in a neutral atom. It is also part of the atomic structure which identifies the element. Example the atomic number of hydrogen is 1 since it has only one proton. A magnesium atom has 12 protons in the nucleus and therefore its atomic number is 12.

Since the number of protons is equal to the number of electrons in the atoms, the atomic number also indicates the number of electrons in the atom.

The official symbol for atomic number is Z. Atomic number is usually written as a subscript in front of the symbol for the atom. Example Chlorine, where Z = 17, is written as $_{17}Cl$.

MASS NUMBER (A)

Mass number is the sum of the numbers of protons and neutrons in the nucleus of an atom. The mass number is usually written as a superscript infront of the symbol for the atom. Example Nitrogen atom, where A=14 is written ^{14}N. The official symbol for mass number is A.

It is possible to get the number of neutrons and number of electrons of an atom if its mass number and atomic number are given.

Example 12.1

Atom M has a mass number of 16 and an atomic number 8. What is the neutron number? What is the number of electrons in atom M?

Solution.

Mass number = 16, atomic number = 8.

(a) Neutron number = Mass number – Atomic number
= 16 – 8 = 8

(b) Number of electrons = number of protons
= Atomic number = 8

Example 12.2

Complete the table below by indicating the number of protons, electrons and neutrons of the atoms shown. The atomic number and mass numbers are given;

Atom	Atomic number	Mass number	Protons	Electrons	Neutrons
Sulphur	16	32			
Chlorine	17	35.5			
Argon	18	40			
Potassium	19	39			
Calcium	20	40			

NUCLIDE NOTATION.

Atoms of different elements can be represented by symbols that indicate their respective atomic numbers and mass numbers. Using an arbitrary element X, the mass number (A) is placed on its upper left end while its atomic number (Z) is placed on the lower left end. Then element X is represented as ${}^{A}_{Z}X$. This is known as nuclide notation..

Example 12.3

You are given an atom ${}^{39}_{19}X$

a. Write its electronic configuration.
b. Identify the number of neutrons and protons in the atom.
c. Draw its electronic diagram.
d. Identify the element X by its symbol and give the nuclide notation.
e. Show the representation of the nucleus of the atom.

Solution:

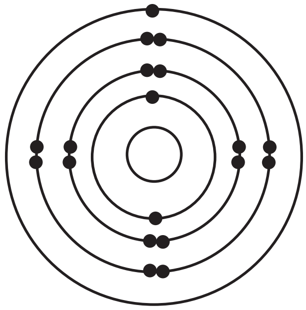
(a) Electronic configuration 2:8:8:1

(b) The number of neutrons = Mass number – Atomic number.

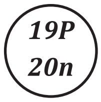
= 39 – 19 = 20.

Number of protons = atomic number = 19.

(c) Electronic diagram



(d) Element X is potassium; K and its nuclide notation is $^{39}_{19}K$.

(e) 

ISOTOPES

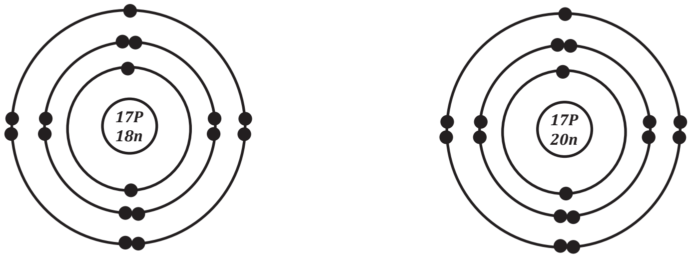
Isotopes are atoms of the same element with the same number of protons but different number of neutrons. Isotopes have the same number of protons, the same number of electrons and the same atomic number. They have the same chemical properties but may have slightly different physical properties. Isotopes have different mass numbers because they have different numbers of neutrons.

Chlorine for example has two common isotopes which are explained in the table below.

Table 12.3 isotopes of chlorine.

Element	Atomic number	Number of protons	Number of neutrons	Atomic mass
Chlorine	17	17	18	35
	17	17	20	37

The isotopes of chlorine can thus be written as $^{35}_{17}Cl$ and $^{37}_{17}Cl$ which are known as chlorine -35 and chlorine -37 respectively.



Chlorine -35 ***Chlorine -37***

Fig. 12.5; Isotopes of chlorine

Having learned about Isotopes, it is also important to understand the concept of Isotopy. Isotopy is the tendency of atoms of the same element to contain the same number of protons but different number of neutrons or mass number. Examples of common elements that display isotopy are hydrogen, oxygen, chlorine and neon.

Consider the below table showing common isotopes.

Table 12.4 Examples of common isotopes.

Element	Symbol	Atomic number	Isotopes	Abundance
Hydrogen	H	1	$^{1}_{1}H$ (Protium)	99.99%
			$^{2}_{1}H$ (Deuterium)	0.01%
			$^{3}_{1}H$ (Tritium)	very rare
Carbon	C	6	$^{12}_{6}C$	98.9%
			$^{13}_{6}C$	1.1%
			$^{14}_{6}C$	trace
Chlorine	Cl	17	$^{35}_{17}Cl$	75%
			$^{37}_{17}Cl$	25%
Oxygen	O	8	$^{16}_{8}O$	99.8%
			$^{17}_{8}O$	0.37%
			$^{18}_{8}O$	0.20%
Neon	Ne	10	$^{20}_{10}Ne$	90.5%
			$^{21}_{10}Ne$	0.3%
			$^{22}_{10}Ne$	9.2%

RELATIVE ATOMIC MASS (R.A.M)

The relative atomic mass of an element is the average mass of one atom of the element relative to $\frac{1}{12}^{th}$ the mass of one atom of carbon -12.

$$\text{R.A.M} = \frac{\text{Average mass of atom of an element}}{\frac{1}{12}^{th}\ \text{the mass of carbon -12 atom.}}$$

The relative atomic mass has no units.

The carbon atom was chosen as the standard atom (reference atom) and its mass was arbitrarily chosen as 12 units. Then using a machine called mass spectrometer, all the other atoms were compared to this standard atom. This reference is called the ***carbon -12 scale***. For example it was found that:

i. The magnesium atom was twice as heavy as the reference atom, so its mass was put at 24.
ii. The hydrogen atom was 1/12 as heavy as the reference atom, so its mass

was put at 4.

iii. The helium atom was 1/3 as heavy as the reference atom, so its mass was put at 4.

Example 12.4

A sample of chlorine gas contains 75% of the isotope $^{35}_{17}Cl$ and 25% of the other isotope, which is $^{37}_{17}Cl$. What is the relative atomic mass of chlorine?

Solution

Relative atomic mas (R.A.M) is given by the formula

$$\textbf{R.A.M} = \left(\text{Mass number of an Isotope} \times \text{Relative abundance for an isotope}\right) + \left(\text{Mass number of an Isotope} \times \text{Relative abundance for an isotope}\right)$$

$$\textbf{R.A.M} = \left(35 \text{ x } \frac{75}{100}\right) + \left(37 \text{ x } \frac{25}{100}\right)$$

$$= \frac{(75 \text{ x } 35) + (25 \text{ x } 37)}{100} = 35.5$$

Therefore, the R.A.M of chlorine is 35.5

Example 12.5

Carbon has two main isotopes $^{13}_{6}C$ and $^{12}_{6}C$ with abundance of 1.11% and 98.89% respectively. Calculate its relative atomic mass (R.A.M).

Solution

$$\textbf{R.A.M} = \left(13 \text{ x } \frac{1.11}{100}\right) + \left(12 \text{ x } \frac{98.89}{100}\right)$$

$$= \frac{(13 \text{ x } 1.11) + (12 \text{ x } 98.89)}{100} = 12.1$$

Therefore, the R.A.M of Carbon is 12.01

Example 12.6

Neon has three isotopes; $^{20}_{10}Ne$ (90.5%), $^{21}_{10}Ne$ (0.3%) and $^{22}_{10}Ne$ (9.2%). Calculate the relative atomic mass of neon.

Solution.

$$\textbf{RAM} = \left(20 \times \frac{90.5}{100}\right) + \left(21 \times \frac{0.3}{100}\right) + \left(22 \times \frac{9.2}{100}\right)$$

$$= \frac{(20 \times 90.5) + (21 \times 0.3) + (22 \times 9.2)}{100} = 20.187$$

Therefore, atomic mass of Neon is 20.187.

SUMMARY

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

(b). The concept of atom was first studied by John Dalton in his theory named Dalton atomic theory.

(c). The three sub atomic particles are protons, neutrons and electrons.

(d). The arrangement of electrons around the nucleus is also known as electronic configuration.

(e). Each shell in an atom can hold a maximum number of electrons given by the formula $2n^2$

(f). Atomic number is the number of protons in the nucleus of an atom.

(g). The mass number is the sum of the numbers of protons and neutrons in the nucleus of an atom.

(h). Isotopes are atoms of the same element with the same number of protons but different number of neutrons.

(i). Relative atomic mass of an element is the average mass of one atom of the element relative to 1/12th the mass of one atom of carbon -12.

REVIEW QUESTIONS.

1. Choose the best answer for each of the following questions.

(i) The smallest particle of an element that can take part in a chemical reaction is called.

A. Matter B. Substance C. Atom D. Element.

(ii) The first modern scientist to come up with ideas about building blocks of matter which he gave the name "atoms" was……………..

A. John Dalton B. Michael Faraday

C. Mendelēev D. Thomas Edison.

(iii) The three sub – atomic particles are

A. Electron, nucleus and protons. B. Protons, nucleus and electrons.

C. Electrons, nucleons and protons. D. Protons, neutrons and electrons.

(iv) Electronic arrangement in atoms is also known as

A. Electronic configuration B. Electronic arrangement

C. Electronic shells D. Electron cloud.

(v) Is the tendency of atoms of the same element to contain the same number of protons but different number of neutrons or mass number.

A. Isotopes B. Isotopy C. Allotropy D. Atomic number.

2. Match the items in List A by writing a letter of the correct statement of List B in the space provided.

LIST A	LIST B
i. Proton	A. $^{37}_{17}Cl$ and $^{35}_{17}Cl$
ii. Neutron	B. Negatively charged particle of an atom
iii. Electron	C. Positively charged particle of an atom.
iv. Carbon-12	D. Neutral charged particle of an atom.
v. Isotopes	E. The reference for relative atomic mass.
	F. Partially charged particle of an atom.

3. Write TRUE for a correct statement and FALSE for an incorrect statement.

i. Atoms can be created or destroyed or split by means of nuclear reactions……………….

ii. An atom is made up of even smaller sub atomic particles called protons, neutrons, and electrons.

iii. The atoms always combine in simple ratios……………….

iv. The atoms of any one element are identical and have the same chemical properties and the same mass………………..

v. Dalton never imagined that any one would ever be able to see an atom.

4. (a) What is electronic configuration?
   (b) An atom has mass number of 24 and neutron number 12;
      (i) What is its atomic number.
      (ii) Write down its electronic configuration.

5. Describe the changes in properties across the period from sodium to argon.
6. A certain element P, was found by mass spectrometry – a sophisticated technique for measuring the relative masses of element or compounds – to have Isotopes.

^{204}P , ^{206}P , ^{207}P and ^{208}P .

Their relative abundance were, respectively 2,24,22 and 52, as determined by a mass of the element P, can you suggest a name for the element? (P is not the real symbol of the element).

7. Draw the structure of an atom according to
   (a) Dalton (b) Rutherford.

8. State the number of protons, neutrons and electrons in the following Isotopes.

(i) $^{12}_{6}C$ (ii) $^{14}_{6}C$ (iii) $^{3}_{1}H$ (iv) $^{87}_{38}Sr$ (v) $^{238}_{92}U$

9. An Isotope of carbon has a mass number of 13 and an atomic number of 6.
   (i) Write its nuclide notation. (ii) How many neutrons does it have?
   (iii) How many electrons does it have?

10. Copy and complete the table below;

Element	Atomic number	Relative atomic mass	Number of protons	Number of neutrons
Hydrogen	1	1		
Helium	2	4		
Lithium		6.9		
Beryllium	4	9		
Boron	5	10		
Carbon	6	12		
Nitrogen		14		
Oxygen		16		
Fluorine	9	19		
Neon	10			

Element	Atomic number	Relative atomic mass	Number of protons	Number of neutrons
Sodium	11	23		
Magnesium	12			
Aluminium	13	27		
Silicon				
Phosphorus	15	31		
Sulphur	16	32		
Chlorine	17			
Argon	18			
Potassium	19			
Calcium	20			

Chapter 13

Periodic classification

Periodic classification is the systematic grouping of elements into periods and groups in the periodic table according to their similarities and differences in electro negativity, ionization energy, melting point, boiling point, density, atomic radius and reactivity,metallic character and non - metallic character.

Many elements were known since the ancient times but up to the middle of seventeenth century, the ideas of elements were disordered and the elements classification did not exist.

13.1 THE PERIODIC TABLE

Periodic table is the arrangement in which all known elements are arranged in their respective periods and groups in order of increasing atomic number. The first periodic table was devised by the Russian chemist Mendelēev in 1869. He devised it to help his students understand chemistry

Mendel ēev did the following:

- He arranged the known elements in groups according to their properties in order of increasing atomic mass, but left gaps for undiscovered elements.
- He predicted the properties of the missing elements from the properties of the elements above and below them in the table.
- He proposed a periodic law, which states

  "The properties of elements are a periodic function of their atomic masses"

When elements were arranged in order of increasing atomic weights, the properties of elements fell into repeating patterns. The vertical columns of similar elements were called groups. The horizontal rows were called periods.

- He listed separately (in group 8) some elements which did not appear to fit into any group i.e. iron, cobalt, nickel etc.

