Basic Principles

MSCE

Chemistry

Senior Secondary Students.

Students book Solution manual 3&4

Chuzu Elisha C

Basic Principles.

Chemistry.

Junior Secondary Students.

Form 3&4.

Self-Learning model.

Book Preparations.

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Table of Contents.

Foreword.

This book is a continuation of Basic Principle Chemistry for junior students. This contains special basics that will help students to understand Chemistry easily as they are studying to prepare Malawi School Certificate of Examination (MSCE). Most students consider this subject, Chemistry as the most difficult due to lack of resources. Therefore, this book has written the scientific language of Chemistry to its simplest form. This will help the students to conduct **self-learning process** through practicing the given questions provided at the end of each text. It is believed that ***when you do, you easily remember***. Therefore, the users (students) are advised to attempt all questions given after the texts for it helps to equip students with all necessary techniques and skills for answering examinations. This carries almost 80% preparation of both theoretical and practical examinations.

The book has been written following the research from various recommended texts book of Chemistry, unpublished past papers of physical science and different minds from students who have sincerely means the scope and this coverage during my experience when I was teaching. Their excellent questions they posed during the lesson have greatly assisted to clarify my methodologies. However, the concept has responded the Ministry of Education, Science and Technology Teaching Syllabus in Malawi.

The approach of the methodologies of this book and the way it has been organized, it is very suitable materials to be used by Chemistry students with the aim to develop self-learning widely from Community Day Secondary Schools (CDSS) to the national levels through. However, it can also be used as guide materials to any other students and teachers.

The basis of sciences is understanding and application; therefore, this book has also taken part in practical experiments, which is based on ***How to do rather than*** **what to do** because the Chemistry science makes sense when you have done the practical, which is the main body of any science course.

This is an examination-guiding book, therefore, the manual script, has interpreted the terminologies used by the most examiners when asking questions in order to help users on how to answer the questions guided by its terminology. More of these are, *explain, describe, discuss, define, state, give, list* and *mention* which are commonly used during examinations.

This is very inspired book to the students who wish to study both Veterinary and medical doctor, Chemical engineering, Agriculture and nursing in response to the National Development.

To God Be the Glory.

UNIT 1

Nitrogen, Sulphur & Phosphorus.

Learning objectives.

By the end of this unit, students should be able to**:**

a) Describe sources and properties of nitrogen.
b) Explain the uses of nitrogen and its compounds.
c) Describe sources and properties of Sulphur.
d) Explain the uses of Sulphur and its compounds.
e) Describe sources and properties of phosphorus.
f) Describe uses of phosphorus and its compounds.

1.10 Introduction.

In JCE Chemistry, you learnt about the Elements of the periodic table. Can you list some of these elements? The group and period to which the element belongs determine its properties. In this scope, we are going to concentrate on **Nitrogen, phosphorus** and **Sulphur**. These are very important elements and they form most of the inorganic compounds. Inorganic compounds are the compounds of non-bilogical origin.

Nitrogen, Phosphorus and Sulphur play a major role in pharmaceuticals and laboratories. In agriculture, plants also require the use of these elements for the health growth.

In this unit, we are going to discuss the sources, properties and the uses of Nitrogen, Phosphorus and Sulphur, and their compounds.

1.11 Nitrogen.

Nitrogen is an essential element, which makes up 78% of air. With the reference from the periodic table, what is the group and period of nitrogen element?

Obviously, nitrogen element is in group V and period 2 of the periodic table. This element is also non-metallic substance. The combination power **(valance) of** nitrogen is three. The nucleus of the nitrogen atom is presented as $7^{14}N$ where 14 is a ***mass number (atomic mass)***, 7 is an ***atomic number (protons***) and **N** is a symbol of an element, nitrogen.

1.12 Sources of nitrogen.

Nitrogen element is found in the following.

The air.

Nitrogen makes up 78% of air. Some of the bacteria convert it into nitrates. These bacteria are able to take nitrogen from the atmosphere and convert it into a form in which it can be used. This is done in the process called ***fixing nitrogen***.

The soil.

Nitrogen is found in nitrogen- containing compounds called *nitrate*, which are produced as a result of the effect of lighting and from dead plants and mammals as they decay.

1.13 Properties of nitrogen.

Nitrogen element has both physical and chemical properties.

Physical Properties.

Nitrogen elements have the physical properties as follows:

- It has no smell.
- It is colourless gas.
- Almost insoluble in water.
- Less dense than that of the air.

Chemical Properties.

The nitrogen gas exists as diatomic molecules ($\boldsymbol{N2}$) with its triple covalent bond($N \equiv \mathrm{N}$). This is very strong bond. Therefore, at room temperature, nitrogen gas is unreactive substance. As such, for a reaction to occur, more heat energy must be supplied **to break the bond** of nitrogen molecule to form single atom of nitrogen, which can go in reaction with the other elements. Nitrogen gas posess the different properties at high temperatue as follows:

(a). Reaction with alkali metals.

Nitrogen reacts with alkali metals to form nitrogen alkalides. For example:

Magnesium + Nitrogen $\longrightarrow$ Magnesium nitride

$6Mg_{s}\ 2N_{2g} \longrightarrow 2Mg_3N2$

(b). Reaction with Hydrogen.

Nitrogen React with hydrogen to form Ammonia under 300 atmospheric pressure and 450^0C heat. The equation below summarizes the condition.

Nitrogen + Hydrogen$\geq$ Ammonia

$N_{2g} + 3H_{sg} \geq 2NH_{3g}$

(c). Nitrogen has no effect on litmus paper hence it is a neutral. Its pH is 7.

(d).Nitrogen does not support combustion. Growing sprit fails to continue in a gas jar containing nitrogen.

1.14 Uses of nitrogen in everyday life.

Nitrogen is very important in everyday life. In industry and other places, nitrogen is used as follows;

- Making artificial fertilizer
- Liquid nitrogen is used as refrigerant
- Manufacturing of ammonia and nitric acid
- Nitrogen is used in food packaging to make food fresh since it keeps off oxygen.
- Filled in an empty oil tanker to prevent fire.

1.15 The compound of the nitrogen.

In JCE Chemistry, you learnt about compounds. Can you mention three compounds, which you discussed with your teacher?

Nitrogen React with other elements to form the different compounds. Of all the compounds, we are going to discuss two compounds, **Ammonia**($\boldsymbol{NH3}$) and **nitric acid**($\boldsymbol{HNO3}$). These compounds are very important in the field of chemistry and chemical engineering.

Ammonia.

This is the compound that contains nitrogen(N) and hydrogen (H) atom in its formula. Therefore, its molecularformula is (NH3). At room temperature, it exists an gaseous state. Ammonia can be prepared using is **industrial** and **laboratory** methods.

(a).Industrial preparation of ammonia

The industrial preparaton of ammonia is also called **Haber process.** It simply involves the reaction of nitrogen and hydrogen to form ammonia. For instance;

Hydrogen + Nitrogen $\gtrless$ Ammonia

$3H_2(g) + N_2(g) \gtrless 2NH_3$

The double arrow shows that, reaction is reversible.

Steps of Haber process.

Although the Haber process is a simple reaction of hydrogen and ammonia, but it goes through different steps as follows:

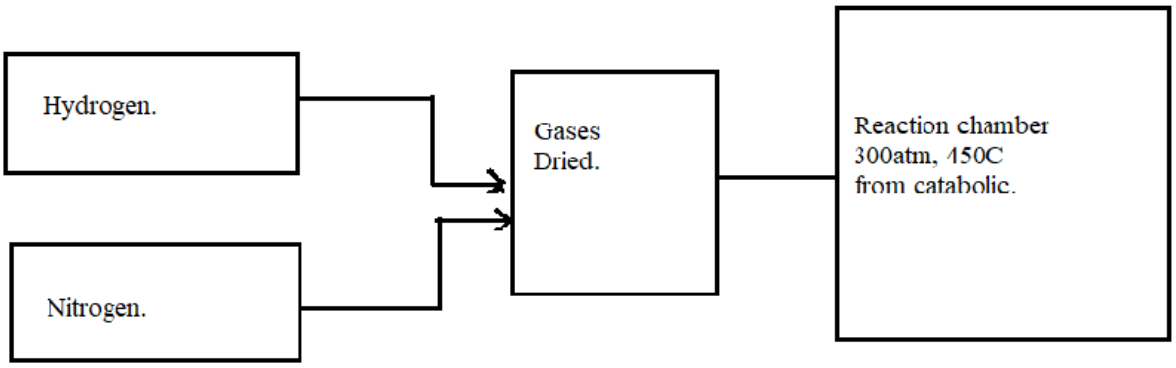
Step 1: Nitrogen obtained **(fraction distillation)** and hydrogen too obtained (reaction of nature gases e.g methane (CH4) and steam).

Step 2: Nitrogen and hydrogen are dired them mixed in the ration of ratio 3:1.

Step 3: At 300 atmosphere pressure and 450^0C, the mixture reacts to form Ammonia (NH_3)

Step 4: The mixture is cooled to liquefy the Ammonia and get it.

The figure summarise the process



NB: Iron catalyst with potassium and aluminium oxide mixed is used to speed up the reaction process.

(b).Laboratory prepation of ammonia.

Experiment

AIM: Preparation of ammonia gas.

Materials.

- Flasks.
- Calcium hydroxide
- Ammonium chloride.
- Red litmus paper.
- Drying tower.

Procedures.

a) Arrange the apparatus as shown in figure below:

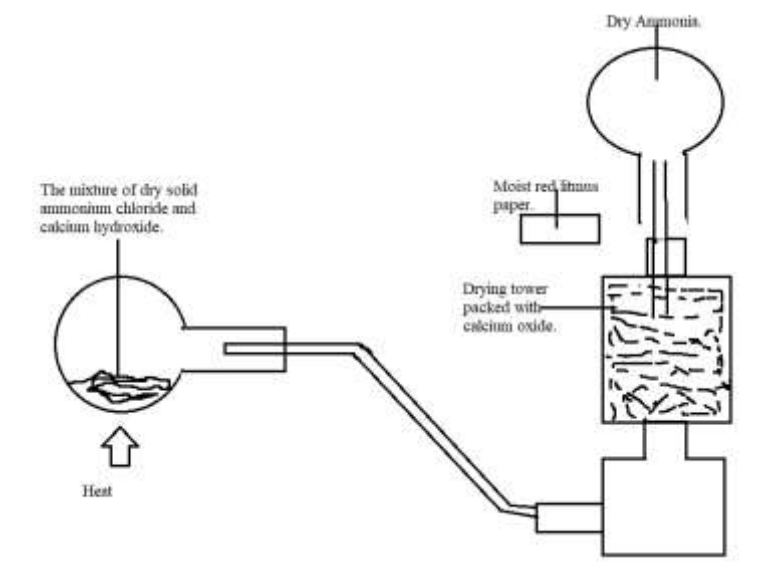


Figure 1.15 b. Laboratory preparation of Ammonia gas.

b) Heat the mixture.
c) Observe the colour of most red litmus paper.
d) Record the results in the table.

Observation.

The red litmus paper will turn blue.

Discussion.

Small quantities of ammonia gas can be produced by heating any ammonium salt, such as ammonium chloride, with an alkali, such as calcium hydroxide.

In the experiment, when the mixture is heated; Ammonium chloride and calcium hydroxide get reacted and produce calcium chloride, water and Ammonia gas.

The equation below summaries the process.

$$Ca(HO)_{2(s)} + 2NH_4Cl_{(s)} \longrightarrow CaCl_{2(s)} + NH_{3(o)} + H_2O_{(i)}$$

This reaction forms the basis of a chemical test to show that a compound contains the ammonium ion (NH4+). if any compound containing ammonium ion is heated with sodium hydroxide, ammonia gas is given off which turns damp red litmus paper blue.

Water vapour is removed from the ammonia gas by passing the gas formed through a drying power containg calcium oxide.

1.16 Uses of ammonia.

Just like nitrogen, ammonia is also useful as follows.

- Used in making plastics, dye and fibres
- Used in the manufacture of fertilizer
- Ammonia gas is used in the manufacture of nitric acid
- Used in making explosives
- It is used in the manufacture of the ammonia chloride used in the dry cells
- It is used to make hard water soft
- Liquid ammonia is used as refrigerant
- It is used as a solvent in cleaning e.g) in launderers)
- To manufacture ammonia carbonate used in smelling salts.

1.17 Nitric acid.

This is the compound that contains Hydrogen, Nitrogen and Oxygen. This has the molecular of HNO_3. Nitric acid is very important strong acid. However, in its concentrated form, it is powerful

acid which is an oxiding agent of both metals and non-metals. The compound of nitric acid is prepared mainly through ***Ostwald process.***

The Ostwald process.

Mainly, this process has three stages.

Step 1: At 230^0C ,the mixture of air and ammonia is heated . this produces nitrogen monoxide (NO) and water.

Ammonia +oxygen⟶ Nitrogen monoxide +water.

$4NH_{3(g)}+O_{2(g)}\rightarrow 4NO_{(g)}+6H_2O_{(l)}$ ▲

step2. The nitrogen monoxide gas react with oxygen from the air to form brown nitrogen dioxide gas

Nitrogen +Oxygen ⟶Nitrogen dioxide

$4NH_{2(g)}+O_{2(g)}\rightarrow 2NO_{2(g)}$

Step 3. The nitrogen dioxide NO is then dissolved in **water** to produce **nitric acid.**

Nitrogen dioxide +Water *Nitric Acid Nitrogen Monoxide.*

$3NO_{2(g)}+H_2O_{(l)}\rightarrow 2HNO_{(aq)}+NO_{(g)}$

Important note.