Table 13.1; Mendelēev’s periodic table

Group	I		II		III		IV		V		VI		VII		VIII
	A	B	A	B	A	B	A	B	A	B	A	B	A	B	
Period 1		H													
Period 2	Li		Be		B		C		N		O		F		
Period 3	Na		Mg		Al		Si		P		S		Cl		
Period 4	K		Ca				T		V		Cr		Mn		Fe,Co,Ni
Period 5		Cu		Zn						As		Se		Br	
Period 6	Rb		Sr						Nb		Mo				Ru,Rh,Pd
Period 7	Ag		Cd		In		Sn		Sb		Te		I		

Limitations of Mendelēev’s classification;

- The appearance of two families in one group, e.g. K and Cu, Ca and Zn.
- Some elements had to be placed in reverse order of their atomic weight so as to fit into groups having similar properties. (e.g. Ar and K, I and Te, Co and Ni)

MODERN PERIODIC TABLE

The modern periodic table of elements is a table of elements arranged systematically according to their ***increasing atomic numbers***. It is a result of several modifications to Mendelēev’s periodic table. The modifications were made as new elements were discovered and new theories developed to explain the chemical behavior of elements.

The ***modern periodic law*** states that;

The properties of elements are a periodic function of their atomic numbers.

13.2 PERIODICITY

Periodicity is the regular periodic changes of elements due to their atomic number.

FEATURES OF THE MODERN PERIODIC TABLE.

The main features of the modern periodic table are;

- The elements are placed in order of their atomic numbers instead of their atomic masses.
- A new group of ‘noble gases’ has been discovered and included on the right hand side of the table as group O.
- Gaps in the table left by Mendelēev, have been filled as new elements e.g. germanium have been discovered. Man-made elements have also been accommodated in the table.
- Some groups, have names as well as numbers. Group I is called the alkali

metals. Group II is called the alkaline earth metals. Group (VII) is called the halogens. Group O is called the noble gases.

- Metals have been clearly separated from non-metals.
- There are five blocks of similar elements in the modern periodic table.

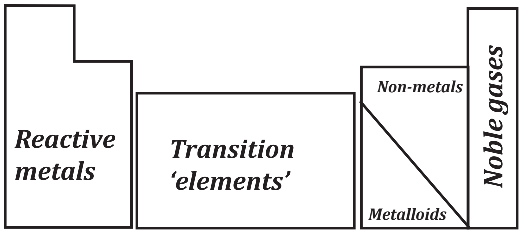


Figure 13.1 Blocks in the modern periodic table.

Table 13.2, position and electronic arrangement of the first twenty elements in the Periodic Table.

Groups	I	II	III	IV	V	VI	VII	O
Period 1	$^{1}_{1}$H							$^{2}_{2}$He
Period 2	^{3}Li 2.1	^{4}Be 2.2	^{5}B 2.3	^{6}C 2.4	^{7}N 2.5	^{8}O 2.6	^{9}F 2.7	^{10}Ne 2.8
Period 3	^{11}Na 2.8.1	^{12}Mg 2.8.2	^{13}Al 2.8.3	^{14}Si 2.8.4	^{15}P 2.8.5	^{16}S 2.8.6	^{17}Cl 2.8.7	^{18}Ar 2.8.8
Period 4	^{19}K 2.8.8.1	^{20}Ca 2.8.8.2						

The modern periodic table can be divided into blocks as it can be explained as follows:

(a) Reactive metals

These are metals of Group I and Group II. They include reactive metals such as potassium, sodium, calcium, and magnesium. They are the alkali metals and alkaline earth metals.

(b) Transition metals

Metals in this block resemble each other slightly across the period and down the group. They are less reactive than the alkali metals, but they have greater tensile strength.

(c) Metalloids

These are semi-metals which are fairly unreactive. They have some metallic and some non-metallic properties. Examples of these metals are: gallium, germanium and indium

(d) Non-metals

Non-metals form a group of elements which are also reactive. They have properties opposite to metals. They react by receiving electrons. With this behavior they have negative ions. Members of this group include nitrogen, chlorine, sulphur and bromine.

(e) Noble gases

Noble gases are also known as inert gases. They are gaseous non-metals which are very unreactive. They form the last group in the periodic table. These elements are stable and unreactive because they have complete outermost shell. Examples are Helium, Neon and Krypton.

13.3 GENERAL PERIODIC TRENDS

The general periodic trends can be explained as follows;

1. Melting point

Melting point is the temperature at which a solid melts to form a liquid. Down a group the melting points increases and decreases along a period from left to right of a periodic table.

2. Boiling Point

Boiling point is the temperature at which a liquid boils to form vapour. Generally metals have high boiling points than non-metals. The boiling point increases down a group and decreases along a period from left to right of the periodic table.

3. Density

Is the degree of compactness of a substance which means it is the mass per unit volume of a substance. Density increases down the group and decreases across the period from left to right.

4. Electronegativity

Electronegativity is the ability of an atom to attract an electron towards itself. Across the period from left to right electronegativity increases. Down the group electronegativity decreases.

5. Ionization Energy

Ionization energy is the energy required to remove an electron from an atom or ion. Across the period from left to right ionization energy increases and down the group ionization energy decreases.

6. Atomic Radius

Atomic radius is the distance between the nucleus of an atom and the outermost stable energy level. The atomic radius of elements across the period decreases from left to right. Atomic radius increases down the group as successive energy levels are filled.

7. Metallic Character

Element in group I, II, III and IV are metals. The metallic character of an element is its ability to lose electrons more easily from the valency shell. When an atom lose electrons from the valency shell it becomes positively charged. The metallic character decreases along a period from left to right. Down the group metallic character increase.

8. Non-Metallic Character

Non-metallic character is the tendency of an atom to accept an electron more easily into its valency shell. Across the period from left to right non-metallic character increases and down the group metallic character decreases.

SPECIFIC GROUP TRENDS.

Arranging the elements in the periodic table results in the elements to fall in groups of similar properties. Example, all elements in group two have the same number of electrons in their outermost shell. This explains why elements have similar properties. If elements of the same group are studied closely, some differences in properties are noted, some properties increase or decrease gradually toward one direction.

1. Group One Elements (Alkali Metals)

This group consists of five metals which are lithium (Li), sodium (Na), potassium (K), rubidium (Rb) and caesium (Cs). These elements have one electron each in their outer energy level.

Francium (Fr) is also an alkali metal but rarely included in the group. It is among the rarest naturally occurring elements.

Lithium, sodium and potassium all react very readily with water or air and are stored under paraffin oil. The three metals have the following properties:

Physical Properties;

i. They are good conductors of heat and electricity.
ii. They are soft metals
iii. They have low density.
iv. They have shiny surface when freshly cut.

Chemical Properties;

i. They burn in oxygen or air with a characteristic flame colour to form white solid oxides.
   Metal + oxygen→ metal oxides.
   The oxides dissolve in water to form alkaline solutions of the metal hydroxide.
   Metal oxide + water → metal hydroxide.
ii. They react vigorously with water to give the alkaline solution and hydrogen gas.
   Metal + water → metal hydroxide + hydrogen.

Table 13.3 trends in group I

Name(symbol)	Atomic number(Z)	Electronic configuration	Atomic radius (picometres)	1^{st} Ionization energy (kj/mol)	Melting point (°C)	Density (g/cm^3)	Electro negativity	Boiling point (°C)
Lithium (Li)	3	2:1	152	526	180	0.54	1.0	1347
Sodium (Na)	11	2:8:1	186	504	98	0.97	0.9	883
Potassium (K)	19	2:8:8:1	231	425	64	0.86	0.8	774
Rubidium (Rb)	37	2:8:18:8:1	244	410	39	1.5	0.8	688
Caesium (Cs)	55	2:8:18:18:8:1	262	380	29	1.9	0.7	678

2. Group Two Elements (Alkali Earth Metals).

Group II consist of six metals which are beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba) and radium (Ra).The elements have two electrons each in their outer energy level. These metals have the following properties.

Physical Properties

i. They are harder metals than those in group I.
ii. They are silvery grey in colour when pure and clean but they tarnish quickly when left in air due to the formation of the respective metal oxide.
iii. They are good conductors of heat and electricity.

Chemical Properties

i. They burn in oxygen or air with a characteristic flame colour to form a solid white oxide
   Metal + oxygen → metal oxide.
ii. They react with water but much less vigorously than the elements in group 1.
   Metal + water → metal hydroxide + hydrogen.
iii. The metals become more reactive as we move down the group.

Table 13.4 trends in group II

Name (symbol)	Atomic number (Z)	Electronic configuration	Atomic radius (picometer)	1^{st} ionization energy (KJ/mol)	Melting point (˚C)	Density (g/cm3)	Electro negativity	Boiling point (˚C)
Beryllium (Be)	4	2:2	112	899	1287	1.85	1.5	2469
Magnesium (Mg)	12	2:8:2	160	738	649	1.74	1.2	1090
Calcium (Ca)	20	2:8:8:2	197	590	839	1.55	1.0	1484
Strontium (Sr)	38	2:8:18:8:2	215	549	769	2.63	1.0	1382
Barium (Ba)	56	2:8:18:18:8:2	217	503	1850	3.51	0.9	1897

3. Group Seven Elements (Halogens)

The halogens in the periodic table consist of five chemically related elements, fluorine (F), chlorine (Cl), bromine (Br), iodine (I) and astatine (At)

Table 13.5 trends in group VII

Name (symbol)	Atomic number (Z)	Electronic configuration	Atomic radius (picometer)	1^{st} ionization energy (kj/mol)	Melting point(˚C)	Density (g/cm3)	Electro negativity.	Boiling point(˚C)
Fluorine (F)	9	2:7	71	1681.0	-219.62	0.017	3.98	-188.12
Chlorine (Cl)	17	2:8:7	99	1251.2	-101.5	0.0032	3.16	-34.04
Bromine (Br)	35	2:8:18:7	114	1139.9	-7.3	3.1028	2.96	58.8
Iodine (I)	53	2:8:18:18:7	133	1008.4	114	4.933	2.66	184.3

Physical Properties;

i. Halogens range from solid (I_2) to liquid (Br_2) to gaseous (F_2 and Cl_2) at room temperature.
ii. Fluorine is a poisonous pale yellow gas, chlorine is a poisonous pale green gas, bromine is toxic and caustic brown volatile liquid, and iodine is shiny black solid which easily sublimes to form a violet vapour on heating.
iii. At room temperature all the halogens exist as diatomic molecules.
iv. The melting points, boiling points, atomic radii and ionic radii all increase on descending the group.

Chemical Properties;

i. They are highly reactive especially with alkali metals and alkaline earth metals forming stable ionic crystals.
ii. They have very high electronegativities.
iii. The halogens oxides are acidic and the hydrides are covalent.
iv. The halogens form organic compounds which are best known for their industrial and environmental impact. Example PVC and DDT.
v. They oxidize metals to form halides.

SUMMARY

(a). In 1817, a German scientist, Dobereiner, noted that there was connection between relative atomic mass and chemical properties of an element.
(b). Dobereiner arranged elements into groups of three and called 'Dobereiner's triads'
(c). In Dobereiner's triads the relative atomic mass of the middle element was an average of the mass of the other two.
(d). It was noted that very few elements formed a group of three and therefore the Dobereiner classification was discouraged.
(e). In 1866, a British chemist, John Newlands, thought of the idea of arranging atomic masses.
(f). Newlands noticed that an element tended to display characteristics similar to the 8th element in front of it.
(g). Newlands arranged the elements in columns according to a law he called the ***Law of octaves***
(h). The Newland classification was unfortunate since he grouped together certain elements which had very different characteristics. Example oxygen was placed in the same group as iron and sulphur and therefore his ideas were rejected.
(i). A Russian chemist, Dimitri Mendelēev, later developed Newland's ideas and convinced other chemists to use them.

REVIEW QUESTIONS

1. The figure below is a skeleton of the periodic table. Only a few elements have been represented. The letters used are not the official chemical symbols of the elements. You may use these letters in attempting this question.
   (a) write the letters of three elements in the same period.
   (b) write the letters of the elements commonly known as;
      (i) Alkali metals (ii) Alkaline earth metals
      (iii) Halogens (iv) Noble gases
      (v) Transition metals.
   (c) write the letters of;
      (i) Seven metals and (ii) Six non-metals.

Z																	P
													U			T	
Y	G													Q			
	R						V				J		M				H
W																	

2. Choose the most correct answer for the following questions;
   i. The first periodic table was devised by the Russian chemist known as;............
      A Dobereiner B Newland C Medvedev D Mendelēev.
   ii. Newland introduced the...................... which states that "elements with similar chemical properties have atomic numbers which differ by 7".
      A. Law of octaves B. law of elements
      C. law of octate D. Periodicity.
   iii. ______is the regular periodic changes of elements due to their atomic number.
      A. Periodic law B. Periodic table
      C. Periodic classification D. Periodicity
   iv. The vertical columns of elements in the periodic table are called;___________
      A. Groups B. Periods C. Metals D. Non-metals
   v. The metallic elements in Group I of the periodic table are called;________
      A. Alkali elements B. Alkali earth metals
      C. Alkali compounds D. Alkali metals.

3. Match the items in LIST A with their correct statements in LIST B;

List A	List B
i. The modern periodic law	A. The temperature at which a liquid boils to form a gas.
ii. Group II metals	B. The temperature at which a solid melts to form a liquid.
iii. Elements which display metallic and non-metallic characteristics	C. The properties of elements are a periodic function of their atomic numbers.
iv. Melting point	D. Alkaline metals
v. Boiling point	E. Alkaline earth metals
	F. Metalloids

4. Write ***TRUE*** for a correct statement and ***FALSE*** for an incorrect statement.
   i. The atomic radii of elements in a periodic table decrease from left to right...........
   ii. Elements to the left of the periodic table show non-metallic properties...........
   iii. Electronegativity decreases from left to right across the period.....................
   iv. Densities of elements increases down the group...
   v. Electronegativity and ionization energy decrease down the group................