The **Nitrogen monoxide** and **Nitrogen dioxide** are very dangerous gases when dissolved into the atmosphere. They react with rainy water to form acid rain. Therefore, their release into the atmosphere should be controlled**.**

1.18 Uses of nitric acid in every day lifes.

The following are the uses of **Nitric acid**

- Used in manufacturing drugs
- Manufacturing nitrates fertilizer
- Manufacturing dyes, plastics and fibers
- Manufacturing explosive like trinitrotoluene, and dynamite
- Used in identification of gold
- Used as oxidizing agent in textile industries
- Used in refining gemstones
- Used as electrolyte in electrolysis.

Self-assessment 1.0

1 a state **two** sources of nitrogen.

b Explain why nitrogen gas is unreactive at low temperature?

c State **three** uses of nitrogen.

2.study the chemical equation below and use it to answer the questions given.

$3NO_{2(g)}+H_2O_{(l)} \rightarrow \mathbf{A}+NO_{(g)}$

a. name the process shown above
b. name the product labelled A
c. mention three uses of compound in 2b
d. state two factors that influence the reaction to occur very quickly

3.Give t**hree** uses of ammonia.

1.19 Sulphur.

This is non-metallic element which has very important role in chemical industry. This element is found in group VI and period 2 of periodic table. Sulphur is found in different forms and sources. This element is extracted from its deposit by the frash process. Its compound, sulphuric acid has also several industrial uses.

1.20 Sources of Sulphur.

Sulphur is found in the followings.

- Crude oil
- Natural gases
- Metal ores e.g. copper pyrites ($CUFeS_2$) and Zinc blende(ZnS)
- Volcanic regions.

1.21 Properties of sulphur.

Sulphur element has both physical and chemical properties.

Physical properties of Sulphur.

- Does not conduct electricity
- Insoluble in water
- Soluble in organic solvent e.g. benzene and toluene
- Yellow brittle solid at room temperature
- Has low melting and boiling point as compared to metals.

Chemical properties of Sulphur.

Sulphur posess different chemical properties at roo temperature as follows:

(a).Reaction with metals.

Sulphur reacts with metals to produce metal sulphides. For example,

Magnessium + Sulphur→ magnesium Sulphide

$Mg_{(g)}+ S_{(g)}\rightarrow MgS_{(s)}$

(b).Reaction with Oxygen.

Sulphur reacts with oxygen to form oxides. For example*;*

Sulphur +Oxygen→ Sulphur dioxide

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

(c).Reaction with nitric acids.

Sulphur reacts with nitric acids to produce sulphuric acid, water and nitrogen dioxide.

Sulphur + Nitric acid ⟶ sulphuric acid+ water+ nitrogen dioxide

$S_{(g)} + 6HNO_{(aq)} \rightarrow H_2SO4_{(aq)} +2H_2O_{(l)}+6NO_{2(g)}$

1.22 Uses of sulphur.

Sulphur is very important in industry, laboratory and other chemical studies.

- Production of sulphuric acid
- Volcanization (Hardening) of rubber
- Manufacturing of matches, gun powder, fireworks, plastic flowers and fungicides
- Manufacturing special concrete called Sulphur concrete.
- Manufacturing the drugs.

Sulphuric acid.

This is the compound of Sulphur that contains ***Hydrogen***, ***Sulphur*** and ***Oxygen***. Its molecular formula is **(H_2SO_4).** In industry, sulphuric acid has several uses. The compound is extracted from the ground by the process called **contact process**. This process is facillited by high temperature (450^0C) and other catalyst like Vansadium oxide (V_2O_5).

Steps in contact proces.

Step 1. *Sulphur reacts with oxygen to form Sulphur dioxide*

Sulphur +oxygen →Sulphur dioxide

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

Step 2: *Sulphur dioxide reacts with atmospheric oxygen to form Sulphur trioxide in the presence of catalyst (Vanadium Oxide, V_2O_5).* This kind of reactions takes place at the temperature of 450^0C.

Sulphur dioxide +atmos. Oxygen←→ Sulphur trioxide

$SO_{2(g)} + O_{2(g)} \rightarrow SO_{3(g)}$

The reaction is reversible. The Sulphur trioxide can dissolve to produce reactants back.

Step 3: *The Sulphur trioxide dissolves in concentrated sulphuric acid (98%) to form substance called OLEUM.*

Sulphur trioxide + sulphuric acid→OLEUM

$H_2SO_{4(aq)} + SO_{3(g)} \rightarrow H_2S_2O_{7(aq)}$

(OLEUM)

Step 4: *Oleum ($H_2S_2O_7$) is added to water in order to obtain sulphuric acid of required concentration.*

Oleum + water→sulphuric acid

$H_2S_2O_{7(aq)} + H_2O_{(l)} \rightarrow 2H_2SO_{4(aq)}$

Uses of sulphuric acid.

The uses of sulphuric acids are;

- Production of detergent and soap
- Tarring leather
- As an electrolyte in car battery
- Cleaning metals
- Manufacturing insecticides and drugs
- Manufacturing fertilizer
- Manufacturing paints, pigments and dyes
- Manufacturing plastics
- Extraction and treatment of metals
- Refining crude petroleum
- Manufacturing explosives
- Manufacturing synthetic fibers and flocculants
- Used as dehydrating agents

Dehydrating agent.

Concentrated sulphuric acid is a powerful dehydrating agent. This means that it will take off water from different substances like sugar (Sucrose) and hydrated salt crystals (hydrated copper II sulphate crystals).

When concentrated sulphuric acid takes off water from sucrose it leaves only Carbon substance. This substance is black in colour. The temperature of the medium (container) of the reaction is raised. This means that the reaction is exothermic.

The dehydrating reaction,

Sucrose + $_{\text{conc. H2S04}}\rightarrow$**Carbon** + water.

$C_{12}H_{22}O_{11(s)} \rightarrow 12C_{(s)} + 11H_2O_{(l)}$

The figure 1.22. The concentrated Sulphuric acid has removed the elements of water from the sugar, leaving black carbon.

When few drops of concentrated sulphuric acid are added to hydrate copper II sulphate crystals, it leaves only anhydrous copper II sulphate crystals. Water is the by-product of the reaction.

Sulphates.

These are called salts of sulphuric acids. They have a general molecular formula of **XSO_4** where **X** is a metal or non- metal. To test for a sulphate, add a few drops of dilute hydrochloric acid to unknown compound. This is followed by addition a few drops of barium sulphate ($BaSO_4$). If sulphate is present, a white precipitates of barium sulphate is produced.

Many sulphate have very important uses in laboratories, chemical engineering and pharmaceuticals.

NB: Sulphate are formed by reaction of acid with carbonate, bases and some metals.

The table:1.22 uses of metal sulphates.

Salt	**Formula**	**Use**
Ammonium sulphates	$(NH_4)_2\ SO_4$	Fertilizer
Barium Sulphate	$BaSO_4$	Use in diagnostic X-ray standers.
Calcium Sulphate	$CaSO_4\underline{1}\ H_2O$ 2	Plaster of Paris used to set bones.
Magnesium Sulphate	$MgSO_4$	In medicine used as a laxative

Self-assessment 1.1.

1. (a) write down two sources of Sulphur

   (b) Mention three properties of Sulphur

   (c) State five uses of Sulphur

2. Study the equation below and answer the given questions

   $C_{12}H_{22}O_{11}(s)$ conce 12c + B

   H_2SO_4

i. Complete the equation
ii. Name the process shown above
iii. Mention three uses of sulphuric acid.

3. (a) Define "sulphate"
   (b) Write down the use of each of the following sulphate:
   - $(NH_3)_2SO_4$
   - $BaSO_4$
   - $CaSO_4$
   - $MgSO_4$

4. Define "contact process"

1.23 Phosphorus.

The non- metallic element, which is found in Group V and period 2 of the periodic table. This element also plays a major role in Agriculture areas. In pharmaceutical, medicine is also produced using the same element.

1.24 Sources of phosphorus.

- Crab shells
- In coal
- In banana peels

1.25 Properties of phosphorus.

This too has both physical and chemical properties.

Physical Properties of Phosphorus.

- It has low melting and boiling points as compared to metal elements.
- It exists in two form (allotropes)
- It is insoluble in water because it is non- polar element
- It is soluble in organic solvent

Chemical Properties of Phosphorus.

(a).Reaction with oxygen.

Phosphorus reacts with oxygen to form phosphorus pentoxide. This means that it burns very readily in oxygen to give bright flame for instance;

$$P_{4(s)} + 5O_{2\,(g)} \longrightarrow 2P_2O_5$$

Phosphorus + oxygen $\longrightarrow$ phosphorus pentoxide

(b).Reaction with chlorine gas.

Phosphorus burns in chloride gas to form phosphorus trichloride (PCl_2) and phosphorus pentachloride (PCl_5) when excess chloride supplied.

$$P_{4(s)} + 10\ Cl_{2\,(g)} \longrightarrow 4PCl_5$$

1.26 Uses of phosphorus.

- Manufacturing match heads
- Manufacturing detergents e.g surf
- Manufacturing fertilizers
- Manufacturing rat poison
- Used in bombs to release energy

1.27 The compound of phosphorus.

The phosphorus has several compounds. Of all the compounds, we are going to focus on phosphates and phosphoric acid in terms of their uses. Phosphates are the salt of phosphoric acid. These two compounds are important in most of the field in chemistry studies.

1.28 Uses of phosphoric acids and phosphates.

- Phosphates are used in baking's.
- Phosphate are used in toothpaste.
- Phosphate are used as a source for bone development in the body.
- Phosphate acids is used as a raw material for manufacturing fertilizer.

Self-assessment 1.3.

1. (a) Give two sources of phosphorus

(b)What are the three properties of phosphorus?

2. Mention three uses of phosphorus
3. (a) Define "phosphate"

(b) Write down one use of each of the following compounds

i. Phosphate
ii. Phosphoric acid

4. Why phosphorus is insoluble in water?

1.29 Summary

necessary for the well- being of plants and animals. It makes up 78% of air. This element is commonly sound is air and soil. Nitrogen is colorless, odorless and unreactive at room temperature. The element is also very important in the production of fertilizers, Ammonia and nitric acids.

- Ammonia and nitric acids are the compounds of nitrogen. Ammonia is formed industrially by Haber process while nitric acid is formed by Ostwald process. These two compounds in industry are used in manufacturing dye, plastics, fertilizer and explosives.
- Sulphur is also another non- metallic element found naturally in crude oil, volcanic region and metal over. It is insoluble in water and yellow- brittle solid at room temperature. Sulphur is used in manufacturing matches, sulphuric acid and volcanisation of rubber.
- Sulphuric acid is a compound of Sulphur formed by contact process. This is used in industry for manufacturing of different things like fertilizer, detergent, soap and drug. The sulphate are also there to aid in diagnostic of disease.
- Phosphorus is another non-metallic element found in rock and crab shells. It is found in the allotropes. This is used to manufacture rat poison, bomb and match head. Phosphate and phosphoric acid are also used in fertilizer production.

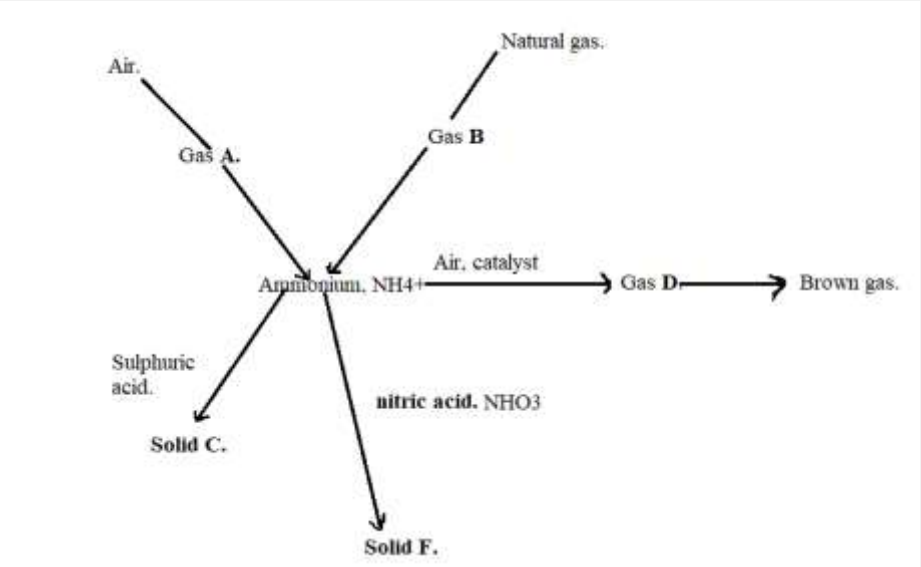
1.30 Revision Exercise

1.a). State two sources of nitrogen.

b). Explain three reason why nitrogen element is very important.

c). Mention three properties of the nitrogen.

2. Study the following reaction scheme.



a) Identify the substance A to F by giving their names and chemical formula.
b) How is gas A obtained from the air?
c) Write the word and balanced chemical equation for the formation of ammonia gas from gases A and B.
d) Write an equation for the formation of solid C from ammonia.
e) Give a use for
   Solid C.
   Nitric acid.
   Solid F.