5. Explain how electronegativity, atomic radius, metallic properties and non-metallic properties vary;
   (a) Across a period from left to right of the periodic table.
   (b) down a group
6. What do you understand by;
   (a) modern periodic table
   (b) periodicity
   (c) transition elements
   (d) modern periodic law
   (e) law of octaves
7. A metallic element Y reacts slowly with water to give a strongly alkaline solution. In which group of the periodic table would you place Y?
8. With two examples briefly explain the meaning of metalloids.
9. (a) Define noble gases.
   (b) Name the group which they belong in a periodic table.
   (c) Explain briefly how noble gases;
      (i) Are arranged in electronic configuration
      (ii) Undergo chemical reaction.

Chapter 14

Formula, Bonding and Nomenclature

14.1 CHEMICAL FORMULAE

A ***formula*** is a method of representing a molecule of a compound. Example the chemical formula of limestone is $CaCO_3$. The formula of a compound expresses the composition of a substance in terms of the relative numbers of atoms that have chemically combined together. In a formula each atom present is represented by its symbol. The symbol of an element represents one atom of the element. A symbol is shown by one or two letters, the first of which is always a capital.

One atom of chlorine is represented by the symbol Cl. When two atoms of chlorine combine, the molecule formed is represented by the formula Cl_2. Water is denoted by the formula H_2O which contains two atoms of hydrogen and one atom of oxygen combined. Two molecules of water are written $2H_2O$; the two infront of the formula multiplies everything that follows it.

Qualitatively H_2O means a water molecule is made of hydrogen and oxygen atoms. The quantitative meaning of water gives us the amounts:

- One molecule of water
- One mole of water molecule
- 18g of water
- Two hydrogen atoms for every one oxygen atom.

Consider the table below showing chemical formulae of some compounds.

Table 14.1 Chemical formulae of some compounds.

Compound	Chemical formula
Calcium hydroxide	$Ca(OH)_2$
Zinc chloride	$ZnCl_2$
Sodium carbonate	Na_2CO_3
Copper(II)sulphate	$CuSO_4$
Sulphuric acid	H_2SO_4
Hydrogen peroxide	H_2O_2
Potassium chlorate	$KClO_3$
Nitric acid	HNO_3

14.2 VALENCY

Valency is the combining capacity of an element. Atoms lose, gain or share electrons which are equal to their valencies in forming compounds.

Consider the following examples.

(1). All elements which have one electron in their outermost shell have a valency of 1. E.g. sodium (2:8:1) and potassium (2:8:8:1). Elements with two, three and four electrons in their outer most shells have valencies of 2, 3 and 4 respectively e.g. calcium (2:8:8:2) has two electrons in its outermost shell, therefore its Valency is 2, Aluminium (2:8:3) has three electrons in its outer most shell and its valency is 3 and carbon (2:4) has four electrons in its outer most shell and its valency is 4.

(2). Elements with electrons greater than 4 in the outer most shells, their valency can be calculated by using the formula; ***V= 8 -n.e***

Where by;

V = Valency

8 = represents 8 electrons in the fulfilled outermost shell.

n.e = number of electrons in the outermost shell.

(a). The electronic configuration of oxygen is 2:6 the number of electrons in the outermost shell is 6.

$V = 8\text{-n.e}$ $\quad V = 8-6$ $\quad V = 2$

Therefore, Valency of oxygen is 2.

(b). Electronic configuration of fluorine is 2:7. The number of electrons in the outermost shell is 7.

$V = 8\text{-n.e}$ $\quad V = 8-7 = 1$

Therefore, valency of fluorine is 1.

14.3 OXIDATION STATES

Oxidation states (also oxidation number) tell us the number of electrons a particular element has lost, gained or shared in forming a compound. When an element is alone, its oxidation state is zero.

There is a close relationship between valency and oxidation state, but they are not the same thing. The oxidation state of an atom in a compound is assigned relative to the other elements in that compound. While oxidation states are an arbitrary assignment, valency is a fixed value. Oxidation state may be assigned a positive or a negative number.

VARIABLE VALENCY

Some elements have more than one valency. Such elements can change their oxidation states. Examples are iron and copper.

Compounds in which iron has oxidation number $^+3$ e.g. $FeCl_3$ and those which the oxidation number is $^+2$ e.g. $FeCl_2$ are well known. Copper has oxidation number $^+2$ in many of its compounds e.g. $CuSO_4$ and only a few compounds with the oxidation number $^+1$ e.g. Cu_2O. If an element has a variable valency, its valency must be shown by a roman numeral in the names of all its compounds.

Consider the table below which show a summary on valencies;

Table 14.2 Summary of valencies

	Valency 1	Valency 2	Valency 3
Metals	Potassium,K, Sodium,Na silver,Ag	Calcium, Ca Magnesium, Mg	Aluminium ,Al Iron (III) Ferric, Fe.
Non-metals	Hydrogen,H, Chlorine ,Cl	Oxygen, O Sulphur, S	Nitrogen,N phosphorus ,P
Radicals	Ammonium NH_4^+ Hydroxide OH^- Nitrate NO_3^- Nitrite NO_2^- Hydrogen carbonate HCO_3^-	Carbonate CO_3^{2-} Sulphate SO_4^{2-} Sulphite, SO_3^{2-}	Phosphate(V)oxide, PO_4^{3-}

Example 14.1;

Give the oxidation number of Cr in $Cr_2O_7^{2-}$

Solution

Total charge on the dichromate ion = -2

For oxygen $^-2 \times 7 = -14$

Then, $2Cr - 14 = -2$

$2Cr = {}^+12$

$Cr = {}^+6$

Therefore, the oxidation number of Cr is $^+6$.

Example 14.2

Find the oxidation state of chlorine in the compound $KClO_3$.

Solution

The oxidation number of potassium is $^{+}1$.

The oxidation number of oxygen is $^{-}2$. With three oxygen atoms the oxidation number is given by; (-2x3) = -6

Therefore, +1 + Cl-6 = 0

Cl = 6-1 = +5

Therefore, Oxidation number of chlorine in $KClO_3$ is $^{+}5$

Example 14.3

Calculate the oxidation number of sulphur in H_2SO_4.

Solution

Total oxidation state of $H_2SO_4 = 0$

Oxidation state for H= +1; O = -2; S =?

H_2SO_4 is a neutral compound and therefore the oxidation number of the molecule is zero.

Therefore;

$2x\ (+1) + S + 4x\ (-2) = 0$

$+2 + S - 8 = 0$

$S - 6 = 0$

$S = +6$

Therefore, the oxidation number of sulphur in H_2SO_4 is +6.

Example 14.4

Calculate the oxidation state of nitrogen in NH_4^+

Solution

Total oxidation state of $NH_4^+ = +1$

Oxidation state for H = +1, N=?

Therefore; $N + (4X\ (+1)) = +1$

$N + 4 = +1$

$N + 4 - 4 = +1 - 4$

$N = -3$

Therefore, the oxidation state of nitrogen in NH_4^+ is -3

Example 14.5

What is the oxidation number of sulphur in the sulphate ion, SO_4^{2-} .

Solution;

The total charge on the sulphate ion is -2.

The oxidation number of oxygen is -2.

Then; $S + (-2x4) = -2$

$S-8 = -2$

$S = 8-2$

$S = +6.$

Oxidation state of sulphur in SO_4^{2-} is +6.

Example 14.6

Calculate the oxidation state of sulphur in HSO_3^-

Solution

Total oxidation state of HSO_3^- = -1

Oxidation state for H=+1, O=-2, S=?

$HSO_3^- = -1$

$+1 + S + 3x(-2) = -1$

$+1 + S - 6 = -1$

$S = +4$

Therefore, the oxidation state of sulphur in HSO_3^- is +4

14.4 RADICALS

A radical is a group of elements which acts like a single atom in forming compounds. Radicals do not have independent existence.

Most radicals are negatively charged. However, a few like ammonium are positively charged.

Consider the table below which shows the radicals and their names, symbol, oxidation state, valency and their compounds.

Table 14.3; the common radicals.

Name of radical	Symbol	Oxidation state	Valency	Example of compound in which it is found.
Ammonium	NH_4^+	+1	1	Ammonium chloride NH_4Cl
Sulphate	SO_4^{2-}	-2	2	Calcium sulphate $CaSO_4$
Carbonate	CO_3^{2-}	-2	2	Zinc carbonate $ZnCO_3$
Hydrogen carbonate	HCO_3^-	-1	1	Sodium hydrogen carbonate $NaHCO_3$

Name of radical	Symbol	Oxidation state	Valency	Example of compound in which it is found.
Hydrogesulphate	HSO_4^-	-1	1	Potassium hydrogen sulphate $KHSO_4$
Nitrate	NO_3^-	-1	1	Silver nitrate $AgNO_3$
Hydroxide	OH^-	-1	1	Potassium hydroxide KOH
Phosphate(V)	PO_4^{3-}	-3	3	Phosphoric (V) acid H_3PO_4
Chlorate(V)	ClO_3^-	-1	1	Sodium chlorate(V) $NaClO_3$
Chlorate(I)	ClO^-	-1	1	Potassium chlorate(I) KClO
Sulphite(IV)	SO_3^{2-}	-2	2	Sulphuric (IV) acid H_2SO_3

14.5 IONS

Metal atoms lose electrons to become ions. The atom will lose as many electrons as its valency. Metal ions are always positively charged.

The general equation for the loss of electrons by metal atoms is:$M \rightarrow M^{n+} + n\bar{e}$

Where;

M = Metal atom

M^{n+} = Metal ion which carries a charge of n+

n = number of electrons lost

$\bar{e}$ = the negatively charged electron.

Non – metal atoms gain electons to become ions. The atom will gain as many electrons as its valency. Ions of non – metals are always negatively charged.

The general equation for the gain of electrons by a non – metal atom is;

$X + n\,\bar{e} \longrightarrow X^{n-}$

Where by;

X = non - metal atom

X^{n-} = non - metal ion which has negative charge.

Consider the below table showing ions of common metals and non – metals.

Table 14.4 Ions of common metals and non metals.

Ionic name	Symbol	Charge	Valency	Example of a compound
Hydrogen	H^+	+1	1	Hydrogen chloride, HCl
Potassium	K^+	+1	1	Potassium bromide, KBr
Sodium	Na^+	+1	1	Sodium chloride, $NaCl$
Copper (I)	Cu^+	+1	1	Copper (I) oxide, Cu_2O
Nickel	Ni^{2+}	+2	2	Nickel sulphate, $NiSO_4$
Silver	Ag^+	+1	1	Silver chloride, $AgCl$
Copper (II)	Cu^{2+}	+2	2	Copper (II) sulphate, $CuSO_4$
Barium	Ba^{2+}	+2	2	Barium sulphate, $BaSO_4$
Calcium	Ca^{2+}	+2	2	Calcium nitrate, $Ca(NO_3)_2$
Magnesium	Mg^{2+}	+2	2	Magnesium oxide, MgO
Zinc	Zn^{2+}	+2	2	Zinc sulphate, $ZnSO_4$
Iron	Fe^{2+}	+2	2	Iron (II) chloride, $FeCl_2$
Lead	Pb^{2+}	+2	2	Lead (II) sulphate $PbSO_4$
Mercury (II)	Hg^{2+}	+2	2	Mercury (II) chloride, $HgCl_2$
Aluminium	Al^{3+}	+3	3	Aluminium oxide, Al_2O_3
Ammonium	NH_4^+	+1	1	Ammonium nitrate, NH_4NO_3
Chloride	Cl^-	-1	1	Aluminium chloride $AlCl_3$
Bromide	Br^-	-1	1	Potassium bromide KBr
Iodide	I^-	-1	1	Silver iodide AgI
Oxide	O^{2-}	-2	2	Magnesium oxide MgO
Sulphide	S^{2-}	-2	2	Lead (II) sulphide, PbS
Hydroxide	OH^-	-1	1	Ammonium hydroxide, NH_4OH
Nitrate	NO_3^-	-1	1	Silver nitrate, $AgNO_3$
Carbonate	CO_3^{2-}	-2	2	Calcium carbonate, $CaCO_3$
Sulphate	SO_4^{2-}	-2	2	Iron (II) sulphate, $FeSO_4$
Hydrogen carbonate	HCO_4^-	-1	1	Sodium hydrogen carbonate $NaHCO_3$
Hydrogen sulphate	HSO_4^-	-1	1	Potassium hydrogen sulphate $KHSO_4$
Sulphite	SO_3^{2-}	-2	2	Sodium sulphite Na_2SO_3

14.6 WRITING A CHEMICAL FORMULAE

The points to remember when writing chemical formulae are explained as follows.

1. Positively charged ions are written before the negatively charged ions.
2. A radical must be treated as a unit.
3. The name of the cation is the same as the neutral element from which it is derived (e.g. Na for sodium)
4. Single elements are not bracketed.
5. The valency of 1 is simply assumed and not written in the formula.

Consider the below examples.

Examples 14.7

Write the formula of Copper (II) chloride.

Step 1. Write down the name: Copper (II) chloride.

Step 2. Write down the ions in the compound, complete with their charges.

$Cu^{2+}Cl^{-}$

Step 3. Draw boxes round the ions and draw arrows to exchange the valencies.

$\boxed{Cu^{2+}}$ $\boxed{Cl^{-}}$

$\boxed{Cu}$ $\boxed{2Cl}$

Step 4. Write the exchanged valencies as subscripts. A valency of 1 should not be written. In this step, radicals must be bracketed if necessary. The complete formula is $CuCl_2$. The compound contains two chloride ions for every one Copper (II) ion.

Example 14.8

Write the formula of aluminium Sulphate.

Step 1. Aluminium sulphate

Step 2. $\boxed{Al^{3+}}$ $\boxed{SO_4^{2-}}$

Step 3. $\boxed{2Al^{3+}}$ $\boxed{3SO_4^{2-}}$

Step 4. $Al_2(SO_4)_3$

The complete formula is $Al_2\ (SO_4)_3$

Example 14.9

Write formula for calcium oxide.

Step 1: calcium oxide

Step 2: $\boxed{Ca^{2+}}$ $\boxed{O^{2-}}$

Step 3: $\boxed{2Ca^{2+}}$ $\boxed{2O^{2-}}$

Step 4: CaO

The complete formula is CaO.