3. a) State two uses of Sulphuric acid.

b. Explain how Sulphuric acid is formed?

c. Give one use of each of the following sulphates.

Magnesium sulphates.

Calcium sulphates.

d. State four uses of Sulphur in everyday life.

c. Describe three properties of Sulphur.

4. a).Explain why phosphorus is very important.

UNIT 2

Chemical bonding II

Learning Objectives.

By the end of this unit, students should be able to:

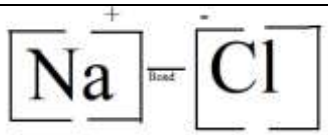
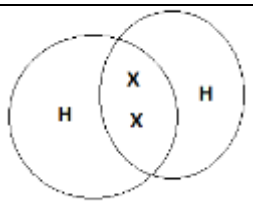
a) Describe properties of ionic and covalent compounds.
b) Explain the structural different between ionic and covalent compounds.
c) Different between poral and no- poral bond.
d) Relate intermolecular force to properties of ionic and covalent compounds.
e) Explain the physical similarities between diamond and silicon oxide.
f) Describe the use of metals in relation to their properties.
g) Explain the physical properties of alloys.
h) State the use of alloys.

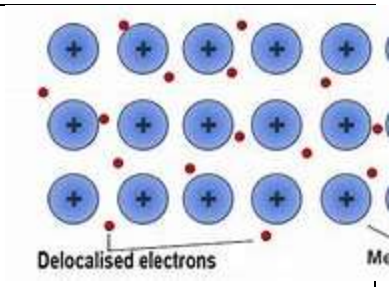
2.10 Introduction.

You have already learnt some aspects of chemical bonding in JCE chemistry. In this book, we will build upon this knowledge to discuss the properties of each type of chemical bond in this topic and other related topics in some chapters.

In JCE chemistry, chemical bond was defined as a force of attraction between two particles. Therefore, we further discussed three types of chemical bond as shown in table **2.10.**

Table 2.10. Type of bond, description and its diagram.

Types of chemical bond.	**Description.**	**Diagram.**
Ionic Bond.	• Involves between metal and non-metal atoms. • The metal atom loses its electrons in the outermost energy level to the outermost shell of non- metal atom. • In this case, metal atom becomes positively charged (oxidized) and non-metal atom becomes negatively charged (reduced). • The strong force of attraction created between these two opposite charged.	
Covalent	• The bond is between non-metal atoms. • The non- metal atoms allow their outer energy to overlap. A molecule of non- metal is formed, with two of them sharing a pair of electron into the overlapped region. • This shared pair of electrons forms a **single covalent bond**	

Metallic.	• This bond occurs among metals. • When metals are closely packed together, each loses its outer electron(s). This forms the sea of electrons (delocalized-mobile electrons). • The metal atoms now become positively charged ions. Therefore, the ions are embedded in the sea of electrons. • There is attraction between the ions and the electrons. This kind of attraction between them is called **metallic bond**.	

In this unit, you are going to study the properties, allotropes and the uses of metals and its alloys.

Self-Assessment 2.1

1. a).Define a chemical bond.

   b).Name three types of the bond.

2. Explain how ionic bond is formed?

2.11 Properties of ionic compounds/ substances.

Ionic compounds have the following properties:

- They are usually solid at room temperature
- They have high melting and boiling points. The strong force of attraction between oppositely charges ionic requires more heat energy to be broken down.
- Ionic compound dissolves in water. This is because water molecules are able to bond positive and negative charges, which break up the lattice and keep the ions apart.
- They usually conduct electricity when in aqueous solution. The force of attraction between the ions is weakened and the ions are free to move. This allows the electrons solution.
- Ionic compounds are insoluble in organic solvent like benzene or prop none

2.12 properties of covalent compounds.

Covalent substances/ compounds have the following properties*:*

- They have low melting and boiling point as compared with ionic compounds. This is because that, the force of attraction that hold molecules together are very weak, hence easily broken down with less heat energy.

- They do not conduct electricity in molten or solution form. They have no free ions that can allow current to pass through since they are not charge.
- They are insoluble in water due to lack of charges that can bonds with water.
- They exist in different states of matter at room temperature. Some are gases (hydrogen H_2), some are liquids (ethanol) and others are solids (sugar).
- Covalent compounds are soluble in organic solvent like benzene and propanone.

The table 2.11 summarises the properties of ionic and covalent substances.

Ionic substances.	**Covalent substances.**
High melting and boiling point	Low melting and boiling point
Soluble in water	Insoluble in water
Solid at room temperature	Liquid or gas at room temperature
Insoluble in organic solvent e.g benzene	Soluble in organic solvent e.g benzene
Conduct electricity in solution form	Do not conduct electricity in solution form

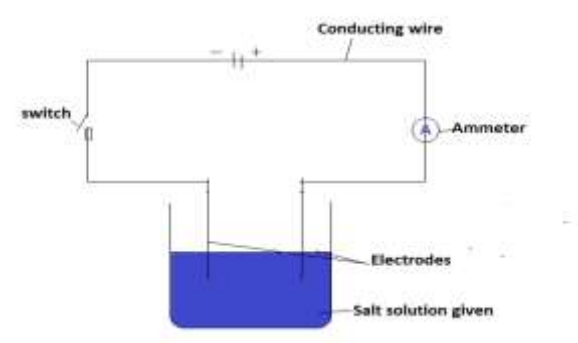
Experiment 2.0

Aim: To classify sugar, salt and cooking oil as an ionic or covalent Compounds

Materials:

- Sugar, salt, cooking oil,
- ammeter, measuring cylinder,
- three beakers,
- Circuit (cells, wires, and switch)

Methods/ procedures.



(a) Set up the apparatus as shown in above (figure
(b) Pat 50cm^3 of salt solution in the beaker
(c) Close the switch
(d) Record the observation in the table of results
(e) Repeat the procedures b-d with 50cm^3 of sugar solution and cooking oil.

Table 2.12. Results recorded from the experiment.

The substance (s)	Effect on Ammeter readings
Salt	Reading is observed (counts)
Sugar	No readings
Cooking oil	No readings

Discussion.

When switch is closed in the presence of sugar solution, Ammeter reading starts taking counts. This is because salt dissociate completely in water to form positive negative ions that can allow current to pass through. This is possible because water is able to bond salt. Ammeter is used since it takes smallest reading unlike bulb. Cooking oil and sugar ammeter do not take any reading because current is not passing through them since they have no ions.

In circuit or conducting wires current is passing through e electrons and in solution through ions.

Conclusion.

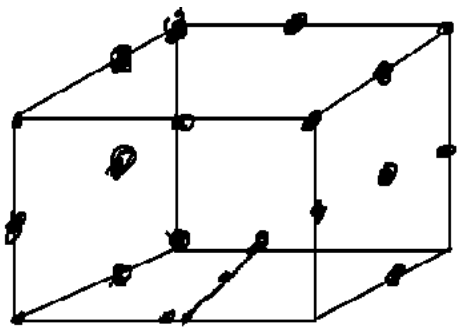
Salt is ionic while sugar and cooking oil are covalent substances.

2.13 The structure of ionic and covalent substances.

Ionic Compounds

In this type of compound, the particles are closely packed together in regular arrangements. This order forms a crystalline structure called **lattice** (*regular three dimensional arrangement of atoms in solid*).

The particles are held together with strong forces of attraction between positive and negative charges. The charged ions are found in a specific arrangement that requires more heat energy to break the force between atoms hence melting and boiling point increases.



The figure 2.13a. Arrngement of particles in ionic compounds.

Covalent compounds

In covalent structure, particles are loosely held together with weak intermolecular forces. This forms irregular shaped structure like that of the *diamond, silicon oxide and graphite*. The forces that hold the molecules together in these compounds are molecular covalent bonds. This force is easily broken down with less amount of energy; hence constitute low melting and boiling points.

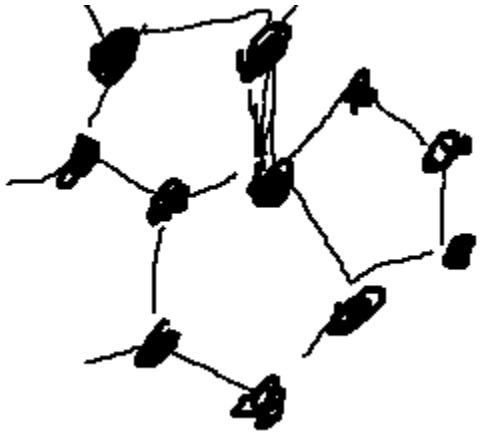


Figure 2.13b. The giant covalent structure.

Table 2.13. **Structure different between ionic and covalent compounds.**

Ionic compound.	**Covalent compound.**
Giant structure called lattice	Simple and molecular giant structure
Are bound by strong ionic force	Held by break intermolecular forces
High melting and boiling point	Low melting and boiling point
irregular crystalline structure	irregular structure

Melting and boiling points of ionic and covalent substances.

As the number of molecules forming a compound become very large, giant structures are formed. Consequently, the forces between them increases of which will require more heat energy to be broken down. This increases melting and boiling points.

Self-Assessment 2.2

1. Identify the type of bond in each of the following compounds
   i. NaCl
   ii. Cu
   iii. HNO_3

2. (a) Explain why covalent substances do not conduct electricity in its liquid state?

   (b) State three properties of ionic substances.

(c) Explain why ionic compounds have high melting and boiling point?

1. (a) Explain how metallic bond is formed?
   (b) Mention three properties of covalent substances

i. Draw the arrangement of particles in sodium chloride (NaCl).

ii. Explain why sodium chloride crystals dissolve in water?

3. Describe three structure difference between ionic and covalent substances.

2.14 Polar bond.

This is the type of covalent bond. The word polar means charged. Therefore, the bond has ionic characteristics, though it is the covalent type.

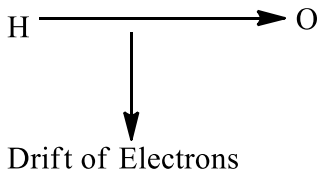
What Happens in This Type of Bond?

Two types of atoms are involved, electronegative and non- electronegative atom. Non-electronegative atom is the atom with less affinity for electron attraction, hence it partially loses the electrons. In this case, electrons are not equally shared. The electros shared are closely attracted to the more electronegative atom.

The electronegative atom now becomes surrounded with more negative charges hence it develops partial negative charge symbolized as^{-} ð .The non- electronegative atom is deprived of electrons and is surrounded with few electrons. This result into development of partial positive charges. The symbol for partial positive charge is $^{+}$ ð.

The permanent dipole is created by the drift of electron cloud in the direction of electronegative atom. This result into the bond called **polar.**

Let's consider water (H_2O). Hydrogen (H) shares its electrons with oxygen (O) to form covalent bond. The shared electrons are closed to the side of more electronegative charges. (O^{-} ð). The hydrogen atom is deprived of its electrons and it develops partial positive charge (H^{+} ð.) as shown below.



The permanent dipole is created by drift of electrons to oxygen atom. This result into the bond called **polar (charged).**

However, the bond has ionic characteristics. The greater difference in electronegative of the atoms results into the greater ionic character of the bond. The bond compound that has a polar bond is said to be ***polar covalent substance*** and is easily dissolved in water to conduct electricity. Therefore, polar compound conduct electricity. Examples of some polar compounds are ***Water*** (H_2O) and ***Hydrochloric acid*** (HCl).

Dative bond.

The dative bond is formed between positive ions and a compound having a lone pair of electrons. Lone pair of electrons has non- bonding electrons. These electrons do not take part in chemical reactions. The dative bond is also called **co- ordinate bond**. The bond formed between **hydrogen ions H^+** and **ammonia NH_3 to form ammonium NH^+_4** is a good example of this type of bond.

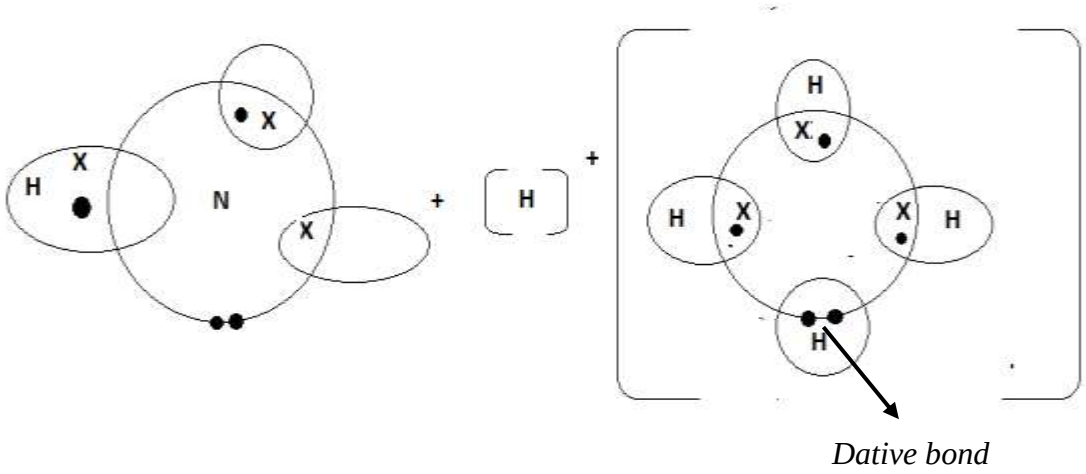


Figure 2.14. Formation of dative bond.