Example 14.10

Write the formula for zinc chloride

Step 1: Zinc chloride

Step 2: $\boxed{Zn^{2+}}$ $\boxed{Cl^{-}}$

$\boxed{1Zn}$ $\boxed{2Cl^{-}}$

Step 4: $ZnCl_2$

The complete formula is $ZnCl_2$.

Activity 14.1

Write the chemical formula of the compounds formed when the following combine.

(a). Sodium ion and carbonate ion.
(b). Sodium ion and cyanide ion.
(c). Ammonium ion and sulphate ion.
(d). Hydrogen ion and chloride ion.
(e). Copper ion and nitrate ion.
(f). Magnesium ion and chloride ion.
(g). Write the names of each of the following compounds.
  (i). $CuCl_2$
  (ii). Na_2SO_4
  (iii). NH_4NO_3.

14.7 EMPIRICAL AND MOLECULAR FORMULAE

Chemical formula can basically be divided into three types, namely empirical formula, molecular formula and structural formula. For this level we shall concentrate on the empirical and molecular formulae.

Empirical formula

The empirical formula of a compound is the simplest formula which expresses its composition by mass.

Molecular formula.

The molecular formula is the formula which shows the actual number of each kind of atom present in a molecule. The molecular formula is equal to n(empirical formula), where n is a whole number.

CALCULATING THE EMPIRICAL AND MOLECULAR FORMULAE.

Example 14.11

Sodium sulphate has the following composition by mass; Sodium 32.4%, Sulphur 22.5% and Oxygen 45.1%. What is its empirical formula?

Solution

Elements	Na	S	O
% Composition	32.4	22.5	45.1
Divide by relative Atomic mass	$\frac{32.4}{23} = 1.14$	$\frac{22.5}{32} = 0.70$	$\frac{45.1}{16} = 2.82$
Divide by smallest number	$\frac{1.14}{0.70} = 2$	$\frac{0.70}{0.70} = 1$	$\frac{2.82}{0.70} = 4$
	2	1	4

Therefore the empirical formula is Na_2SO_4

Example 14.12

What is the empirical formula for a compound of mass 8.1g if it consists of 4.9 g of magnesium and 3.2g of oxygen?

Solution

Elements	Mg	O
Relative Masses	4.9	3.2
Divide by relative Atomic mass	$\frac{4.9}{24} = 0.20$	$\frac{3.2}{16} = 0.20$
Divide by smallest number	$\frac{0.2}{0.2} = 1$	$\frac{0.2}{0.2} = 1$
	1	1

Therefore the empirical formula is MgO

Example 14.13

An experiment showed that 13.88g of calcium chloride were obtained from the combination of 5g of calcium with unknown relative mass of chlorine. What is the formula of calcium chloride?

Solution.

Mass of calcium chloride = 13.88g

Mass of calcium = 5.00g

Mass of chlorine =?

Mass of chlorine = mass of calcium chloride – mass of calcium.

Mass of chlorine = (13.88 – 5.00)g

Hence the mass of chlorine = 8.88g.

Elements	Ca	Cl
Relative masses	5.00	8.88
Divide by relative Atomic mass	$\frac{5.00}{40} = 0.125$	$\frac{8.88}{35.5} = 0.25$
Divide by smallest number	$\frac{0.125}{0.125} = 1$	$\frac{0.25}{0.125} = 2$
	1	2

Therefore the empirical formula is $CaCl_2$

Example 14.14.

A compound contains 15.8% carbon and 84.2% sulphur. Calculate its empirical formula. If its relative molecular mass is 76, what is its molecular formula?

Solution.

Elements	Carbon (C)	Sulphur (S)
% Composition	15.8	84.2
Divide by relative Atomic mass	$\frac{15.8}{12} = 1.32$	$\frac{84.2}{32} = 2.63$
Divide by smallest number	$\frac{1.32}{1.32} = 1$	$\frac{2.63}{1.32} = 2$
	1	2

Therefore the empirical formular is CS_2.

The molecular formular is given by;

Molecular formulae = n (empirical formula)

Where n is a whole number.

Then; Molecular formula = n (CS_2)

Now Relative molecular Mass = n (sum of relative atomic mass)

76 = n (12+ (2x32))

76 = n (12 + 64)

76 = 76n

n = 1

Hence the molecular formula is CS_2.

Example 14.15

A certain gaseous compound contains 82.8% of carbon and 17.2% of hydrogen by mass. The vapour density of the compound is 29. Calculate its molecular formula.

Elements	C	H
% Composition	82.8	17.2
Divide by relative Atomic mass	$\frac{82.8}{12} = 6.90$	$\frac{17.2}{1} = 17.2$
Divide by smallest number	$\frac{6.90}{6.90} = 1$	$\frac{17.2}{6.90} = 2.5$
Multiply by 2 to get whole number only	1 x 2 = 2	2.5 x 2 = 5

Empirical formula is C_2H_5

Relative molecular mass of gas = 2 x Vapour density.

= 2 x 29 = 58

Molecular formula is given by:-

(Empirical formula) n = relative molecular mass

(C_2H_5) n = 58

$$[(12x2)+ (1x5)]\ n = 58$$
$$(24 + 5)\ n = 58$$
$$\frac{29n}{29} = \frac{58}{29}$$
$$n = 2$$
$$(C_2H_5)_2 = C_4H_{10}$$

The molecular formula of the gas is C_4H_{10}

Example 14.16

A certain compound contains 22.7% of zinc, 11.0% of sulphur, 22.3% of oxygen and the rest being water of crystallization. Calculate

(i) Empirical formula.

(ii) Molecular formula of a compound if its molar mass is 287g.

Solution;

Elements	Zn	S	O	H_2O
% Composition	22.7	11.0	22.3	44
Divide by relative Atomic mass	$\frac{22.7}{63} = 0.36$	$\frac{11.0}{32} = 0.34$	$\frac{22.3}{16} = 1.39$	$\frac{44}{18} = 2.44$
Divide by smallest number	$\frac{0.36}{0.34} = 1$	$\frac{0.34}{0.34} = 1$	$\frac{1.39}{0.34} =4$	$\frac{2.44}{0.34} = 7$
	1	1	4	7

The empirical formula is $ZnSO_4{\bullet}7H_2O$

(ii) Given molar mass 287g

$$(\text{Empirical formula})\, n = \text{molar mass}$$

$$[\,(65 + 32 + (16x4) + (7x18)\,]n = 287$$

$$287n = 287$$

$$n = 1$$

The molecular formula is $ZnSO_4{\bullet}7H_2O$

14.8 BONDING

Bonding is the process of joining two or more things or substances together. A link that holds two or more things together is called a bond.

Chemical substances are made of atoms that are held together by chemical bonds.A chemical bond is a force of attraction that holds atoms together to form molecules.

CHEMICAL BONDING

Chemical bonding is a process by which atoms become more stable by donating, gaining or sharing electrons. Chemical bonding involves forces of attraction between electrons in the outermost energy levels of atoms.

When the outermost levels are filled, the atoms are said to be stable. The stable atoms are generally unreactive and can exist freely as single atoms. Unstable atoms cannot exist freely as single toms. For unstable atoms to become stable, they should acquire electronic arrangements similar to those of noble gas. This means that they either lose, gain or share electrons.

Types of chemical bonding;

- There are two main types of chemical bonds, these are ionic and covalent bonds. Ionic bond involves the transfer of electron from a metal atom to the non-metal atom. Covalent bond involves the sharing of a pair of electrons in which, each bonding atom donates sharing electrons.

ELECTROVALENT BONDING (IONIC BONDING)

An ***electrovalent bond*** is the bond due to the transfer (gain or loss) of one or more electrons from one atom or radical to another. The electrovalence of an atom or a group of atoms is the number of electrons which it transfers or receives. Electrovalent bonding usually occurs between a metal and a non-metal.The metal loses electrons and the non-metal gains electrons.

Example; when sodium and chlorine react to form sodium chloride, the sodium atoms must each lose an electron to acquire a stable neon structure, while chlorine atoms must each gain an electron to acquire a stable argon structure, this results in positively charged sodium ions and negatively charged chloride ions to form sodium chloride molecules.

Consider the following ***"cross"*** and ***"dot"*** diagrams in which a cross (x) represents the electrons in a non- metal atom, the dot (•) represents the electron transferred from a metal atom.

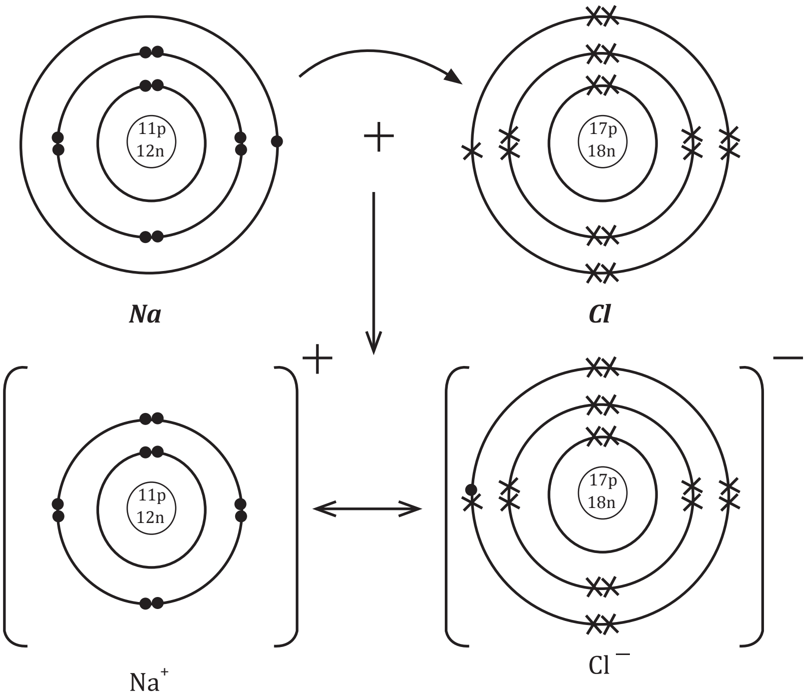


Figure 14.1 electrovalent bond in sodium chloride.

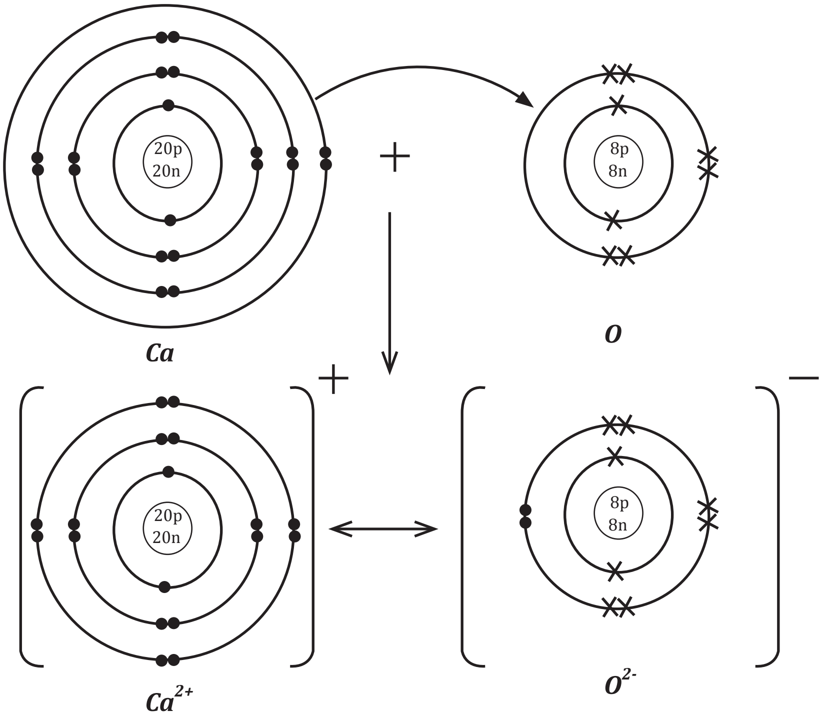


Figure 14.2; electrovalent bonding in calcium oxide.

Properties of electrovalent compounds

(i). They are generally soluble in water

(ii). Conduct electricity when in solution or molten form, but not when in solid form

(iii). They are usually crystalline solids at room temperature.

(iv). They have high melting and boiling points.

(v). They are generally insoluble in non-polar solvents.

(vi). They consist of ions that are positively and negatively charged.

Task 14.1

1. Cations are slightly smaller than their neutral atoms, where as anions are slightly larger than their neutral atoms. Explain why.
2. Briefly explain how magnesium and oxygen atoms combine to form magnesium oxide molecule. In your answer highlight
   (a). the charge on a magnesium ion (b) The charge on an oxide ion
3. Illustrate using a dot and cross diagram the electron transfer involved in forming potassium chloride bond.

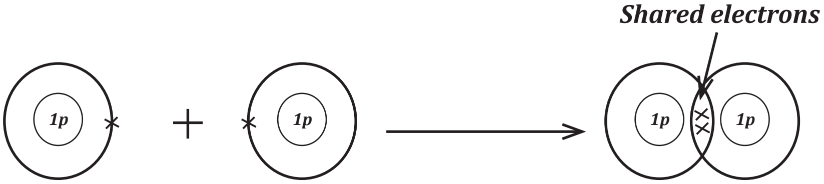
COVALENT BONDING

A ***covalent bonding*** is a type of bonding which involves the sharing of electrons between atoms. In covalent bonding, the atoms share electrons and the atoms within each molecule are held together by the attraction between their positive nuclei and these shared electrons.

A molecule is thus formed in each atom which has attained an energetically stable arrangement of electrons. Covalent bond occurs when non-metal atoms combine with other non-metal atoms. Molecules of covalent compounds are discrete or distinct. It means that the atoms forming the molecules cannot exist freely in the compound. They remain bound together in molecules and their electrons are not free to form bonds with other atoms.

The following examples of covalent bond formation are concerned.