The one pair of electrons of nitrogen atom that is not bonded with hydrogen atom in ammonia is called **lone pair of electrons**. In dative bond, the bond exists between this lone pair of electrons and ions.

Dative bond also occurs in the formation of hydronium H_3O^+since water has oxygen atom which is only bonded with two hydrogens. Therefore, oxygen has two pair of lone electrons.

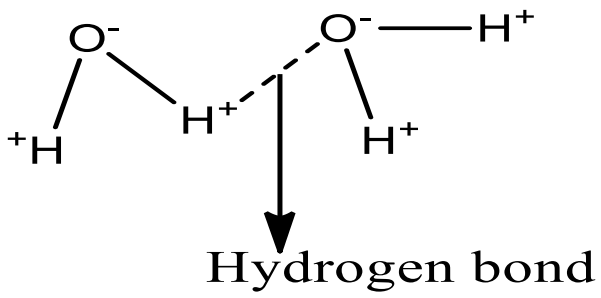
Hydrogen bonding.

Water substance has multiple dipole bonds. This creates permanent dipole- dipole interactions between adjacent molecules. This means that the positive charge in one molecule attract negative charge from another molecule. This constitute **hydrogen bonding.**

What is the hydrogen bond?

This is the strong force of permanent dipole- dipole interactions that is formed when hydrogen atom (H) is bonded to a strong electronegative atom like nitrogen (N), oxygen (O) and chlorine (Cl) atom.

Hydrogen bond may also occur between lone pair of electrons of one molecule and partially positively charged hydrogen atom of one molecule. Lones pair is **potential site**, in which positive charges are attracted.



The figure 2.14b. The hydrogen bond.

Non- polar bond.

Electrons in this bond are equally shared. This means that these two atoms have the same electronegativity. This is also a pure example of covalent bond. The non-polar bond occurs in substances like alkane, chlorine, and Sulphur chloride and hydrogen molecule.

Non- polar substances have weak force; therefore, they are easily broken down by less amount of heat energy. This constitutes low melting and boiling point.

The table 2.14. The differences between polar and non-polar substances.

Polar substance	**Non- polar substance**
Soluble in water	Insoluble in water
High surface tensing	Low surface tensing
High melting and boiling point	Low melting and boiling point

2.15. Intermolecular forces.

These are forces of attraction between molecules. The size of intermolecular force increases as the size and charge of an atom increases. In the JCE chemistry, you learnt that moving across the periods of the periodic table, the charges increases. This result into strong intermolecular forces.

The polar substances have strong intermolecular forces because they are charged unlike non- polar substances which are not charged.

Melting and boiling point of polar and non-polar substances.

As the charge of polar substances increases, the intermolecular forces also become stronger. Therefore, more heat energy is needed to break these forces. This increases melting and boiling point of the polar substances.

Polar substances also have hydrogen bond. This is the strong covalent bonds which require more heat energy to be broken hence it increases melting and boiling point.

Within non- polar substances, intermolecular forces increase as the size of molecules increase. **Why?** More energy will be needed to hold molecules together. This also constitute high melting and boiling point because more heat energy will be needed to break the force between molecules of non-polar substances.

Vander Waals forces.

These are weak forces of attraction which hold molecules together. This force is created by temporally induced- dipole force which is produced as a result of movement of electrons within the molecule of an atom. The Van der Waals forces are easily broken down by less amount of heat energy since they are weak forces. This constitutes low melting and boiling points.

In your further studies of Chemistry, you will learn that Hydrogen bond and Vander Waal forces are very important weak force in biochemical reaction of the body. This will be studied in Biochemistry. (*Beyond this scope*).

The effects of size and kind of atoms in strength to the van der Waals forces.

The size of molecule determines the strength of the van der Waals forces. The size of van der Waals forces (strength) increases as the size of molecules increase. For example, along an organic family series, the molecular get bigger and the intermolecular forces get stronger. Larger molecular have more electrons which induce stronger intermolecular forces.

The kind of atoms in molecules also determine the strength of van der Waals forces. This is because the kind of an atoms in the molecules will determine the size of the charge in the whole molecule. For instances, the amount of charge in Nitrogen (N_2) is less than in iodine(I_2). Hence, there will be stronger van der Waals forces in iodine than in nitrogen.

Self-Assessment. 2.3

1. (a) Identify the type of covalent bond in each of the following molecules.
   i. Water (H_2O)
   ii. Carbon dioxide (CO_2)
   iii. Hydronium (H_3O+)

   (b) Explain why water has high melting and boiling point than carbon dioxide.

2. (a) Define “dative bond”

   (b) Draw the dative bond formed between hydrogen ion and Ammonia

   (c) Mention three differences between polar substance and non- polar substances

3. (a) Define “van der Waals force”

(b) Explain giving examples, how the strength of Van der Waals forces depend on Molecular size and the kind of atoms in the molecule.

4. (a) Define "hydrogen bonding"
   (b) Explain the intermolecular force in relation to the polar and non- polar substances.
   (c).Define lone pair of electron.
   (d). Draw the hydrogen bond in between water molecule and Hydrochloric acid.

2.16 Allotropy.

Allotropy by definition is refers to the existence substance of an element in two or more different form in the same physical state. Therefore, **Allotropes** are the different forms of an element.They exhibit different physical properties.However, identical chemical compounds can be formed from the various allotropes of the same element.

The name of allotropy is **polymorphism.** Allotropes can be used differently. **Why?** *They have different physical properties*. Examples of elements, which show allotropes are; ***Sulphur, tin, iron and carbon.*** Based on your syllbus, you are encouraged to concentrate on the allotropes of **carbon** and **Sulphur**.

Allotropes of Sulphur.

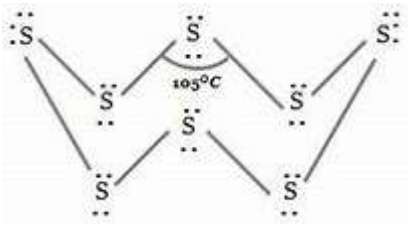
These are different forms of Sulphur.The element of Sulphur has four allotropes which are made up of the simple S_8 rings. These are:

- Rhombic sulphur
- Monoclinic sulphur
- Plastic sulphur
- Amorphous Sulphur

However, in this book, you are going to concentrate on **Rhombic** and **Monoclinic sulphur.**

(a).Rhombic sulphur.

This has interlocking arrangement of s_8 rings, with each other. It consists of large yellow crystals with an octahedral shape. This allotrope crystallizes from a solution of carbon disulphide at room temperature. Therefore, they are stable below 96^{o}c. Rhombic sulphur is also called **alpha (a) sulphur**.



The figure 2.16a. Interlocking patterns of s_8 molecules in rhombic Sulphur.

(b).Monoclinic sulphur.

In this allotropes, the particles are stable above 96^0C. The s_8 rings of molecules are stacked. This kind of allotropes is also called ***beta sulphur*** (**B- sulphur**).

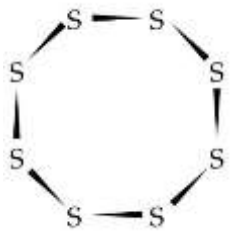


Figure 2.16b: Stacked pattern of arrangement of s8 molecule in monoclinic Sulphur.

The structure of Rhombic and monoclic Sulphur.

Of the two allotropes of sulphur, rhombic sulphur is more stable because it has interlocking arrangement of sulphur s_8 rings while in monoclinic s_8 ring are stacked arranged.

Melting and boiling point of allotropes of Sulphur.

As we have looked at the structure arrangement of both llotropes of Sulphur, we will use such knowledge to describe the melting and boiling point of the llotropes of Sulphur.

Therefore, in terms of melting and boiling point, Rhombic sulphur requires more heat energy to break strong intermolecular forces between them hence constitute high melting and boiling point unlike in monoclinic sulphur where by particles are easily dispersed with less amount of heat energy.

Self-Assessment 2.4

1. a).Define the term allotropy.
   b).State four allotropes of Sulphur.
   c).With the aid of diagrams, explain why rhombic sulphur is more than monoclinic Sulphur?

2. a).Why rhombic Sulphur has high melting and boiling points than monoclic Sulphur?

   b). Define polymorphisms.

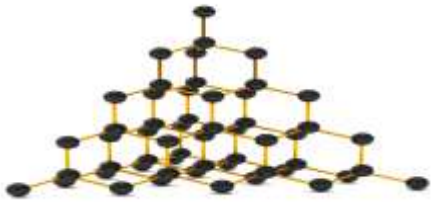
2.17 Allotropes of carbon.

The carbon atom has two allotropes;

- Diamond
- Graphite

(a).Diamond.

In this kind of llotropes, the carbon atom is covalently bonded to the other four carbon atoms. This forms giant structure, of clear, colourless crystal of tetrahedral arranged atoms. The allotrope has no delocalized electrons (mobile). As such, it does not conduct electricity.Diamond is extremely hard because the atoms are closely packed together.



The figure 2.17a: The structure of diamond.

Physical properties of diamond.

Following the arrangement of atoms in diamond, it exhibit the following properties;

- Has high melting and boiling point
- Does not conduct electricity
- Extremely hard substances at room temperature
- Has high densities unlike graphite
- It is the least compressible substances

Uses of the diamonds.

- Making cutting tools in industry and surgery.
- As heat absorbers
- For decoration purposes
- Making jeweler

(b).Graphite.

In graphite, the carbon atom is covalently bonded to the other three carbon atoms. This forms plane layer.This result into fused hexagonal ring arranged in an infinite three – dimensional structure. The layers joined by weak forces.

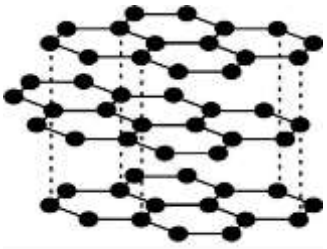


Figure 2.17b: The structure of graphite.

Physical properties of graphite.

- It conducts electricity.
- It is soft materials with slippery feel
- It has low density unlike diamond. Its density is approximately equal to $2.25glm^3$
- It has low melting and boiling point

Uses of graphite.

- Making lubricants
- Making pencils
- Making electrodes

Melting and boiling point of Diamond and graphite.

Diamond has tetrahedral shape in which carbon atoms are closely packed together with strong intermolecular covalent bond. This is very strog bond, as such, it requires more heat energy to be broken down.

In graphite, less force is required to break weak bond in fuses hexagonal rings hence particles are easily dispersed with less amount of heat energy.

2.18 Similarities between diamond and silicon oxide.

Diamond and silicon oxide are very similar based on the followings;

- Both have tetrahedral arrangement of atoms
- Both have high melting and boiling point approximately (1160^0C)

- Both are very hand substances
- Particles in both are very closely packed together
- Both substance do not conduct electricity
- Both are insoluble in water

Important note.

Silicon gets bonded tetrahedrally to four oxygen atoms, resulting into giant structure. Examples of silicon oxide which are common are **sand** or **silica**.

Self-Assessment 2.5

1.a) i . Mention two allotropes of carbon.

ii . Give three physical properties of graphite.

iii . Explain why diamond do not conduct electricity.

(b) Mention two uses of each of the followings:

i. Diamond_____

ii. Graphite_____

2. (a) Explain why diamond has high melting and boiling point than graphite

(b) Give two physical properties of graphite.

(c).Mention two similarities between diamond and silicon oxide

(c) State two examples of silicon oxide

3. With the aid of diagram, explain the differences between **diamond** and **graphite**

2.18 Properties of metals.

Metals have the following properties:

- They are malleable and ductile.
  - ❖ Malleable means that metals can be hammered into different shape.
  - ❖ Ductile means metals can be pulled into thin wires.
- Sonorous. Produces sound which is ringing.

- They are good conductors of heat and electricity. This is possible because they have mobile electrons within metal structures in which current is passed.
- Metals have high melting and boiling point. The strong force of attraction between positive metal ions and the mobile sea of electrons requires more heat energy to be broken down.
- Metals have high density. Particle within the structures are very closely packed in regular manner.
- They are insoluble in polar and non- polar solvents.

2.19 Uses of the metals.

Metals are used in industry;

- For making cooking pots
- For electrical wiring
- For making bell
- Making bridge structures
- Making jewelry and ornament e.g. silver and gold

2.20 Alloys.

This is the mixture of two or more metals. The examples of alloys which are commonly found are **Brass** and **steel.**

Brass is the mixture of zinc and copper while **Steel** is the alloy of iron and carbon, sometimes with other metal included. The alloys are formed by mixing the molten substances thoroughly.

Physical prperties of alloys.

- Malleable e.g. pewter.
- Does not rust e.g. stainless steel.
- Light and tough e.g. magnesium.
- Does not corrode
- They are harder e.g. Bronze and Brass.
- Tough and brittle e.g. Hard steel.
- Sonorous e.g. Bronze.

Uses of alloys

The following are the uses of alloys in industry;

- Making car bodies e.g. stainless steel
- For making plate e.g. pewter
- Construction of air craft e.g. magnesium

- Making coins e.g. cuprous- nickel
- Making cutting tools like razor blades and chisel
- Making musical instruments
- Making statues and ornaments
- Making bells

Self-Assessment.2.6

1. (a) I. Define "alloy"

   II. Explain how alloys are formed?

   (b). What are the three physical properties of alloys

   (C). Mention **three** uses of alloys in everyday life.

2. (a) Mention **three** physical properties of metals

   (b) Explain why metals are good conductors of heat and electricity?