1. Hydrogen atom needs one electron to acquire the helium structure. It can combine with another hydrogen atom and share the electrons so that each acquires a stable helium structure. This results in the formation of a hydrogen molecule. The kind of bond formed between the two atoms is called a covalent bond.



or $H^{\bullet} + H^{\bullet} \rightarrow H : H$ or $H - H$

Figure 14.3 covalent bonding in hydrogen molecule.

2. Fluorine molecule is formed by means of establishing a covalent bond between two fluorine atoms. Each of the fluorine atoms donates single electrons from its outer most shell.

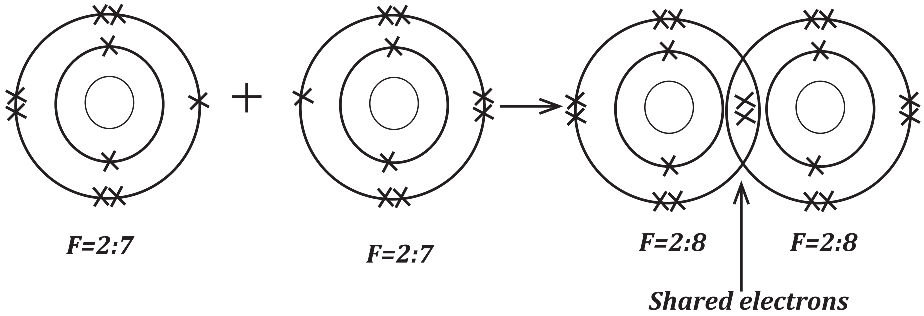
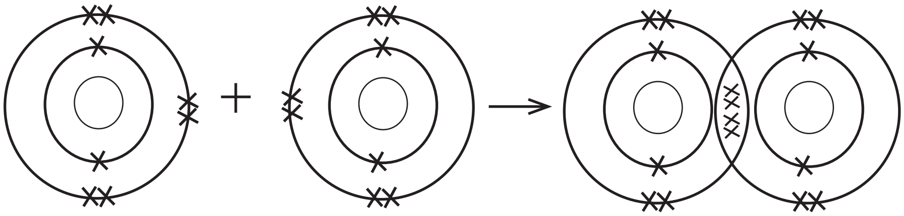


Figure 14.4 covalent bonding in fluorine molecule.

In fluorine molecule, each atom has 6 electrons of its own in the outer most orbits plus the shared pair, which revolves around both atoms. i.e. They have a complete outer octet and the stable neon structure (2:8).

3. Oxygen requires two electrons to acquire the stable atom structure. Therefore two atoms of oxygen combine and share four electrons between them.



two atoms of oxygen ***Oxygen molecule***

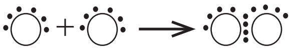
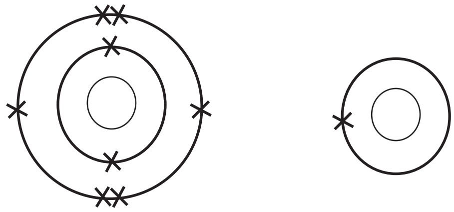
Or  or O = O

Figure 14.5 covalent bonding in oxygen. (double covalent bond).

4. In the water molecule (H_2O) each of the unpaired electrons in the outer most shell of oxygen atom is used for covalent bonding with the electrons of the two hydrogen atoms. Therefore, there are two single covalent bonds in the water molecule.



O=2:6 ***H=1***

Then, two hydrogen atoms combine with one atom of oxygen to form water molecule.

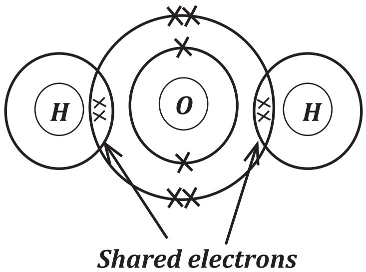


Fig 14.6 covalent bonding in water.

Both oxygen and hydrogen attain a noble gas configuration. O^{2-} = 2:8 and H^{-} = 2

5. Consider a triple covalency such as nitrogen gas (N_2). It is a diatomic. A triple bond is formed between the two atoms. Electronic configuration for nitrogen is 2:5

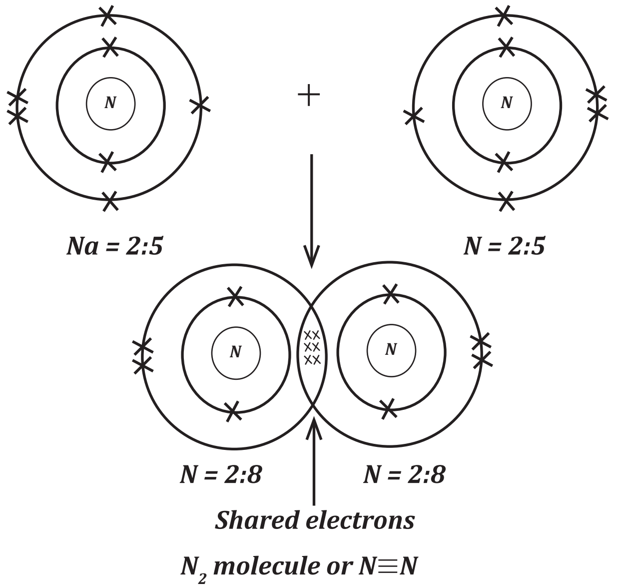


Fig. 14.7 Covalent bonding in nitrogen

Properties of covalent compounds

(i). The melting and boiling points are usually low

(ii). They are usually liquids or gases at room temperature.

(iii). They are generally insoluble in water but soluble in organic solvents such as ether, ethanol or benzene.

(iv). They do not conduct electricity when solid, liquid or in solution because they do not have ions.

(v). They consist of molecules with two or more different atoms linked together by covalent bonds.

Experiment 14.1

Aim; to find out the electrical conductivity of covalent and electrovalent compounds.

Procedures;

1. Use crucible or beaker to hold the solution which is to be tested.
2. Arrange the electric circuit as in figure 14.8 . A 6-volt battery or any other 6-volt direct current (d.c) supply is connected by a copper wire to a switch and either a 6-volt bulb in a bulb holder or an ammeter. You can include a variable resistor if necessary. Two carbon rods take the current into and out of the solution being tested.

3. Fill the container with one solution under test. Put on the switch to complete the circuit. If the bulb glows or the ammeter shows a reading, the solution is a conductor of electric current. Switch off at once.
4. Use the following solutions; diesel, ethanol, dilute hydrochloric acid, sodium hydroxide, sodium chloride, sugar solution and dilute sulphuric acid and clean the two carbon rods after each experiment. If we are to investigate the conduction of electricity by molten substances we should use the same circuit as that used for investigation of the conduction of electricity by a variety of substances in the solid and liquid states. The substances to be used for testing are such as wax, sugar, lead (II) bromide, potassium iodide, lead (II) iodide.

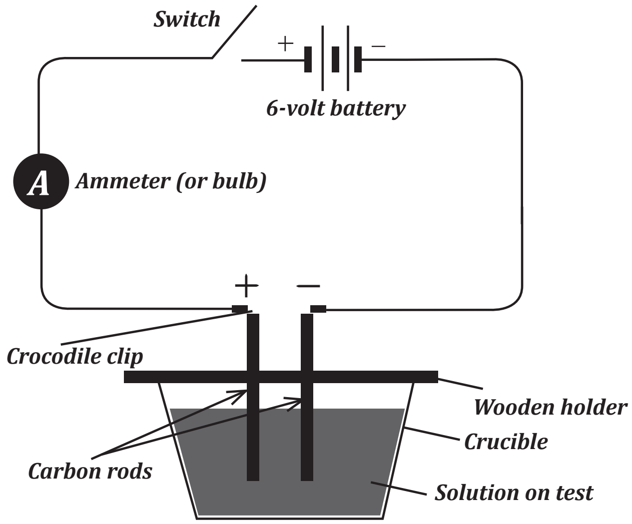


Figure 14.8; to find out which solids and solutions conduct electricity.

From your results of the experiment, divide the solutions into two groups:

(a). The solutions which conduct electric current and cause light in the bulb or deflection of the ammeter.

(b). Solutions which neither conduct electric current nor cause the bulb to glow or no any ammeter deflection. Which of these solutions are covalent compounds and which are electrovalent compounds?

Table 14.5. Differences between covalent and electrovalent compounds.

Covalent compounds	Electrovalent compounds.
They are formed by mutual sharing of electrons	They are formed by complete transfer of electrons
They are made up of molecules	They are made up of ions
They are usually liquids or gases	They are usually crystalline solids
They are soluble in non-polar solvent like benzene or carbon tetrachloride and insoluble in polar solvents like water	They are usually soluble in water but insoluble in non-polar solvents like carbon tetrachloride
They generally have low melting and boiling points	They have generally high melting and boiling points
They do not conduct electricity.	They are good conductors of electricity in the molten state and in aqueous solutions but insulators in the solid state.

14.9 NOMENCLATURE OF BINARY INORGANIC COMPOUNDS

Nomenclature is a systematic way of assigning names to items or substances that belong to a particular group or classification.

A binary compound is a chemical compound that contains exactly two different elements. ***Inorganic compounds*** fall mainly into two main categories, namely ionic and covalent.

BINARY IONIC COMPOUNDS

A binary ionic compound is a salt consisting of only two elements in which both elements are ions, a cation (which has a positive charge) and anion (which has a negative charge).

The following are steps to be considered when naming binary ionic compounds:

(1). Name the metallic ion that appears first in the formula using the name of the element itself.

(2). The second part of the formula, which is usually an anion in the compound, will end in “ide”. Example, chlorine become chloride, fluorine become fluoride and hydroxyl become hydroxide.

Example 14.17

What is the name of the compound of formula CuS?

Solution;

Step 1.

Let x be the valence of a Cu atom.

Step 2. Sulphur has a -2 charge

Step 3. $1(x) + 1(-2) = 0$

$x = +2$ for Cu

Step 4. Write the name copper and place (II) in brackets beside it.

Step 5. Use the name sulphur but change the last two letters to "ide".

Therefore the compound is copper (II) sulphide.

Example 14.18

What is the name of the compound with formula $FeCl_3$?

Solution;

The total charge of the molecule is zero and Cl^- has a negative charge.

Then;

Step 1. let x be the valence of the Fe atom

Step 2. $1(x) + 3(-1) = 0$

Step 3. $x = +3$

Step 4. So the Fe is in the +3 oxidation state.

Write the name "iron" and place (III) in brackets beside it.

Step 5. Use the name chlorine but change the last three digits to "ide".

Hence the name is iron (III) chloride.

More examples of binary ionic compounds are;

CaO is named Calcium oxide

AlN is named aluminium nitride

K_2S is named potassium sulphide

MnO_2 is named manganese (IV) oxide.

Note; Manganese can have more than one charge, but each oxide ion has a charge of -2. In order for the compound to be neutral, Mn must have a +4 charge.

BINARY COVALENT COMPOUNDS.

A binary covalent compound is composed of two different non-metal elements. Example, a molecule chlorine trifluoride (ClF_3) contains 1 atom of chlorine and three atoms of fluorine. A set of prefixes is used to identify the subscript in the binary covalent compound formula.

Consider the table below: Table 14.6. List of prefixes

Number	prefix
1	mono
2	di
3	tri
4	tetra
5	penta
6	hexa
7	hepta
8	Octa-
9	Nona-
10	Deca-

The steps to be considered when naming binary covalent compounds are;

1. Give the name of the first element
2. Give the name of the second element with the ending changed to –ide.
3. If more than one compound is possible between the two elements, give prefixes to indicate the number of atoms of each element.

Example 14.19

Name the compound N_2O_4.

(i). Use the prefix "ide" in front of nitrogen
(ii). Use the prefix "tetra" in front of the oxygen.
(iii). We drop "-ygen" and replace with "ide".

Therefore the name is dinitrogen tetraoxide

Example 14.20

What is the name for PCl_3?

(i). Since there is one phosphorus atom, we use it as the first part of the name.
(ii). There are three chlorine atoms, so we use "tri" in front of chlorine. We then drop the "-ine" in chlorine and replace with "ide".

Therefore the name is phosphorus trichloride.

Consider the table below which gives the formulae and names of some binary covalent compounds.

Table 14.7 some binary covalent compounds

Formula	Name
CO_2	Carbon dioxide
CO	Carbon monoxide
P_2O_5	Diphosphorus pentoxide
N_2O	Dinitrogen monoxide
SiO_2	Silicon dioxide
CBr_4	Carbon tetrabromide
SO_3	Sulphur trioxide
PBr_5	Phosphorus pentabromide
ICl_3	Iodine trichloride
NI_3	Nitrogen triiodide
N_2O_3	Dinitrogen trioxide
N_2O_4	Dinitrogen tetraoxide
SO_2	Sulphur dioxide
NO	Nitrogen monoxide
NO_2	Nitrogen dioxide
As_2O_5	Diarsenic pentoxide
PCl_3	Phosphorus trichloride
CCl_4	Carbon tetrachloride
H_2O	Dihydrogen monoxide (most call it water)
SF_6	Sulphur hexafluoride
XeO_3	Xenon trioxide
P_4S_3	Tetraphosphorus trisulphide

CHEMICAL NAMES OF COMMON SUBSTANCES

Chemical names are usually used to give an accurate description of the composition of a substance.

The table below gives a list of common names for some chemicals.

Table 14.8 chemical names of common substances

Common name	Chemical name
Common salt/table salt	Sodium chloride ($NaCl$)
Washing soda/soda ash	Sodium carbonate ($Na_2 CO_3$)
Limestone/marble/chalk	Calcium carbonate ($CaCO_3$)
Quick lime /lime	Calcium oxide (CaO)
Lime water	Calcium hydroxide $Ca(OH)_2(aq)$ solution
Slaked lime	Calcium hydroxide ($Ca(OH_2)$)
Salt petre	Potassium nitrate (KNO_3)
Plaster of Paris (POP)	Calcium sulphate -1-water ($CaSO_4 \bullet H_2O$)
Gypsum	Calcium sulphate-2-water ($CaSO_4 \bullet 2H_2O$)
Caustic soda	Sodium hydroxide ($NaOH$)
Caustic potash	Potassium hydroxide (KOH)
Brine	Concentrated sodium chloride solution (($NaCl(aq)$))
Iodine tincture	A mixture of iodine and sodium iodide dissolved in a 50-50 solution of ethanol and water.
Silica	Silicon (IV) oxide (SiO_2)
Blue copper	Copper sulphate-5-water ($CuSO_4 \bullet 5H_2O$)
Rock salt	Naturally occurring sodium chloride (NaCl impure)
Red oxide	Lead tetraoxide (Pb_3O_4)
Chile salt petre	Sodium nitrate ($NaNO_3$)
Sugar	Sucrose ($C_{12}H_{22}O_{11}$)
Aspirin	Acetyl/salicyclic acid ($C_9H_8O_4$)
Baking soda	Sodium bicarbonate ($NaHCO_3$)
Vitamin C	Ascorbic acid ($C_6H_8O_6$)
Asbestos	Magnesium silicate ($MgSiO_3$)
Fluorspar (fluorite)	Calcium fluoride (CaF_2)
Soda lime	Mixture of sodium and calcium hydroxide ($Ca(OH)_2 + NaOH$)

SUMMARY

(a). Valency is the combining capacity of an element.
(b). Oxidation state of an element indicates the number of electrons lost, gained or shared by an atom of the element with respect to its neutral atom.
(c). The neutral atom has no charge.
(d). A radical is a group of elements which act like a single atom in forming compounds.
(e). Chemical bonding is a method by which atoms become more stable by donating, gaining or sharing electrons.
(f). The two main types of chemical bonds are covalent bond and electrovalent bond.
(g). A covalent bond involves sharing of electrons.
(h). An electrovalent bond involves transfer of electrons.
(i). The empirical formula of a compound is the simplest formula which expresses its composition by mass.
(j). Molecular formula is the formula which shows the actual number of each kind of atom present in a molecule.
(k). A binary compound is a chemical compound that contains exactly two different elements.

REVIEW QUESTIONS

1. Choose the most correct answer for the following questions;

(i). .... is a method of representing a molecule of a compound.
(A) Valency (B) Element
(C) A formula (D) Oxidation

(ii). The combining capacity of an element is called; ....
(A) Oxidation state (B) Valency
(C) Mole (D) Ion

(iii). A group of elements which acts like a single atom in forming compounds is called;................................
(A) Molecule (B) Compound
(C) Radical (D) Formula.

(iv). The formula of copper (II) chloride is..............
(A) $CuCl$ (B) Cu_2Cl (C) Cu_2Cl_2 (D) $CuCl_2$

(v). The system of naming compounds is also known as; ....
(A) Nomenclature (B) Naming methodology
(C) Classification (D) Binary compound

2. Match the items in LIST A with their correct statement in LIST B

List A	List B
(i). Oxidation state ________ (ii). Covalent bond _______ (iii). Electrovalent bond_____ (iv). Binary compound _____ (v). Empirical formula_______	A. Sharing of electrons B. Transfer of electrons C. A chemical compound that contains exactly two different elements. D. The simplest formula which expresses its composition by mass E. The formula which shows the actual number of each kind of atom present in one molecule of a compound. F. The number of electrons lost, gained or shared by an atom

3. Write **TRUE** for a correct statement and **FALSE** for an incorrect statement.
   (i). The melting and boiling points of covalent compounds are usually low............
   (ii). Electrovalent compounds are usually liquids or gases at room temperature
   (iii). Electrovalent compounds are generally crystalline solids at room temperature...................
   (iv). Electrovalent compounds have high melting and boiling points ....
   (v). Electrovalent compounds are generally soluble in water ....
   (vi). Covalent compounds consist of ions that are positively and negatively charged..................
   (vii). Covalent compounds are usually crystalline solids at room temperature.......

4. Name the elements which have these symbols.
   K, Na, Ag, Pb, Zn, Ba, Sn, Ar, P and Hg
5. Calculate the oxidation number of the underlined elements in the following compounds or radicals;
   a. $Na_3\underline{P}O_4$
   b. $\underline{S}O_3^{2-}$
   c. $\underline{N}O_2^{-}$
6. Write the formulae of the following compounds;
   (a) Iron (III) chloride (b) Lead (II) oxide
   (c) Potassium nitrate (d) Barium hydroxide
   (e) Potassium hydrogen carbonate
7. Here is a list of formulae for some iron compounds;
   Fe_2O_3, $FeSO_4$, $FeCl_3$, $Fe_2(SO_4)_3$

(a) Which compounds are binary compounds

(b)Which compound contains iron in oxidation state +2?

8. From the following list, O_2, Al, Na^+, CH_4, I^-, N select;

(a) Two atoms (b) two molecules (c) two ions.

9. Name the following compounds;

(a) MgI_2 (b) N_2O_5 (c) CCl_4

(d) $FeBr_2$ (e) Cu_3P_2 (f) H_2S

(g) Hg_2Cl_2 (h) K_2O (i) PCl_5

(j) SF_6

10. Explain the differences between the properties of electrovalent and covalent compounds using sodium chloride and water as examples.

11. Copy and complete this table;

Formula	Systematic (IUPAC) name
Cu_2O	
FeS	
$Pb\ (NO_3)_2$	
$Fe_2(SO_4)_3$	
$KClO_3$	
$Cu(NO_3)_2$	
MnO_2	

Questions for Revision and Practice

1. (i) Mention at least ten products that can be produced through application of chemistry knowledge
   (ii) Explain how the knowledge of chemistry is used to produce the named products in (i) above.
2. Explain how the following professionals can apply the knowledge of chemistry in their respective fields: Geologist, nurse, medical doctor, engineer, science teacher, laboratory technician, pharmacist, driver, farmer, horticulturists, floriculturists, and housewife.
3. (i) List at least ten examples of laboratory apparatus.
   (ii) Name the nature and use of each apparatus you have listed in (i) above
   (iii) Draw all the apparatus you have listed in (i) above
4. Explain how you could go about when rendering first aid to accidents of bruise, burns, fainting, shock, suffocation, choking, electric shock, severe bleeding, light bleeding and vomiting respectively
5. (i) Define first aid and first aid kit respectively..
   (ii) Mention at least eleven items found in a first aid kit.
   (iii) Explain the uses of items you have mentioned in (ii) above.
6. Draw chemical warning signs expected to be seen on the following items:
   (i) A can of petroleum
   (ii) A bottle of concentrated sulphuric acid
   (iii) A bottle of zinc granules
   (iv) A bottle of sodium lump
   (v) A bottle of uranium metal
   (vi) A container of mercury metal
   (vii) A container with hydrogen gas
7. (i) Define heat and flame respectively
   (ii) Draw well labeled diagrams of luminous and non-luminous flames respectively
   (iii) Outline five(5) differences between luminous and non-luminous flames
   (iv) Juma is working in a welding industry in Dar es Salaam. What type of flame is he supposed to use in his work?
   (v) Mwajuma is in a sitting room at home when the electric supply is cut off suddenly. Advise her on what type of flame she might use to illuminate the sitting room.
8. Explain how you can produce the hottest flame using a Bunsen burner
9. (i) What do you understand by control experiment?
   (ii) Design an experiment and explain how you would use scientific procedure to go about the experiment.

10. (i) Outline five (5) differences between Physical and Chemical changes.
(ii) In tabular form, name at least ten (10) examples of physical and chemical changes respectively.
11. Define:
(i) Brownian motion (ii) Particulate nature of matter
(iii) Diffusion (iv) Osmosis
(v) Kinetic molecular behavior
12. Define:
(i) Element (ii) Compound (iii) Mixture
(iv) Molecule (v) Homogenous mixture (vi) Heterogeneous mixture
13. Name and define three types of solutions as far as saturation is concerned.
14. (i) Define solutions, suspensions and emulsions respectively.
(ii) Mention at least two examples on the terms you have defined in (i) above
15. Name the method you would use to separate the following mixtures:
(i) Muddy water (ii) Mixture of salt and water
(iii) Mixture of sand and water (iv) Mixture of rice and water
(v) Mixture of ethanol and water (vi) Kerosene and water
(vii) Oil in ground nuts (viii)Tea leaves and water
(ix) Mixture of dyes (x) Protein and blood
16. (i) Mention and explain three components of a fire triangle.
(ii) With the aid of a chemical equation, define Rust and Rusting respectively.
(iii) Explain how you would go about in using a portable fire extinguisher in your school laboratory in case of fire outbreak.
(iv) Mention and explain at least five (5) methods of preventing rusting
17. Explain how Oxygen is important to the following people: housewife, mountain climber, deep sea diver, rocket scientist, premature baby, factory worker and laboratory technician.
18. Maduhu is working in a multi-product industry in Tanga. Explain how he will make use of hydrogen gas to produce different products.
19. (a) Define: (i) Water (ii) Water cycle
(iii) Water treatment (iv) Water purification
(b) (i) Explain how you can use water in day-to-day activities
(ii) Explain how you can purify domestic as well as you can treat urban water. Use one page narration.
20. Define Isotope and Isotopy respectively.
21. What do you understand by periodic table?
22. Define: (i) Bonding (ii) Chemical formula
(iii) Nomenclature (iv) Binary compound

Appendix A; Periodic table

non-metals

metals

Group I	II											III	IV	V	VI	VII	O
			$^{1}_{1}$H hydrogen														$^{4}_{2}$He helium
$^{7}_{3}$Li lithium	$^{9}_{4}$Be beryllium											$^{11}_{5}$B boron	$^{12}_{6}$C carbon	$^{14}_{7}$N nitrogen	$^{16}_{8}$O oxygen	$^{19}_{9}$F fluorine	$^{20}_{10}$Ne neon
$^{23}_{11}$Na sodium	$^{24}_{12}$Mg magnesium	transition elements										$^{27}_{13}$Al aluminium	$^{28}_{14}$Si silicon	$^{31}_{15}$P phosphorus	$^{32}_{16}$S sulphur	$^{35\cdot5}_{17}$Cl chlorine	$^{40}_{18}$Ar argon
$^{39}_{19}$K potassium	$^{40}_{20}$Ca calcium	$^{45}_{21}$Sc scandium	$^{48}_{22}$Ti titanium	$^{51}_{23}$V vanadium	$^{52}_{24}$Cr chromium	$^{55}_{25}$Mn manganese	$^{56}_{26}$Fe iron	$^{59}_{27}$Co cobalt	$^{59}_{28}$Ni nickel	$^{64}_{29}$Cu copper	$^{65}_{30}$Zn zinc	$^{70}_{31}$Ga gallium	$^{73}_{32}$Ge germanium	$^{75}_{33}$As arsenic	$^{79}_{34}$Se selenium	$^{80}_{35}$Br bromine	$^{84}_{36}$Kr krypton
$^{85}_{37}$Rb rubidium	$^{88}_{38}$Sr strontium	$^{89}_{39}$Y yttrium	$^{91}_{40}$Zr zirconium	$^{93}_{41}$Nb niobium	$^{96}_{42}$Mo molybdenum	$^{98}_{43}$Tc technetium	$^{101}_{44}$Ru ruthenium	$^{103}_{45}$Rh rhodium	$^{106}_{46}$Pd palladium	$^{108}_{47}$Ag silver	$^{112}_{48}$Cd cadmium	$^{115}_{49}$In indium	$^{119}_{50}$Sn tin	$^{122}_{51}$Sb antimony	$^{128}_{52}$Te tellurium	$^{127}_{53}$I iodine	$^{131}_{54}$Xe xenon
$^{133}_{55}$Cs caesium	$^{137}_{56}$Ba barium	$^{139}_{57}$La lanthanum	$^{178\cdot5}_{72}$Hf hafnium	$^{181}_{73}$Ta tantalum	$^{184}_{74}$W tungsten	$^{186}_{75}$Re rhenium	$^{190}_{76}$Os osmium	$^{192}_{77}$Ir iridium	$^{195}_{78}$Pt platinum	$^{197}_{79}$Au gold	$^{201}_{80}$Hg mercury	$^{204}_{81}$Tl thallium	$^{207}_{82}$Pb lead	$^{209}_{83}$Bi bismuth	$^{210}_{84}$Po polonium	$^{210}_{85}$At astatine	$^{222}_{86}$Rn radon
$^{223}_{87}$Fr francium	$^{226}_{88}$Ra radium	$^{227}_{89}$Ac actinium															

$^{140}_{58}$Ce cerium	$^{141}_{59}$Pr praesodimium	$^{144}_{60}$Nd neodimium	$^{147}_{61}$Pm promethium	$^{150}_{62}$Sm samarium	$^{152}_{63}$Eu europium	$^{157}_{64}$Gd gadolinium	$^{159}_{65}$Tb terbium	$^{162}_{66}$Dy dysprosium	$^{165}_{67}$Ho holmium	$^{167}_{68}$Er erbium	$^{169}_{69}$Tm thulium	$^{173}_{70}$Yb ytterbium	$^{175}_{71}$Lu lutecium
$^{232}_{90}$Th thorium	$^{231}_{91}$Pa protoactinium	$^{238}_{92}$U uranium	$^{237}_{93}$Np neptunium	$^{242}_{94}$Pu plutonium	$^{243}_{95}$Am americium	$^{247}_{96}$Cm curium	$^{247}_{97}$Bk berkelium	$^{251}_{98}$Cf californium	$^{254}_{99}$Es einsteinium	$^{253}_{100}$Fm fermium	$^{256}_{101}$Md mendelevium	$^{254}_{102}$No nobelium	$^{257}_{103}$Lw lawrencium