3. (a) Explain why magnesium metal has high melting and boiling point

   (b) Metals are sonorous. Explain why this is important

4. Mention three uses of metals

2.21 Summary.

- There are three types of chemical bonding; ionic, covalent and metallic bonding. Ionic substances dissolve in water while covalent substance does not. In polar covalent bonds shared electrons are closely attracted to the more electronegative atom while in non-polar, electrons are shared equally.
- Ionic structure is lattice which is three dimensional arrangements of particles while in covalent, atoms form simple giant structure in which they are irregularly arranged.
- Polar covalent bonds have strong force of attraction because of charges. In non-polar, attraction forces increase as the size of molecules increase.
- Sulphur has four allotropes; Rhombic, monoclinic, plastic and amorphous while carbon has two allotropes diamond and graphite. Diamond is used for making jewelry while graphite is used in pencils. Diamond is similar to silicon oxide since both they possess tetrahedral arrangement of particles.
- Metals are ductile, malleable and shiny as such; they are used to make plates, coins and electrical wires. Alloy is mixture of metals. These are tough and do not corrode hence they are used in construction and making tools.

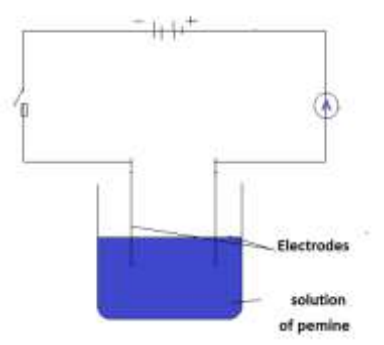
2.22 Revision Exercise.

1. Describe three properties of each of the following compounds.
   Ionic compounds.
   Covalent compounds.
2. a).Define polar bond.
   b).Describe three difference between polar and non-polar substances.
   c).Why water molecules has high melting and boilig point than butane substances?
3. a).Define allotropes.
   b).Name two allotropes of
   i).Sulphur
   ii).Carbon.
4. a).With the aid of diagram, describe the structure of
   i).Rhombic and monoclic Sulphur.
   ii).diamond and graphite.
   b).State three properties of
   i).Diamond.
   ii).Graphite.
5. Describe three differences between diamond and silicon oxide.
6. State three uses of graphite.
7. Explain. Why diamond do not conduct electrity?
8. Mention three properties of metals.
9. Why metals are very important? State three reasons.
10. a).Define Alloys.
    b).Describe three properties of alloys.
    c).Name three uses of alloys.

11. With the aid of well labelled diagram, describe an experiment to demonstrate that salt solution conducts electricity than sugar solution.

2.23 Practical Question.

1. You are provided with the following materials, the cells, conducting wires, 50cm^3 of cooking oil, sugar, sodium chloride (NaCl), petrol, diesel, magnesium hydroxide ($Mg(OH)_2$, ammonium chloride (NH_4Cl), pentane (C_5H_{12}), two beakers and measuring cylinder.
   a. Set up the material as shown below



(a) Measuring 50ml of sodium chloride (NaCl) using the measuring cylinder
(b) Put the measured solution into the beaker as shown in the diagram above
(c) Close the switch
(d) Observation the result and record in the table
(e) Repeat the procedures b- e with sugar, petrol, diesel, pentane, magnesium hydroxide and ammonium chloride
(f) Complete the table below by putting current/ no current ***6marks***

Substance (s)	**Ammeter reading(s)**
Sugar	
Cooking oil	
Sodium chloride (Nacl)	
Petrol	
Diesel	
Magnesium hydroxide($Mg(OH)_2$)	
Ammonium chloride (NH_4Cl)	
Pentane (C_5H_{12})	

- Classify the compounds as;
  a. **Ionic compounds**_____
  _____3 marks
  b. **Covalent compounds** _____
  _____5 marks

Explain why ammeter is used instead of a bulb. 1mark

Mention two sources of errors in this practical.

UNIT 3

Stoichiometry

Learning objectives.

By the end of this unit, students should be able to:

a) Write a balanced equation.
b) Work out the relative formula mass of a compound.
c) Define the mole of a substance.
d) Convert moles into other units of measurements.
e) Determine the percentage of water in molecular and hydrated ionic compounds.
f) Deduce empirical and molecular formulae from relevant data.
g) Calculate concentration of a solution using titration.
h) Prepare the standard solution.
i) Determine the concentration of a solution using titration.
j) Determine the yield in a chemical reaction.
k) Define the term exothermic and endothermic in relation to heat changes.
l) Describe temperature changes in exothermic and endothermic reaction.
m) Draw energy level diagram for exothermic and endothermic reaction.
n) Explain why bond breaking is endothermic and bond making is exothermic.
o) Determine whether a reaction is exothermic or endothermic using bond energy.

3.10 Introduction.

Stoichiometry is the branch of chemistry that is based on the *law of conservation of mass* which states that ***"the total mass of the reactants equals the total mass of the products"***, leading to the insight that the relations among quantities of reactants and products typically form a ratio of positive integers. This means that if the amounts of the separate reactants are known, then the amount of the product can be calculated. Conversely, if one reactant has a known quantity and the quantity of the products can be empirically determined, then the amount of the other reactants can also be calculated.

Stoichiometry measures these quantitative relationships, and is used to determine the amount of products and reactants that are produced or needed in a given reaction. Describing the quantitative relationships among substances as they participate in chemical reactions is known as ***reaction stoichiometry.***

This theory is widely used in drug companies, school and clinical laboratories and by chemical engineers to calculate quantities like Mass, Moles and Volume given the reactants and percentage yields. Stoichiometry also predicts how elements and components are diluted in standard solution.

In this unit, we are going to calculate different quantities and concentration of solutions based on the outlined objectives.

3.11 Chemical reaction and Chemical Equation.

In JCE chemistry, you learnt that the paper that is burnt is called **reactant** and the ash produced from it is generally called **product**. This is a kind of chemical changes. When burnt, the substances usually combine chemically with oxygen in the air.

The word **chemical reaction** means *the re-arrangement of atoms to form new substances*. To understand this definition, let us consider chemical reaction in the experiment below.

Experiment 3.1

Aim: To demostrate Reaction of Sodium chloride and lead nitrate.

Materials.

- 30ml of Sodium chloride.
- 30ml lead nitrate.
- Test tubes.
- 20ml Measuring cylinder.

Procedures.

a) Clean the test tube with water.
b) Put 10ml of Sodium chloride in the cleaned test tube.
c) Add 5ml of Lead nitrate in the same test tube.
d) Leave the set up for 15 minutes.
e) Observe the tube after this given time and record the results.

Observation.

White precipitates are observed in the tube.

Discussion.

When the sodium chloride (NaCl) and lead nitrate $Pb(NO_3)_2$ are mixed, they react to form lead chloride ($PbCl_2$) and sodium nitrate ($NaCO_3$) with the white precipitates. In this case, Sodium chloride and lead nitrate are **reactants** while lead chloride and sodium nitrate are **products** of the reaction. This is summarised as:

Sodium chloride + Lead nitrate $\longrightarrow$ Lead chloride + Sodium nitrate.

Using chemical symbol.

$Pb(NO_3)_2\,(aq) + 2NaCl\,(aq) \longrightarrow PbCl_{2s} + 2NaNO_3\,aq$

For the reaction to take place, the reactants dissociate and produce ions as follows;

$Pb^{+2}\ NO^{-3} + Na^{+}\ Cl^{-}$

Therefore, the oppositely charged ions are attracted in the solution as shown below,

$Pb^{+2}\ NO^{-}_{3} + Na^{+}\ Cl^{-}$

Then, they produce new substance as they are re- arranged to be stable. In this case, they exchange the valence.

$Pb^{+2}Cl + Na^{+} + NO_3^{-1}$

$PbCl_2 + NaNO_3$

Therefore, the whole equation is

$Pb\ (NO_3)_{2(aq)} + 2NaCl_{aq} \longrightarrow PbCl2_{\ (s)} + 2NaNO_{3\ aq}$

The product formed is not equal to the reactant. Therefore, they may have different **physical** and **chemical properties.**

Conclusion.

From the discussion, when the chemical react, new substances are produced.

Chemical quation.

In JCE chemistry, you have also looked that chemical equation can be presented using word, chemical symbol and diagram. Therefore, the chemical reaction of Sodium chloride plus lead nitrate to produce lead chloride and sodium nitrate is presented by both word and chemical symbol respectively:

Sodium sschloride + Lead nitrate $\longrightarrow$ Lead chloride + Sodium nitrate.

The same equation can be written using the chemical symbol as:

$Pb\ (NO_3)_2\ (aq) + 2NaCl\ (aq) \longrightarrow PbCl_{2s} + 2NaNO_3\ aq$

This presentation of chemical reaction is called chemical equation. Therefore, the word **chemical equation**, is defined as *the symbolic form of chemical reaction in which formulas are used to indicate reactants and products.*

Rules for writing Balanced chemical equation.

In Stoichiometry sciences, it is very important to work with well balanced chemical equation in order to get good results of what we are working. Therefore, chemical equation should be written based on the following rules:

- Make sure that reaction should take place based on experiential fact.
- Write the correct formula for reactants and products.

- Balance the number of atoms on both side of equation using coefficient (number written in front of any formulae).
- Write the physical state of each reactants and products; solid (s), liquid (l), gas (g) and aqueous (aq).

NB: Start with small number when balancing equation.

Example 1.

Methane reacts with Oxygen to produce Carbondioxide and water. Write the balanced equation.

Solution.

Methane + Oxygen $\longrightarrow$ Carbondioxide + water.

Using formulas.

$CH_{4(g)} + O_{2(g)} \longrightarrow CO_{2(g).} + H_2O_{(l)}$

Left hand side (Reactants)	Right hand side (Products).
No. of atoms	No of atoms.
C=1	C=1
H=4	H=2
O=2	O=3

The equation is not balanced.

Put 2 **(coefficient)** in front of **O** (g) **why2?** It is a small number.

$CH_{4(g)} + \mathbf{2O_{2(g)}} \quad CO_{2(g).} + H_2O_{(l)}$

Now, the number of C, O, and H atoms in the reactants are;

C = 1

H = 4

O = 4

Put 2 in front of **H_2O** to balance the number of oxygen atoms (O) in the products.

$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g).} + \mathbf{2H_2O_{(l)}}$

No of atoms (Reactants)	No. of atoms (Products).
C=1	C=1
H=4	H=4
O=4	O=4

The equation is now balanced.

Self-Assessment 3.11

1. (a) Define the following:
   i. Chemical reaction
   ii. Chemical equation

   (b) Write the balanced equation of sodium chloride (NaCl) plus water (H_2O).

2. Balance this equation.

   $Mg(NO_3)_{2\,(aq)} + Na_2SO_{4(aq)} \longrightarrow NaNO_{3(aq)} + MgSO_{4(aq)}$

3. (a) What does stoichiometry states

   (b) Write down the symbol that represents the following physical states

   Solid_____

   Liquid_____

   Gas_____

   Aqueous_____

4. Write down the chemical formula of sodium chloride

3.12 Relative formula mass.

The relative formula mass (R.F.M) is the sum of the relative Atomic Mass of all those elements shown in the formula of the substances. This is often refered to as the relative molecular mass (Mr). This is the mas of one mole of a substance. This is also called molar mass. Relative formula mass is abbreviated as R. F. M. This has Sl unit of **glmol**.

To work out RFM of a compound, the following condition must be achieved.

Write the correct formula using

- Symbol of elements and formula of radicals.
- Valencies of elements and radicals
- Sum up relative atomic mass (RAM).

How can we calculate relative formula mass R. F. M.? Of a given compound?

Remember in junior section, you learnt how you can write chemical formula. Therefore, we can use that basic knowledge to calculate **RFM** by adding up the Relative Atomic masses of each atom in a given formula.

Example 2.

Find the Relative formula mass (RFM) of the Ammonium nitrate (NH_4NO_3) given the Relative Atomic Mass (RAM) of (N= 14 H=1 O= 16)

Solution.

i. Ammonium Nitrate

Chemical formula. **NH_4NO_3**

ii. **2N + 4(H) + 3(O)**

The number written in the front of each element locate the number of atoms presents in the formula.

iii. *2 (14) + 4(1) + 3 (16)*
= 28 + 4 + 48
=80gmol
Now RFM is 80gmls

3.13 The mole.

Chemists often need to know how much of a substance has been formed or used up during a chemical reaction. This is a particulry important in the chemical industry, where the substances being reacted (reactants) and the substances being produced (products) are worth thousands of pounds.

To solve this problem chemists need a way of counting atoms, ions or molecules. Atoms, ions or molecules are very tiny particles and it is impossible to measure a dozen or even the hundred of them. Instead, chemists weigh out a very large number of particles. This is $6x10^{23}$ atoms, ions or molecules and is called **Avogadros constant** after the famous Italian scientist Amedeo Avogadro (1776-1856).

The word **Mole** *is the amount of substance, which contains $6.023x10^{23}$ elementary particles*. The elementary particles are atom, ions or molecules as stated. The mole is abbreviated as **mol**. This is the weighing unit of very smallest particles. A mole is just like any other unit for measurement we use in science.

Calculations of Moles.

We have already seen that we can compare the masses of all the other atoms with the mass of carbon atoms. This is the basis of the **relative atomic mass scale**. Chemists have found by experiments that if you take the relative atomic mass of an element in grams, it always contains $6x10^{23}$ or one mole of its atoms. This means that:

1 mole of substance = 6.023 x 10^{23} particles.