Appendix B. A table of elements

Name	Symbol	Atomic number	M.P (°C)	B.P (°C)	Density	Discovery (Year)	Ionization energy (eV)
Hydrogen	H	1	-259	-253	0.09 (g/l)	1766	13.5984
Helium	He	2	-272	-269	0.18 (g/l)	1895	24.5874
Lithium	Li	3	180	1347	0.53 (g/cm^3)	1817	5.3917
Beryllium	Be	4	1287	2469	1.85 (g/cm^3)	1797	9.3227
Boron	B	5	2300	2550	2.46 (g/cm^3)	1808	8.298
Carbon	C	6	3500	4827	2.26 (g/cm^3)	Ancient	11.2603
Nitrogen	N	7	-210	-196	1.25 (g/l)	1772	14.5341
Oxygen	O	8	-218	-183	1.43 (g/l)	1774	13.6181
Fluorine	F	9	-220	-188	1.7 (g/l)	1886	17.4228
Neon	Ne	10	-249	-246	0.9 (g/l)	1898	21.5645
Sodium	Na	11	98	883	0.97 (g/cm^3)	1807	5.1391
Magnesium	Mg	12	649	1090	1.74 (g/cm^3)	1755	7.6462
Aluminium	Al	13	660	2467	2.7 (g/cm^3)	1825	5.9858
Silicon	Si	14	1410	2355	2.33 (g/cm^3)	1824	8.1517
Phosphorus	P	15	44	280	1.82 (g/cm^3)	1669	10.4867
Sulphur	S	16	113	445	1.96 (g/cm^3)	Ancient	10.36
Chlorine	Cl	17	-101	-35	3.21 (g/l)	1774	12.9676
Argon	Ar	18	-189	-186	1.78 (g/l)	1894	15.7596
Potassium	K	19	64	774	0.86 (g/cm^3)	1807	4.3407
Calcium	Ca	20	839	1484	1.55 (g/cm^3)	1808	6.1132
Scandium	Sc	21	1539	2832	2.99 (g/cm^3)	1879	6.5615
Titanium	Ti	22	1660	3287	4.54 (g/cm^3)	1791	6.8281
Vanadium	V	23	1890	3380	6.11 (g/cm^3)	1830	6.7462
Chromium	Cr	24	1857	2672	7.19 (g/cm^3)	1797	6.7665
Manganese	Mn	25	1245	1962	7.43 (g/cm^3)	1774	7.434
Iron	Fe	26	1535	2750	7.87 (g/cm^3)	Ancient	7.9024
Cobalt	Co	27	1495	2870	8.9 (g/cm^3)	1735	7.881
Nickel	Ni	28	1453	2732	8.9 (g/cm^3)	1751	7.6398
Copper	Cu	29	1083	2567	8.96 (g/cm^3)	Ancient	7.7264
Zinc	Zn	30	420	907	7.13 (g/cm^3)	Ancient	9.3942
Gallium	Ga	31	30	2403	5.91 (g/cm^3)	1875	5.9993
Germanium	Ge	32	937	2830	5.32 (g/cm^3)	1886	7.8994
Arsenic	As	33	81	613	5.72 (g/cm^3)	Ancient	9.7886
Selenium	Se	34	217	685	4.79 (g/cm^3)	1817	9.7524
Bromine	Br	35	-7	59	3.12 (g/cm^3)	1826	11.8138
Krypton	Kr	36	-157	-153	3.75 (g/l)	1898	13.9996
Rubidium	Rb	37	39	688	1.53 (g/cm^3)	1861	4.1771
Strontium	Sr	38	769	1382	2.63 (g/cm^3)	1790	5.6949
Yitrium	Y	39	1523	3337	4.47 (g/cm^3)	1794	6.2173

Name	Symbol	Atomic number	M.P (°C)	B.P (°C)	Density	Discovery (Year)	Ionization energy (eV)
Zirconium	Zr	40	1852	4377	6.51 (g/cm^3)	1789	6.6339
Niobium	Nb	41	2468	4927	8.57 (g/cm^3)	1801	6.7589
Molybdenum	Mo	42	2617	4612	10.22 (g/cm^3)	1781	7.0924
Technetium	Tc	43	2200	4877	11.5 (g/cm^3)	1937	7.28
Ruthenium	Ru	44	2250	3900	12.37 (g/cm^3)	1844	7.3605
Rhodium	Rh	45	1966	3727	12.45 (g/cm^3)	1803	7.4589
Palladium	Pd	46	1552	2927	12.02 (g/cm^3)	1803	8.3369
Silver	Ag	47	962	2212	10.5 (g/cm^3)	Ancient	7.5762
Cadmium	Cd	48	321	765	8.65 (g/cm^3)	1817	8.9938
Indium	In	49	157	2000	7.31 (g/cm^3)	1863	5.7864
Tin	Sn	50	232	2270	7.31 (g/cm^3)	Ancient	7.3439
Antimony	Sb	51	630	1750	6.68 (g/cm^3)	Ancient	8.6084
Tellurium	Te	52	449	990	6.24 (g/cm^3)	1983	9.0096
Iodine	I	53	114	184	4.94 (g/cm^3)	1811	10.4513
Xenon	Xe	54	-112	-108	5.9 (g/l)	1898	12.1298
Caesium	Cs	55	29	678	1.87 (g/cm^3)	1860	3.8939
Barium	Ba	56	725	1897	3.51 (g/cm^3)	1808	5.2117
Lanthanum	La	57	920	3469	6.15 (g/cm^3)	1939	5.4569
Cerium	Ce	58	795	3257	6.69 (g/cm^3)	1803	5.5387
Praseodymium	Pr	59	935	3127	6.64 (g/cm^3)	1885	5.473
Neodymium	Nd	60	1010	3127	7.01 (g/cm^3)	1885	5.525
Promethium	Pm	61	1100	3000	7.3 (g/cm^3)	1945	5.582
Samarium	Sm	62	1072	1900	7.35 (g/cm^3)	1879	5.6437
Europium	Eu	63	822	1597	5.24 (g/cm^3)	1901	5.6764
Gadolinium	Gd	64	1311	3233	7.9 (g/cm^3)	1880	6.1501
Terbium	Tb	65	1360	3041	8.23 (g/cm^3)	1843	586.38
Dysprosium	Dy	66	1412	2562	8.55 (g/cm^3)	1886	5.9389
Holmium	Ho	67	1470	2720	8.79 (g/cm^3)	1867	6.0215
Erbium	Er	68	1522	2510	9.07 (g/cm^3)	1842	6.1077
Thulium	Tm	69	1545	1727	9.32 (g/cm^3)	1879	6.1843
Ytterbium	Yb	70	824	1466	6.57 (g/cm^3)	1878	6.2542
Lutetium	Lu	71	1656	3315	9.84 (g/cm^3)	1907	5.4259
Hafnium	Hf	72	2150	5400	13.31 (g/cm^3)	1923	6.8251
Tantalum	Ta	73	2996	5425	16.65 (g/cm^3)	1802	7.5496
Tungsten	W	74	3410	5660	19.25 (g/cm^3)	1783	7.864
Rhenium	Re	75	3180	5627	21.04 (g/cm^3)	1925	7.8335
Osmium	Os	76	3045	5027	22.6 (g/cm^3)	1803	8.4382
Iridium	Ir	77	2410	4527	22.65 (g/cm^3)	1803	8.967
Platinum	Pt	78	1772	3827	21.09 (g/cm^3)	1735	8.9587
Gold	Au	79	1064	2807	19.32 (g/cm^3)	Ancient	9.2255

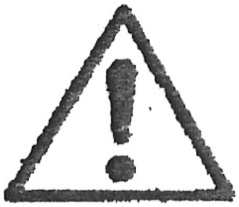
Name	Symbol	Atomic number	M.P (°C)	B.P (°C)	Density	Discovery (Year)	Ionization energy (eV)
Mercury	Hg	80	-39	357	13.55 (g/cm^3)	Ancient	10.4375
Thallium	Ti	81	303	1457	11.85 (g/cm^3)	1861	6.1082
Lead	Pb	82	327	1740	11.34 (g/cm^3)	Ancient	7.4967
Bismuth	Bi	83	271	1560	9.78 (g/cm^3)	Ancient	7.2856
Polonium	Po	84	254	962	9.2 (g/cm^3)	1898	8.417
Astatine	At	85	302	337	7.0 (g/cm^3)	1940	9.3
Radon	Tn	86	-71	-62	9.73 (g/l)	1900	10.7185
Francium	Fr	87	27	677	1.87 (g/cm^3)	1939	4.0727
Radium	Ra	88	700	1737	5 (g/cm^3)	1898	5.2754
Actinium	Ac	89	1050	3200	10.07 (g/cm^3)	1899	5.17
Thorium	Th	90	1750	4790	11.72 (g/cm^3)	1829	6.3067
Protactinium	Pa	91	1568	4027	15.4 (g/cm^3)	1913	5.89
Uranium	U	92	1132	3818	19.05 (g/cm^3)	1789	6.1947
Neptunium	Np	93	640	3902	20.45 (g/cm^3)	1940	6.2657
Plutonium	Pu	94	640	3235	19.84 (g/cm^3)	1940	6.0262
Americium	Am	95	994	2607	13.67 (g/cm^3)	1944	5.9738
Curium	Cm	96	1340	3100	13.5 (g/cm^3)	1944	5.9915
Berkelium	Bk	97	986	2900	14.78 (g/cm^3)	1949	6.1979
Californium	Cf	98	900	1472	15.1 (g/cm^3)	1950	6.2817
Einsteinium	Es	99	860	____	8.84 (g/cm^3)	1952	6.42
Fermium	Fm	100	1527	____	____	1952	6.5
Mendelevium	Md	101	827	____	____	1955	6.58
Nobelium	No	102	827	____	____	1958	6.65
Lawrencium	Lw	103	1627	____	____	1961	4.9
Unnilquadium	Unq	104	2100	5500	23 (g/cm^3)	1964	5.79

APPENDIX C: HAZARD WARNING SYMBOLS


OXIDISING


HARMFUL/IRRITANT


DANGEROUS FOR THE ENVIRONMENT


GENERAL DANGER

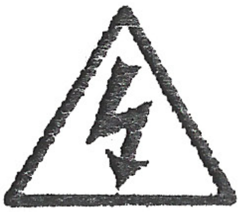

RISK OF FIRE

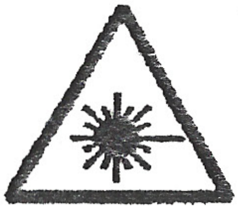

RISK OF EXPLOSION


TOXIC HAZARD


CORROSIVE SUBSTANCE


RISK OF IONISING RADIATIONS


RISK OF ELECTRIC SHOCK


LASER BEAM


NON-IONISING RADIATION


CHEMICAL WEAPON


BIOHAZARD


NO OPEN FLAME

GLOSSARY

1. Acid – a substance that produce hydrogen ions (H^+) or hydroxonium ions (H_3O^+) in solution.
2. Acidic oxides – Oxides of non-metals which react with water to form acidic solution.
3. Alkali- a soluble base
4. Alkali earth metals- group II elements
5. Alkali metals- group I elements
6. Analysis- the process of examining data in detail in order to understand it.
7. Anhydrous- describe a substance that contains no water.
8. Anion- an ion with a negative charge.
9. Anode- positively charged electrode.
10. Anthocyanin- the pigment in the juice of red cabbage that can be used as an indicator.
11. Aqueous solution- a solution obtained by dissolving a solute in water (solvent).
12. Apparatus- special tools and equipment used in a laboratory.
13. Atmospheric pressure- force per unit area exerted on object on the earth as a result of the attraction of the earth.
14. Atmosphere- the mass of air surrounding the earth.
15. Atom- the smallest particle that retain the property of the element.
16. Atomic number- the number of protons in an atom.
17. Atomic mass - the weighted average of the atomic masses for the mixture of all naturally occurring isotope of an element.
18. Atomic mass unit (amu)- the unit for expressing relative masses of atoms with reference to one- twelfth of the mass of a carbon-12 atom.
19. Atomic radius - The distance between the nucleus of an atom and the outermost stable energy level.
20. Balance-Apparatus for weighing substances
21. Barrel - the cylindrical tube that is the upper part of a Bunsen burner
22. Base - A metal oxide or hydroxide that neutralize an acid to form salt and water only.
23. Basic oxide-An oxide of metal which react with water to form alkaline solutions.
24. Beaker-lipped cylindrical glass or plastic container used for holding, treating and mixing liquids
25. Biogas - Gas produced by the decay of organic matter in the absence of air.
26. Binary compound-A compound which contain two types of ions
27. Biotechnology-The use of living organism such as bacteria, fungi, and cells from plants and animals for industrial and scientific purpose to make product
28. Boiling point-The temperature at which a liquid boils to form vapour
29. Brownian motion-An irregular motion of tiny particles suspended in fluids
30. Bunsen burner-A small adjustable gas burner commonly used in laboratories.
31. Cardiopulmonary resuscitation (CPR)-Ultimately applying chest compression and lung ventilations to a person having difficulty in breathing.
32. Catalyst - A substance that speeds up the rate of a chemical reaction but itself remains unchanged at the end of reaction.
33. Cathode-The negatively charged electrode.
34. Cation-positively charged ion
35. Chemical bonds-force of attraction

that hold atoms or ions together in chemical compound

36. Chemical equation-symbolic representation of a chemical reaction showing chemical formulae of substance present before and after a chemical reaction.
37. Chemical change –An irreversible change that affects the chemical properties of substances and there is formation of new products.
38. Centrifugation-spring a sample around very quickly in a machine called a centrifuge, which causes sediment to settle and separate quickly from the liquid
39. Chromatography-A method of separating two or more solute from the liquid they are dissolved in: Example, separating dyes in colored ink
40. Chemical warning signs-safety symbol found on chemical containers.
41. Chemist-A person who studies chemistry
42. Chemistry-The branch of science which deals with composition, decomposition, structure and properties of matter.
43. Collar –A band near the base of a Bunsen burner that can be rotated to control air supply to the burner.
44. Colloid-contains very small particles of one substance dispersed through out another substance usually in different state. Example: Tiny particles of a solid dispersed in a liquid.
45. Combustion- burning of a substance in the presence of oxygen.
46. Combustible- a substance which ignites and burns easily
47. Combination reaction- two or more substances combine to form a single compound.
48. Composition- the way in which a substance is made of different parts.
49. Compound - a pure substance made up of two or more kinds of elements that are chemically combined together in fixed proportions.
50. Compressible- capable of being condensed or reduced in volume.
51. Concentrated solution- a solution containing a relatively large quantity of the solute.
52. Concentration- a measure of the quantity of solute dissolved in a particular volume of solution.
53. Conclusion- summary of the results of an experiment.
54. Condensation- the cooling of a vapour to form a liquid.
55. Contaminant- any potentially harmful substance in the environment.
56. Control experiment- in a control experiment all the independent variables are kept constant.
57. Coordinate covalent bond- a chemical bond formed when one atom donates both of the electrons shared in covalent bond.
58. Covalent bond- a pair of electrons shared by two atoms in a molecule.
59. Crystal- a solid having plane faces at definite angles and whose atoms, ions, or molecules have a regular three dimensional arrangement.
60. Controlled variable- the factor in an experiment that is kept constant and does not affect the outcome of the experiment.
61. Covalency- the process of sharing electrons by which many non-metals and sometimes certain metals, join to form compounds.
62. Cosmetics- products used for cleaning, beautifying, promoting attractiveness or altering ones appearance.
63. Crystallization- the formation of

crystals from a homogenous solution.