Mathematically;

Number of moles $= \frac{\text{mass of substances}}{\text{molar mass of substance (RFM)}}$

Example 3.

Work out number of moles in 64g of sodium chloride (NaCl). RAM; Na = 23 (Cl=35.5)

Solution.

Molar mass (RFM) of NaCl =1xNa+1xCl

= 1x23 + 1 x 35.5

= 23+ 35.5

= 58.5 glmol

Number of moles = *Mass of a substance (g)*

Molar mass (RFM) glmol

$= \frac{64g}{58.5g\backslash mol}$

= 0.1 mol

64g of NaCl contains 0.1 mole.

Example 4

Calculate the mass of 0.1 mole of sodium carbonate ($Na_2 CO_3$)

(RAM= Na= 23 C =12 O=16)

Solution.

Molar mass (RFM) of Na_2CO_3 = 2xNa +C + 3xO

= 2x23+12+3x16

= 46+ 12+ 48

= 106gmol

Number of moles = 0.1 mole

But:

Moles = Mass of Na_2CO_3

RFM (molar mass)

0.1 = Mass of Na_2co_3

106glmol.

By cross multiplication:

Mass of Na_2CO_3 = 0.1x 106glmol

= 10.6g

0.1 Mole of Na_2CO_3 has a mass of 10.6g.

Example 5

Calculate the number of atoms in 3 moles of calcium.

Solution

1 mole = 6.0 $x10^{23}$ atoms
3 moles = x

By cross multiplication

$$\frac{X\ Moles}{1 mole} = \frac{3\ \text{moles x}\ 6.0\text{x}1023\ \text{atoms}}{1\ mole}$$

X = 3x $6.0x10^{23}$ Atoms

= 1.8 x 10^{24} atoms

3 Moles contain 1.8 $x10^{24}$ atoms.

Example 6.

Calculate the number of atoms in 80g of sodium hydroxide (NaOH).

(RAM: Na = 23 O = 16 H = 1)

Solution.

Molar mass (RFM) of NaOH

= Na +O + H

= 23 + 16 + 1

= 40glmol

$$\text{No of moles} = \frac{\text{Mass of NaoH}}{\text{Molar mass}}$$

Mass of NaoH= 80g

Then, No of moles = $\frac{80glmol}{40glmol}$

= 2 moles.

But 1 mole = 6.10^{23} atoms

2 mole = more

By cross multiplication.

X x 1mole = 2moles x 6.0 x10^{23}

X = 12x 10^{23} atom

= 1.2 x10^{24} atom

80g contains 1.2x10^{24} atoms.

Self-Assessment 3.12

1.a. Define the following terms.

i).Relative atomic mass

ii). Moles.

b. Work out the number of moles in 80g of magnesium chloride (MgCl2).

RAM: Mg=24, Cl=35.5

2. Calculate the number of atoms in 20g of Calcium hydroxide (Ca(OH)2).

RAM: Ca=40, O=16, H=1.

Mole and gases.

Many substances exist as gases. If we want to find the number of moles of a gas we can do this by measuring the volume rather than the mass. Chemists have shown by experiments that one mole of any gas occupies a volume of approximately 24dm3 at room temperature (25^{0}C or 298K) and one atmospheric pressure. At standard temperature (0^{0}C or 273K) and one atmospheric pressure, one mole of gas occupies 22.4dm^3. In general;

1 mole =24dm3, at room temperature (rtp).

1 mole=22.4dm3, at standard temperature (std).

Therefore, it is relatively easy to convert volumes of gases into moles of gases into volumes using above relationships. *The volume occupied by one mole of gas at standard temperature, atmospheric pressure and room temperature* is called **Molar gas volume**. The volume occupied by one of any gas must contains 6 x 10^{23} molecules. Therefore, it follows that equal volume of all gases measured at the temperature and pressure must contain the same number of molecules. This idea was first put forward by Amedeo Avogadro and is called **Avogadro's Law**.

Mathematically:

$$\textbf{Number of moles of gas} = \frac{\boldsymbol{Volume\ of\ the\ gas\ (in\ dm3\ at\ rtp)}}{\boldsymbol{24dm3}}$$

Number of moles of gas= $\frac{\boldsymbol{Volume\ of\ the\ gas(in\ dm3\ at\ std\ temp)}}{\boldsymbol{22.4dm3}}$

Example 7.

Calculate the volume occupied by 12g of oxygen (O_2) gas at 25 ^oC.

Solution.

Molar mass of O2 = 2x16

= 32glmol

No. of moles = $\frac{Mass\ of\ Oxygen}{Molar\ mass\ of\ compound}$

Mass of oxygen gas = 12g

No. of moles = $\frac{12g}{32g/mol}$

=1mol

= 0.375 mol.

At 25 ^oC, 1 mole = 22 .4 $dm^{3.}$

Then;

1 mole = 22.4dm^3

0.375 mole = **x**

$\frac{X\ mole}{1\ mole} = \frac{0.375\ mole\ X\ 22.4dm3}{1\ mole}$

X= 0.375 x 22.4 dm^3

= 8.4dm^3

12g of oxygen gas occupies 8.4dm^3 at 25^{0}C.

Example 8.

How many moles of hydrogen gas (H_2) are in 24cm^3 at room temperature?

Solution.

1mole = 24dm^3

Change volume from cm^3 to dm^3 by diving with 1000

Volume = $\frac{24cm3}{1000cm3}$ X 1dm^3

= 0.024 dm^3

By simple proportion

1 mole = 24dm^3

X mole = 0,024 dm^3

$$\frac{24x}{24dm3} = \frac{0.024dm3\ x\ 1\ mole}{24dm3}$$

X = 0.001 mole

24cm^3 occupies 0.001mol of hydrogen gas.

Another method.

No. of moles= $\frac{Volume\ of\ hydrogen\ gas}{24dm3}$

Vomule=$\frac{24}{1000}$x 100

0.024dm^3

No. of moles= $\frac{0.024}{24}$

0.001 moles of H_2

Example 9

Work out the mass of 100dm^3 of hydrogen gas at standard temperature.

Solution.

At 0^0C, 1 mole = 22.4dm^3

Then

1 mole= 22.4 dm^3

X mole = 100dm^3 (By simple proportion)

$$\frac{22.4dm3\ x\ 1mole}{22.4dm3} = \frac{1\ mole\ X\ 100dm3}{22.4dm3}$$

X = 4.464 moles

Molar mass of H_2 = 1 x 2

= 2glmol

Then

Mass of H_2 in 100dm3 = $\frac{2g\ X\ 4.464mole}{1\ mole}$

= 8.928g

Therefore, 100dm^3 occupies 8.928g at H_2 at 0 ^{0}C.

Example 10.

Calculate relative molecular mass of a gas 100cm^3 of the gas volume which has a mass of 0.067g. (At 25^{o}C, 1 mole occupies 24dm^3.)

Solution

Volume = $\frac{100cm3\ x\ 1dm3}{1000cm3}$

= 0.1 dm^3

No. of moles:

If $24dm^3$ occupies 1 mole

$0.1dm^3$ occupies? **Less**

By simple proportion

$$X = \frac{0.1dm3\ X\ 1\ mole}{24dm3}$$

$$= 0.004mol$$

But

$$RAM\ (molar\ mass) = \frac{mass\ of\ a\ gas}{NO.\ of\ moles}$$

$$= \frac{0.067g}{0.004mole}$$

$$= 16.75g\backslash mole$$

The relative molecular mass of gas is 16. 75glmol.

Self-Assessment 3.13

1. Define Molar gas Volume.
2. Work out the mass of $100dm^3$ of nitrogen gas at standard temperature.
3. Calculate relative molecular mass of a gas $50cm^3$ of the gas volume which has a mass of 0.9g. (At 25°C, 1 mole occupies $24dm^3$.)
4. How many moles of Oxygen gas (H_2) are in $24cm^3$ at room temperature?
5. Calculate the volume occupied by 12g of oxygen (O_2) gas at 25 °C.

3.14 Percentages of water in hydrated ionic and molecular compounds.

Water is incorporated into the structure of substances as they crystallize. For example, copper ll sulphate pentahydrate Cu SO. $5H_20$. This is called **water of crystallization.** Sometimes it is necessary to work out the percentages, by mass of water crystallization in a hydrated salt. So it is simply called **the percentage of water**.

Mathematically.

$$\textbf{The percentage of water} = \frac{\textbf{Mass of water}}{\textbf{Total mass of compounds}} \text{ X } 100\%$$

When the crystals of sugar compounds contain water molecular, the calculation will be called **percentage of water in sugar.**

Example 11

Calculate the percentage by mass of water in the salt hydrate $MgSO_4.7H_2o$. (RAM: H= 1, O=16, Mg=24, S=32)

Solution

- Mass of $MgSo_4.7H_2O$
  $=Mg + 11\ (O) + S+14(H)$
  $= 24+ 11(16) +1(32)+ 14(1)$
  $= 24.6g|mole$

Mass of $7H_2O$

$= 14(1) + 7(16)$

$= 14 + 112$

$= 126g\backslash mole$

The percentage of water $= \frac{\text{mass of } 7H2O}{\text{Mass of } Mg.7H2O}$

$= \frac{126g\backslash mole}{245g/mole} X\ 100\%$

$=51.2\%$

Therefore, the percentage of water in salt hydrated $MgSO_4.7H_2O$ is 51.2%

Self-Assessment 3.14

Calculate the percentage by mass of water in the salt hydrate $CaSO_4.5H_2O$. (RAM: C= 64, S=32, O=16, H=1)

3.15 Empirical and molecular formula.

If we have 1 mole of a compound, then the formula shows the number of moles each element in that compound. For example, the formula for lead (II) bromide is PbBr2. This means that 1 mole of lead II contains 1 mole of lead ions and 2 moles of bromide ions.

If we do not know the formula of a compound, we can find the masses of the elements present experimentally and these masses can be used to work out the formula of that compound.

The word **empirical formula** *means the formula showing the simplest ratio of atoms present while* **molecular formula** *is the formula showing the actual number of atoms of each element present in one molecule*.

Empirical formula only tells us the type of atoms present in a compound, but it does not tell us the exact number of atoms available in the compound. Therefore, only the molecular formula tells exact number of atoms present in the compound.

How do we determine the empirical formula?

- Find the mass of each element in the given sample. For the composition that is given as percentages assume that **100% is the same as 100g.** We take percentages as equal to the mass of the element
- If one mass/ percentage is missing, find it by subtracting from the total mass or percentage of the compound
- Convert the mass of each element into **moles.**
- Find the simplest whole number of the moles of each element by diving all with smallest number of moles.
- Then, you get **empirical formula**

Let us consider the different examples below:

Example 12

Calculate the empirical formula of an organic compound contains 48g of carbon, 12g of hydrogen and 32g of oxygen. (RAM: (C = 12, H = 1, O = 16).

Solution.

Using the table

Element	**Number of moles**	**Simplest ratio of number of moles**
C	$\frac{48g}{12glmole} = 4mole$	$\frac{4mol}{2mol} = 2$
H	$\frac{12g}{1glmole} = 12moles$	$\frac{12mol}{2mol} = 6$
O	$\frac{32g}{10glmole} = 2mole$	$\frac{2mole}{2mole} = 1$

The smallest number of mole is 2 moles. then divide by it throughout the table to get the ratio.

The empirical formula is C_2H_6O

Example 13

Find the empirical formula of the compound that has the following percentage composition by mass of element C=40%, H = 6.67% and O = 53.33%.

(RAM: C = 12, H= 1 O =16)

Solution

By using the table.

Elements	Number of moles	Simplest ratio
C	$\frac{40g}{12glmo} = 3.33mol$	$\frac{3.33mo}{3.35m} = 1$
H	$\frac{6.67}{1} = 6.67mol$	$\frac{6.67}{3.33} = 2$
O	$\frac{53.33g}{16glmo} = 3.33mol$	$\frac{3.33}{3.33} = 1$

Divide by 3.33 mole throughout to get the ratio

The empirical formula is **CH_2O**.

Example 14.

A certain compound contains 74.19% sodium (Na) and the rest is oxygen (O). If the relative molecular mass (RMM) is 124. Work out its molecular formula.

Solution

Using the table:

Elements	No. of moles	Simplest ratio
Na	$\frac{74lg}{23glmole} = 3.23mole$	$\frac{3.23mole}{1.61mole} = 2$
O	$\frac{25.61}{16mole} = 1.16mole$	$\frac{1.161mole}{1.16lmole} = 1$

Percentage of oxygen = % of total - % of Na

= 100% - 74.19%

= 25.818% of oxygen

Using the table;

The empirical formula is Na_2O.

Molecular formula = *No. of empirical formula* X *Empirical Formula*

No. of empirical formula = $\frac{mass\ of\ compound}{mass\ of\ empirical\ formula}$

Mass of empirical formula = 2 x (Na) +1 x (O)

= (2 x 23) + (1 x16)

= 62

Mass of compound given is 124g

No. of empirical formula=$\frac{124g}{62g}$

=2

Molecular formula = 2 (Empirical formula)

= 2 (Na_2O)

= Na_4O_2

Molecular formula is Na_4O_2.

Self-Assessment 3.15

1. (a) Define the following:
   i. Empirical formula
   ii. Molecular formula

(b) A certain compound contains 20% of Carbon, 5% of Hydrogen and the rest is Oxygen. If the relative molecular mass is 180. Work out its molecular formula.

3. Calculate the empirical formula of the compound containing 24g of Carbon, 6g of Hydrogen and 16 g of Oxygen (RAM: C=12, H=1, O=16)

Reacting masses.

When we write a balanced chemical equation the numbers of moles of reactants and products involved in the chemical reaction. If we consider the reaction between magnesium and oxygen.

Magnesium + Oxygen $\longrightarrow$ magnesium oxide.

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

This shows that 2 moles of magnesium react with 1 mole of oxygen to give 2 moles of magnesium oxide. Using the ideas of moles and masses we can use this information to calculate the equatntities of different chemicals involved.

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

2moles	1mole	2moles
2x24	1x (16x2)	2x(24+16)
48g	32g	80g

You will notice that the total mass of reactants is equal to the total mass of of products. This is true for any chemical reaction and it is known as the Law of conservation of mass. This law was understood by the Greeks but was first clearly formulated by Antoine Lavoisier in 1774. Chemists can use this idea to calculate masses of products formed and reactants used in chemical processes before they are carried out.s

The calculations becomes easier after balancing chemical equation as stated.

Example 15

10g of calcium react with dilute sulphuric acid which is excess to produce calcium sulphate and hydrogen gas.

- **Write down balanced equation**.

  $Ca_{(s)} + H_2SO_{4(aq)}$ $CaSO_{4\,(aq)} + H_2$ (g)
- **Work out number of moles of calcium metal used up in the reaction**.

No. of moles in log of Ca $= \frac{mass\ of\ calcium(Ca)}{molar1\ mass}$

Molar mass of calcium $= \frac{10g}{40g|mole}$

$=0.25$mole

Iii) How many moles of sulphuric acid are needed in the reactions?

From equation 1
1 mole of Ca reacts with 1 mole of H_2SO_4
Then
0.25mole of calcium (Ca) will react with x mole of (H_2S0_4)

$\frac{1\ \text{mole}}{1\ \text{mole}} = \frac{0.25\text{mol}}{X}$

X = 0.25mol.

Therefore, 0.25mole of H_2SO_4 will needed.

IV What volume of hydrogen evolved at standard temperature?

From the balanced equation
1 mole of Ca produces 1 mole of H_2

0.25mole of Ca will produce x

By simple proportion

$$\frac{1\ mole}{1\ mole} = \frac{0.25\ mole}{X}$$

X= **0.25 mole**

1 mole of H_2 of St.p occupies 22.4dm3

Then 0.25 mole of H_2 will occupies Xdm^3

By simple proportion

$$\frac{X\ mole}{1\ mole} = \frac{22.4dm3}{1\ mole}$$

X = 5.6 dm^3.

The volume of hydrogen will be 5. 6dm^3.

V. What is the mass of calcium sulphate formed?

1 mole of Ca produce 1 mole of $CaSo_4$

0.25 mole of ca forms X moles of $CaSo_4$

By simple proportion

$$\frac{1\ mole}{1\ mole} = \frac{0.25mol}{X}$$

X = 0.25 mole

Molar mass of $CaSO_4$ = 10+32+41(16)

= 136glmole

Mass of $CaSO_4$ = 136glmol X 0.25mol

= 34g.

The mass of calcium sulphate formed will be 34g.

Self-Assessment 3.16

15 g of magnesium react with dilute sulphuric Acid (H_2SO_4) which is excess to produce Magnesium sulphate and hydrogen gas.

i).Write down balanced equation.

ii).Work out number of moles of calcium metal used up in the reaction.

iii).How many moles of sulphuric acid are needed in the reactions?

iv).What volume of hydrogen evolved at standard temperature?

v).What is the mass of calcium sulphate formed?

3.15 Concertation of a solution.

You might agree that when drinking squash (Sobo), you put half of the squash in the cup and the other half is filled with pure water. This is done to reduce its concentration to the level that you can drink. This is similar when you add too much sugar in the cup of tea, you reduce its concentration by adding water. From this basic knowledge, **concentration** is defined as *the amount of solutes dissolved in a specific volume of the solvent*.

In JCE chemistry, you learnt that the substance that is dissolved is called **solutes** while the substance that dissolve another substance is called **solvents**. The salt and water are good examples of solutes and solvent respectively. Therefore, the mixture of solutes and solvents is called **solution**.

Chemists often need to know the concentration of a solution. Sometimes it is measured in grams per cubic decimeter (gdm^{-3}) but more often concentration is measured in moles per cubic decimeter (mol/dm^{-3}).

When one mole of a substance is dissolved in water and the solution is made up to 1 dm^3 ($1000cm^3$), **a 1 molar (1M)** solution is produced called **molar solution**. Chemists do not always need to make up such large volumes of solution.

Ways of expressing concentrations.

In chemistry science, concentration of the solution is expressed in different ways as follows:

- As percentage by mass or volume.
- As mass per volume e.g. $gldm^{3.}$
- As molarity.

Molarity.

As we already stated that molarity is one way of expressing concentration. It is defined as *the number of moles of solutes dissolved in 1dm^3 of a solution*. Mathematically;

$$\text{Molarity}=\frac{\text{Nomber of moles}}{\text{Volume (dm3)}}$$

This follows that:

No. of moles = Molarity X volume (dm^3)

1 liter= 1 dm^3= 1000cm^3 = 1000ml.

Example 16

Calculate the molarity of a solution made by dissolving 8g of sodium chloride (NaCl) in 200cm3 of water. (RAM: Na= 23 a= 35.5)

Solution

Molar mass (RFM) of mass = 1xNa+1xCl

= 1x23+K35.5

= 58.5glmol

Mass given of Na= 8g.

Then;

No. of moles = Mass of Nacl

Molar mass

$= \frac{8g}{58.5g\backslash mole}$

=0.14mol

Change volume to dm^3 by diving by $100cm^3$

Them; volume = $\frac{200cm3}{1000cm3}$ X $1dm^3$

== $0.2dm^3$

But:

Molarity = No of moles of NaCl

Volume (dm^3)

$= \frac{0.14mole}{0.2dm3}$

= $0.7mol/dm^3$

The molarity of 8g of NaCl is $200cm^3$ water is $0.7mol/dm^3$

However; 1 $mole/dm^3$ = 1M

Then, $0.7mol/dm^3$ = 0.7M

Example 17

A300mg of aspirin was completely dissolved in 10ml of water. The molecular formula of drug is $C_9H_8O_4$. Calculate its concentration.

Solution: Molar mass of C_9H_8O4 = 9xC=8xH+4xO

- = 9x12+8x1x4x16
- = 108+ 8+64
- = 180glmol

Change mass from mg to g.

$=\frac{300mg}{1000mg}$ X 1g

= 0.3g

= No of moles $=\frac{\text{Mass of } C_9H_8O_4}{\text{Molar mass}}$

$=\frac{0.3g}{180glmol}$

=0.0016mol

Change volume to ℓ

Then$=\frac{10m\ell}{1000m\ell}$ X $1dm^3$

=0.01dm3

Concentration$=\frac{0.0016mol}{0.01dm3}$

The concentration of 300mg of asprin in 10ml is 0.16ℓ mole/dm^3

Sometimes chemists need to know the mass of a substance that has to be dissolved to prepare a known volume at a given concentration. A simple method of calculating number of moles and so the mass of substance needed is by using the relationship:

Number of moles=**Concentration x volume of solution.**

(in moldm-3) **(in dm3)**

Example 18

Calculate the mass of potassium hydroxide (KOH) that need, to be used to prepare $500cm^3$ of a 2M solution in water (RAM: K= 29 O= 16 H= 1)

Solution

Molar mass of KOH= 1xK+Ox1+1xH

= 1x39+Ox1+1x1

= 40+16

= 56glmol

Change volume to dm^3.

Volume (dm^3) = $\frac{500cm3}{1000cm3}$ X $1dm^3$

= $0.5dm^3$

Molarity = $\frac{\text{No of moles}}{\text{Volume}}$

2M $= \frac{\text{No of moles}}{0.5}$

By simple proportion.

No of moles = 0.5dm3 X2mole/dm3

= 1 mole

But;

Mass = No of moles X molar mass (RFM)

= 1 mole X $\frac{56g}{1mole}$

= 56g

2M of KOH in 500cm^3 of water has a mass of 56g.

Self-Assessment 3.17

.Define the terms.

i. Concentration.
ii. Molarity.
iii. State three ways of expressing concentration.

2.Calculate the concentration (in moldm^{-3}) of a solution of sodium hydroxide (NaOH) which was made by dissolving 10g of solid sodium hydroxide in water and making up to 250cm3. (RAM: Na=23, O=16, H=1).

3.Calculate the mass of potassium nitrate (KNO_3) which needs to be used to prepare 200cm^3 of a 20moldm^{-3} solution.(RAM: K=39, N=14, O=16).

Concentration in percentages.

Some Acid have their concertation in percentage e.g. 98% sulphuric acid, 32% of HCl. Therefore, molarity of these percentages can be found by treating the **% as mass per 100mℓ of water.** 98% of sulphuric acid means that 98g of sulphuric acids have dissolved in 100mℓ of water.

Example.19

Calculate the molarity of 32% hydrochloric acid (HCl). (RAM: H=1, Cl=35.5).

Solution

32% means 32g in 100ml of water, therefore 32g of HCl has dissolved in 100ml of water.

Molar mass (RFM) of HCl = 1xH+1xCl

= 1x1+1x35.5

= 36.5glmol

No of moles $= \frac{\text{Mass of HCl}}{\text{Molar mass of HCl}}$

$= \frac{32g}{36.5glmol}$

$= 0.88moles$

Then;

Molarity $= \frac{\text{No of moles}}{\text{Volume } (\ell)}$

Volume $= \frac{100m\ell \text{ X} 1\ell}{1000m\ell}$

$= 0.1\ell$

Molarity $= \frac{0.88moles}{0.1\ moles}$

$= 8.8mol/dm^3$ (88M)

The concertation is 8.8M

Example 20

Calculate the molarity of 5% ethanoic Acid (CH_3COOH). (RAM: C=12, H= 1, O=16)

Solution

- 5% of CH3OOH means 5g of CH3COOH has dissolved in 100ml of water
- Molar mass (RFM) of CH3CooH

= 2xC+2xO+4(1)

= 2(12) + 2(16) +4(1)

= 60glmol

Change volume to ℓ by dividing by 1000ml and multiply by ℓ

Volume $= \frac{100ml}{1000ml}$x 1ℓ

$0.1e$

No of moles $= \frac{\text{Mass of C3HCOOH}}{\text{RFM (Molar mass)}}$

$= \frac{5g}{60glmol}$

Then;

Molarity $= \frac{No.of\ moles}{\text{Volume } (\ell)}$

$= \frac{0.083mole}{0.1e}$

= 0.83mol/ℓ (0.83M)

The concentration is 0.83M.

Self-Assessment 3.18

Calculate the morality of 20% sulphuric acid (H_2SO_4). RAM: H=1, S=32, O=16.

3.16 Standard solution.

From the previuos explaination, the molarity 3M (3mol/dm^3) of sodium hydroxide (NaOH) means that 3moles of NaOH have been dissolved in 1000cm^3 (1dm^3) of water. Therefore, the concentration of sodium hydroxide (NaOH) is known, which is 3M.

In chemistry, the solution of known concentration is called **standard solution**. In laboratories, both school and clinical, we can prepare standard solution using different ways.

Ways of preparing starndard solution.

- By dissolving solutes
- By dilution of concentrated standards solution.

Preparing standard solution by dissolving the solutes;

Example 21

Describe how 500cm3 of 0.2M solution of sodium chloride (NaCl) can be prepared using sodium chloride crystals (RAM: Na = 23, Cl= 35.5)

Solution

Molar mass of NaCl =1xNa+ 1xCl

= 23+35.5

58.5glmol

Change volume to dm^3

Volume = $\frac{500cm3}{1000cm3}$ x 1dm^3

= 0.5dm^3

Molarity = 0.2M (0.2 mol/dm^3)

But

Molarity = $\frac{\text{No of moles}}{\text{Volume (dm3)}}$

$\frac{0.2mol}{dm3} = \frac{\text{No.of mole}}{0.5mole}$

= 0.1mol

Find mas of NaCl.

Mass of NaCl = No of moles X molar mass

$$= 0.1\text{mol X } \frac{58.5\text{g}}{1\text{mole}}$$

= 5.85g of NaCl

Procedures.

- Measure 5.85g of NaCl in a beaker using triple bean balance
- Pour this into 500cm^3 volumetric flasks
- Add 500cm^3 of distilled water
- Stir using a glass rod until all NaCl crystals dissolved completely.

Example 22

Describe how 250cm^3 of 1M copper sulphate solution could be prepared using hydrated copper sulphate ($CusO_4.SH_2o$). The molar mass of $CusO_4.5H_2O$ is 250glmol.

Solution

Molar mass = 250glmol

$$\text{Volume} = \frac{250\text{cm3}}{1000\text{cm3}} \text{ X } 1\text{dm}^3$$

$= 0.25dm^3$

No of mole = molarity X volume (dm^3)

Molarity = 11 mole/dm^3 (1M).

$$\text{No of moles} = \frac{1\text{mol}}{1\text{dm3}}\text{X } 0.25(\text{dm}^3)$$

= 0.25moles

Mass of $CuSO4.5H_2O$ = No of moles X Molar mass

= 0.25moles X 250gmol

= 62.5g of $CusO_4.5H_2O$

Procedures.