64. Crystallization point- the temperature at which crystals form from homogenous solution.
65. Data- information that is analyzed so as to make decisions after experimentation.
66. Decantation- pouring off a liquid into a separate container, leaving sediment behind.
67. Dependent variables -Variables that cannot be changed by an investigator or researcher.
68. Decomposition- breaking down of chemical compound into elements or smaller compounds
69. Decomposition reaction- a compound breaking down into two or more simpler substance
70. Density- the degree of compactness of a substance.
71. Detergent- liquid or solid compound or mixture of compounds used to assist in cleaning.
72. Diatomic elements - Elements which form molecules composed of two atoms.
73. Diffusion- process by which a gas spreads out to fill all available space. A solute dissolved in a solvent also diffuses.
74. Disinfect- to make clean and free from infection, especially by use of a chemical.
75. Dissolving- process by which a solute breaks down into tiny particles and spreads throughout the solvent.
76. Displacement reaction- a reaction in which one substance takes the place of another.
77. Dissociation- the process by which ionic compounds separate or split into smaller molecules, ions or radicals.
78. Distillation- separating out a pure liquid from a solute dissolved in it by boiling.
79. Drugs- substances used to treat or prevent diseases, diagnose illnesses or relieve pain.
80. Effervescence- the escape of gas from a liquid.
81. Electrode- a conductor through which an electric current enters or leaves an electrolyte.
82. Electrolysis- the decomposition of an electrolyte by an electric current.
83. Electronegativity- ability of an atom to attract an electron.
84. Electron- the negatively charged particle of an atom.
85. Electronic configuration- symbolic representation of the electron arrangement in sublevels of an atom using the electrons.
86. Electronegative elements - elements like fluorine, oxygen and nitrogen that have a very strong attraction for electrons involved in a chemical bond.
87. Electrovalent bonding- force of attraction between ions.
88. Electrolyte- a solution or molten substance that will conduct electricity and can be electrolyzed.
89. Element- a pure substance that cannot be split into simpler substances by a simple chemical process.
90. Electron shell- a set of energy levels outside the nucleus that contains the atom's electrons.
91. Electrovalency - the process of exchange of electrons which metals and some non-metals join to form compounds.
92. Experiment- controlled investigation used to test or obtain facts, or establish a hypothesis, or illustrate a known scientific law.
93. Empirical formula- the formula that

represents the simplest ratio of the atoms or ions in compound.

94. Emulsifying agent- a substance that can be added to an emulsion to prevent it from separating into the two component liquids.
95. Emulsion- a mixture of liquids that do not mix well.
96. Energy levels- special regions in which electrons orbit the nucleus of an atom.
97. Evaporation- process by which molecules, usually from a liquid, leave the liquid surface and become gaseous.
98. Environment - the surroundings in which something exists or lives.
99. Filtrate- clear solution that passes through the filter paper or other filtration medium during filtration.
100. Filtration- process of separating an insoluble solid from a solution or liquid.
101. Fertilizers- chemicals or natural substances added to soil to increase its fertility.
102. Fire extinguisher- apparatus used for putting out hazardous fires.
103. Fire triangle- three components required to start a fire.
104. First aid- the help given to a sick or injured person before medical assistance from hospital.
105. First aid kit- a small box in which equipment and chemicals needed for the first aid are kept.
106. Flame- a burning gas which produces heat and light.
107. Formula unit- the specific group of atoms or ions symbolized and expressed in the chemical formula.
108. Formula- a collection of symbols and numbers, representing the composition of a compound.
109. Fossil fuel- fuels such as coal and oil originated from the remains of living organisms in the ground millions of years ago.
110. Food supplement- a nutrient added to a food stuff to enhance it.
111. Fractional distillation- separation of two or more liquids of different boiling points by boiling and using a fractionating column.
112. Fractionating column- apparatus used in the distillation of liquid mixtures to separate them into their components or fractions.
113. Fuels- a substance that burns to produce energy.
114. Fume chamber- a ventilated enclosure in a chemistry laboratory in which harmful experiments involving poisonous gases and vapours are carried out.
115. Galvanize- to coat iron or steel with zinc in order to prevent rusting.
116. Gas- a fluid with no definite shape or volume but it is able to expand indefinitely.
117. Group- a vertical column of elements in the periodic table
118. Haber process- the industrial production of ammonia on a large scale by reacting hydrogen with nitrogen.
119. Halogens- group VII elements.
120. Hard water - Water that does not easily form lather with soap.
121. Hazard- something that is dangerous and can cause damage.
122. Heat- energy transferred from one body to another by thermal interactions.
123. Heimlich manouvre - first aid procedure for dislodging an obstruction from a person's throat.
124. Heterogeneous- a substance that is not the same all the way through.
125. Homogenous- a substance that is the same all the way through.

126. Hydrated- a substance which contains water molecules.
127. Hypothesis- a proposal or idea put forward to explain something, but not yet proved.
128. Immiscible- liquids that when mixed do not dissolve in each other to form a solution.
129. Incinerator- A facility in which the combustion of solid wastes take place.
130. Insoluble - a substance that does not dissolve to any easily detectable extent.
131. Independent variable- the factor in an experiment that is manipulated or changed to obtain values being measured.
132. Indicator- a substance that exhibits different colours in solutions of different acidity and basicity.
133. Interpretation- an explanation of the meaning or importance of something.
134. Ion- the electrically charged particle formed when an atom or group of atoms either gains or loses electrons.
135. Ionic compound- A chemical substance that is made up of positive and negative ions.
136. Ionic equation- a balanced chemical equation that shows all water soluble ionic substances written in ionic forms, while insoluble solids and covalently bonded substances are written in molecular form.
137. Ion exchange resin- small beads made of special plastic that remove hardness of water in an ion exchange.
138. Ion exchanger- the device that removes both types of hardness of water by removing all calcium and magnesium ions.
139. Ionization energy- the energy required to remove elections from an atom or ion.
140. Irreversible reaction- a reaction that can only proceed in one direction.
141. Isotopes- atoms of a particular kind of element that have different numbers of neutrons and therefore different mass numbers.
142. IUPAC- The abbreviation for the international union of pure and applied chemistry, which deals with chemical nomenclature.
143. Kinetic molecular behavior- the manner in which particles in matter behave.
144. Laboratory- a special room or building used for carrying out scientific experiments.
145. Law- a rule that must be followed in scientific studies.
146. Limestone- rocks that are mainly composed of insoluble calcium carbonate.
147. Liquefy- to make or to become liquid.
148. Liquid- a fluid with a definite volume and takes the shape of the container holding it.
149. Luminous- bright and full of light
150. Mass- a measure of the quantity of matter in an object.
151. Mass number- the total number of proton plus neutrons in the nucleus of an atom.
152. Matter- anything that has mass and occupies space.
153. Metallic bonding- the bonding of metal atoms in solids where positive metal ions are arranged in a regular three-dimensional way and the loosely held valence electrons are able to move freely throughout the crystal.
154. Metallic solid- a solid that is composed of metal atoms.
155. Metals- shiny, dense substances that conduct electricity well and usually have high melting points.
156. Metalloids- elements that are neither

completely metallic nor completely non-metallic in their properties.

157. Miscible-liquids that mix very well.
158. Mixture- physical combination of two or more substances in any ratio.
159. Molar mass- the mass in grams of a mole of any substance that is the sum of the atomic masses of all the atoms represented in the formula.
160. Molecular formula-the chemical formula that gives the actual number of atoms of each kind present in a molecule of the substance.
161. Molecular weight- the sum of the atomic masses of all atoms in one molecule of a particular compound.
162. Molecule- the smallest part of a substance that can take part in chemical reaction and still retain the properties of that substance
163. Neutralization- the reaction between an acid and a base to give salt and water.
164. Neutron- an electrically neutral particle with a mass of 1 amount that is found in the nucleus of atoms.
165. Noble gases- the chemically stable non-metal elements of group VIII of the periodic table.
166. Nomenclature - the system of names and formula used to identify all chemicals.
167. Non-luminous flame- the blue flame of burner with an adequate supply of oxygen and which produces more heat than the luminous flame.
168. Non-metals- group of elements that do not conduct electricity (except granite).
169. Nucleons- sub-atomic particles found in the nucleus of an atom, namely proton, and neutrons.
170. Nucleus- the tiny dense centre of an atom that contains an atom's protons and neutrons.
171. Nuclide notation- representation of an atom of an element by its chemical symbol, having its mass number on the upper left and it atomic number on the lower left of the symbol.
172. Octet of electrons- used to describe the complete set of eight valence electrons that are present in the outer most energy level of noble gases that follow helium in the periodic table.
173. Octet rule- the tendency for a non-metal atom to gain or share electrons until it has eight valence electrons.
174. Oxidation- the process that occurs when substance combines with oxygen or any chemical process where there is a loss of electrons.
175. Oxidation number- (also called oxidation state) the charge of a simple ion or the "apparent charge" assigned to an atom within a compound or polyatomic ion.
176. Oxidizing agent- a substance that adds oxygen to, removes hydrogen from or loses electrons to, another substance.
177. Oxy-hydrogen flame- a very hot flame produced by the combustion of mixture of oxygen and hydrogen.
178. Particulate nature- relating or in the form of very tiny particles.
179. Percent by mass- the mass of solute divided by the total mass of solution times 100%.
180. Percent by volume- the volume of solute divided by the total volume of solution times 100%.
181. Percent by weight- the number of grams of a particular substance that can be found in 100 grams of the sample.
182. Percent composition- a list of the percents by mass of each element in a compound.

183. Period of elements- a horizontal row of elements in the periodic table.
184. Periodic law- there is a periodic variation in the physical and chemical properties of elements when they are arranged in order of increasing atomic number.
185. Periodic table- an arrangement of all known elements in order of increasing atomic number , set up according to the order in which electron shells are filled and so that elements with similar properties fall into the same vertical column.
186. Periodicity- the regular periodic changes of elements according to their atomic numbers.
187. pH value- a value that indicates the concentration of hydrogen ions in a solution.
188. Pharmacy- a facility or location where drugs and other medical services are sold or dispensed.
189. Physical change- a temporary change in the outward or physical properties of a substance, usually fairly easy to reverse.
190. Pollutant- a contaminant introduced into the environment.
191. Potable- describes water that is safe for people to drink.
192. Precipitate- a solid that forms and separates out of solution as the result of a chemical reaction.
193. Precipitation- water that falls to the ground in the form of rain, snow or hail.
194. Precipitation reaction- a reaction in which two soluble compounds combine to give a soluble compound and an insoluble compound.
195. Prediction- a kind of forecast which suggests what the result is likely to be.
196. Principle- has got the same meaning as law in science.
197. Problem- a scientific question to be answered.
198. Products- the chemical substances formed during a chemical reaction.
199. Properties- attributes or settings that are used to describe an object.
200. Proton- positively charged particle with a mass of 1 a.m.u found in the nucleus of all atoms.
201. Pure substance- an element or compound composed of the same kind of matter with the same kind of particles throughout.
202. Radical- a group of combined atoms that has a positive or negative charge and tends to remain together as unit during chemical reactions.
203. Reactants- the starting materials that enter into a chemical reaction and they appear on the left-hand side of a chemical equation.
204. Reactivity- the vigour with which an element reacts with other substances.
205. Reagent- a substance that takes part in chemical reaction.
206. Recovering position- a position used to prevent an unconscious person from choking in which the body is placed face downwards and slightly to the side, supported by bent limbs.
207. Redox reaction- a reaction in which both reduction and oxidation take place.
208. Reducing agent- a substance that removes oxygen from, gives hydrogen to, or gains electrons from, another substance.
209. Reducing sugar- a sugar that must have available an aldehyde group that can be oxidized to an acid by a mild oxidizing agent.
210. Reduction- the process that occurs when a substance loses oxygen or any

chemical process where there is gain of electrons.

211. Relative atomic mass (R.A.M)- The mass of an atom compared with the mass of an atom of carbon -12, which is taken exactly 12.00000 units.
212. Reservoir - a container designed to hold fluids such as water.
213. Residue- the insoluble material that remains on the filter paper after filtration.
214. Respiration- Breakdown of food substances in the cell to release energy.
215. Resuscitate- Reviving an unconscious person.
216. Reversible reaction- a reaction that can go in either direction of forward or backward.
217. Rust- The reddish-brown substance that forms on the surface of iron or steel due to a chemical reaction with water and air.
218. Rusting- the corrosion of iron and steel by water and oxygen from the air.
219. Sacrificial protection- the process of using some blocks of a more reactive metal to protect large objects made of iron or steel from rusting.
220. Saturated solution- a solution that contains as much solute as it can dissolve at a particular temperature.

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INDEX

A

B

C

D

E

F

G

H

I

K

L