- Measure 62.5g of hydrated copper sulphate crystals in a beaker using triple beam balance
- Pour this amount in 250cm^3 beaker or volumetric flask
- Add 250cm^3 of distilled water
- Using glass rod, stir until all crystals have dissolved completely
- Now you have prepared standard solution.

Example 23

Describe how can you prepare 10% solution of sodium chloride (NaCl).

Materials

- Sodium chloride
- Distilled water
- beaker
- Stirring rod
- 100ml measuring cylinder

Procedures

- A 10% solution of sodium chloride is the same as 10g of sodium chloride (NaCl) in 100ml of water
- Measure 10g of sodium chloride (NaCl) crystals in a beaker using a triple beam balance.
- Add this amount in 100ml of beaker
- Add 100ml distilled water
- Stir using glass rod until all sodium chloride (NaCl) crystal have completely dissolved.

Preparation of standard solution by dilution of concentrated solution.

Dilution is the addition of more solvent to concentrated solution. In dilution; **C1 or M1** is the concertation of concentrated solution while **C_2 or M_2** is the the concentration of diluted solution.

The volume of concentrated solution is taken as **V_1** and diluted is taken as **V_2**.

Therefore, the formula for dilution is **$C_1V_1 = C_2V_2$** or **$M_1V_1 = M_2V_2$**, where **C_1** or **M_1** and **C_2 or M_2** are the initial and final concertation respectively of given solution while **V_1** and **V_2** are the initial and final volume of the respective solutions.

Example 24

A soluble of salt concertation 5glℓ and a volume of 250cm^3 is diluted by adding water to knew volume of 1.5 litres. What is the final concertation?

Solution

Given; C_1= 5glmℓ V_1= 250cm^3 V_2= 1.5ℓ

Change 250cm^3 to litres (ℓ)

Volume = $\frac{250cm3}{1000cm3}$ X1d

= 0.25ℓ

$C_1v_1 = C_2V_2$ or $M_1V_1 = M_2V_2$

C_2=?

By substitution in the formula

$5X0.25 = C_2X1.5$

$C_2 = \frac{5g\ X0.25}{1.5}$

C_2=0.83glℓ

The final concentration is 0.83glℓ.

Example 25

$60cm^3$ of a solution whose concentration is $15glcm^3$ diluted with distilled water by raising its volume $80cm^3$. Calculate the concentration of new solution.

Solution

Given; C_1= $15glm^3$ V_1=$60cm^3$ V^2=$80cm^3$

C_2=?

From the formula

$C_1V_1=C_2V_2$

By substitution of given value;

$15glmc^3X\ 60cm^3= 80cm^3XC_2$

$\frac{15gX60}{80cm3}$= C_2

C_2= $\frac{600g}{80cm3}$

= $11.25glcm^3$

The concentration is $11.25g/cm^3$

Example 26

Describe how you can prepare the standard solution of 0.2M of sodium hydroxide (NaOH) solution from 6M. The volume of 0.2M of sodium hydroxide (NaOH) is $50cm^3$.

Solution

From given data: C_1=6M, C_2= 0.2M, V_2=$50cm^3$ and V_1=?

Using the formula: $C_1V_1=C_2V_2$

$6MxV=0.2Mx\ 50cm^3$

Change volume to dm^3

$\frac{50cm3}{1000cm3}$ X $1dm^3$

Then

$V1 = \frac{0.2mol\ 1dm3X\ 0.05dm3}{6mol\ 1dm3}$

$= 0.01dm^3$

The volume of C_1 is $0.01dm^3$.

The ***procedures*** in dilution are,

a. Measure $10cm^3$ of 6M of sodium hydroxide (NaoH) solution using measuring cylinder.
b. Pour this into $50cm^3$ volumetric flask
c. Add $50cm^3$ of distilled water
d. Now you have prepared standard solution.

Example 27

A laboratory assistant has 2M NaCl solution. Describe how he would prepare $50cm^3$ of 0.5M NaCl from the standard solution.

Solution.

Given; C_1= 2M, V_1= unknown

C_2= 0.5M V_2= $250cm^3$

Using the formula

$C_1V_1 = C_2V_2$, substitute given values

$0.5 \text{ X } 0.25dm^3 = \frac{2Mol x\ V1}{dm3}$

$\frac{0.5mol\ x\ 0.25dm3}{dm3} = \frac{2Mol\ xV1}{dm3}$

$\frac{0.125mol}{2mole/dm3} = V_1$

$V_1 = 0.0625dm^3$ ($625cm^3$)

Procedures

- Measure exactly $62.5cm^3$ of the standard solution
- Place it in a $250cm^3$ volumetric flask
- Add water up to 250cm mark using special distilled water.

Self-Assessment 3.19

1. Define the standard solution.
2. State two ways of preparing standard solution.
3. Describe how 500cm3 of 0.4M solution of magnesium chloride (MgCl2) can be prepared using magnesium chloride crystals (RAM: Mg = 23, Cl= 35.5)
4. Describe how can you prepare 10% solution of sodium chloride (NaCl).
5. A soluble of salt concertation 10glℓ and a volume of 200cm^3 is diluted by adding water to knew volume of 0.5 litres. What is the final concertation?
6. Describe how you can prepare the standard solution of 0.3M of sodium hydroxide (NaOH) solution from 2M. The volume of 0.1M of sodium hydroxide (NaOH) is 40cm^3.

3.17 Tiltration.

Tiltration is commonly used to prepare standard solution and to find concentration of unknown solution. By definition, tiltration is the gradual addition of base to an acid or an acid to a base **until the end point** is reached. **The end point** is the balanced point where by an acid is completely neutralized by a base.

Tiltration process is demonstrated by experiment below.

Experiment 3.2

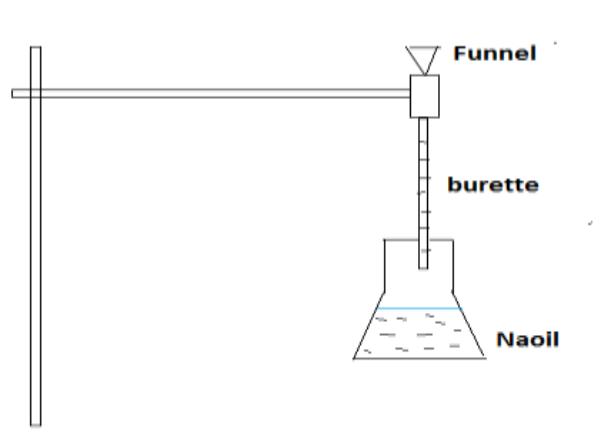
Aim: To demonstrate\determines the concentration of NaOH using 0.3M HCl by titration.

Materials

- Burette,
- three conical flasks,
- funnel
- clamp stand,
- 0.3M HCl,
- phenolphthalein indicator,
- sodium hydroxide (NaOH).

Procedures

a. Set up the apparatus as shown below



b. Pipette 25cm^3 of sodium hydroxide (NaOH) using a pipette into a conical flask.
c. Acid 2-3 drops of phenolphthalein indicator note colour change.
d. Fill the burette 0.3M HCl solution using a filter funnel until it is filled to zero mark
e. Then remove the funnel.
f. Slowly add the HCl from the burette to the solution hydroxide (NaoH) into the conical flask while swilling / shaking until pink colour disappear.
g. Note and record volume of HCl used to the end point.
h. Repeat the experiment until consistent results are obtained

NB: add 0.5cm^3 at each time to get good results.

Table of results

Initial volume of HCl	**Final volume of HCl**	**Volume of HCl used (charge of volume**
0	11.28	11.2
11.2	22.4	11.2

Use the formula = $\frac{\mathbf{M1V1}}{\boldsymbol{M\ acid}} = \frac{\mathbf{M2V2}}{\boldsymbol{Ma\ alkali}}$ **t**o find the concentration of sodium hydroxide (NaoH),

Where.

M_1 = concentration of acid

V_1 = volume of acid used

Macid= Number of moles of acid shown in the balanced chemical equation.

M_2= concentration (molarity) of the alkali used

V_2= volume of alkaline used

M alkali= Number of moles of the alkali shown in the balanced chemical equation.

*What is the volume of sodium hydroxide used**?***

25ml

What is the concentration of hydrochloric acid used?

0.3M

Write down indicator colour change.

purple or pink

Calculate the average volume of HCl used.

Solution.

Initial volume= 11.2ml

Final volume= 11.2ml

Total volume= (11.2+11.2) ml

= 22.4ml

Average volume = $\frac{22.4\text{ml}}{2}$

= 11.2ml

Write balanced equation for the reaction between hydrochloric Acid (HCl) and sodium hydroxide (NaOH).

$$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O(l)$$

Calculate concentration of sodium hydroxide (NaOH) in the solution.

Using the formula: $\frac{\mathbf{M1V1}}{\mathbf{Macid}} = \frac{\mathbf{M2V2}}{\mathbf{Makali}}$

M1=0.3m **M_2**=?

V1=11.2ml **Makal**i= 1mol

Macid= 1mol

Substitute given value into the formula

$\frac{0.3\text{mx}11.2\text{ml}}{1\text{mol}} = \frac{\mathbf{M2}\text{x}25\text{ml}}{1\text{mole}}$

Make **M_2** subject of formula

= 0.1M

The concentration of sodium hydroxide (NaoH) is 0.1m.

Sources of errors

- Material used
- From observer, can record wrong results.

3.18 Determing the yield of chemical reaction.

Here, we apply the same theory of the reaction masses. Let's consider the equation below.

$$Mg(OH)_{2(aq)} + NaCl_{(aq)} \longrightarrow MgCl_2 + NaOH$$

- Balance the equation first then interpret
- The equation means 1 mole of magnesium hydroxide reacts with 1 mole of sodium chloride to produce 1mol of magnesium chloride and 1 mole of sodium hydroxide.

Let's see another example;

$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O\ (l)$

This means; 1 mole of methane (CH4) needs 2moles of oxygen (O) to produce 1mole of carbon dioxide Co_2 and 2moles of water. This interpretation is after balancing the equation.

So what?

We calculate the theoretical yield by

- Changing the statement to moles
- Changing moles to mass

How do we go about that?

- Firstly, we write the chemical equation
- Balance it currently
- Change it to the statement of moles
- Change mole to the statement of mass

NB: Limiting reactant determine the quantity of the products of the chemical reaction.

Example 28

Propane (C_3H_2) burns in oxygen (O_2) to produce carbon dioxide (O_2) and water (H_2O). Calculate the mass of carbon dioxide (CO_2) produced if 88g of propane is used.

Solution

First write equation and balance it:

$C_3H_{8(g)} + 5O_{2\,(g)} \longrightarrow 3CO_{2(g)} + 4H_2o\ (l)$

Change it to statement of mass; using molecular formula mass:

44g of C_2H_8 + 160g of $5O_2 \longrightarrow$ 132g of $3CO_2$ + 72g of $4H_2O$

Then

If 44g C_3H_8 produces 132g CO_2

Then 88g of C_3H_8 produces? **More**

By simple proportion;

$$\frac{44g}{88g} = \frac{132g}{X}$$

(*if we let x to be mass of CO_2 produced)*

Cross multiply

$$\frac{44g \times X}{44g} = \frac{88 \times 132g}{44g}$$

X = 264g of CO_2

Therefore, 88g of C_3H_8 produces 264g of CO_2.

Example 29

Methane (CH_4) burns in excess oxygen (O_2) to produce carbon dioxide (Co_2) and water (H_2O). Work out mass of methane required if 40g of water is produced.

Solution

Chemical equation: $CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O(l)$

Change it to statement of mass: using molecular formula of each substance used.

15g of CH_4 + 64g of $2O_2$ $\longrightarrow$ 44g of CO_2 + 36g of $2H_2O$

Then:

By simple proportion

16g of CH_4 produces 36g of water

Xg of CH_4 produces 40g of water

Cross multiply

$$\frac{30x}{30} = \frac{40g \times 16}{30}$$

$$X = \frac{64}{3}$$

= 21.3g of CH_4

Therefore, 21.3g of CH_4 produces 40g of water.

Self-Assessment 3.20

1. ai Define “Titration”
2. what is the use of phenolphthalein indicator in titration process
3. Butane (C_4H_{10}) burns in excess oxygen (O_2) to produce carbon dioxide (CO_2) and water (H_2O). If 80g of Butane was used, calculate the mass of carbon dioxide produced.
4. Define “end point”
5. Given 0.5M of sodium hydroxide (NaOH) and 20cm^3 of hydrochloric Acid (HCl). Which one is a standard solution?

3.19 Heat of reaction.

In chemical reaction, heat energy is either released or given out to the surroundings. This causes the container in which chemical reaction is taking place to get hot. Heat energy can also be absorbed from the surroundings. Hence, the temperature of the medium is lowered. This causes the reactants or products to be favored depending on the conditions.Therefore, we are going to look at **endothermic and exothermic reaction** in terms of heat energy.

Exothermic reaction.

The reaction that releases heat energy to the surrounding. This causes the temperature of the medium to be rised.

Examples:

The mixture of sulphur acid and ~~water~~ Causes the rise of the temperature to the surroundings**.** $\mathbf{H_2SO_4 + H_2O \quad H^+ + SO_4^-}$

a. The reaction between sodium hydroxide (NaoH) and hydrochloric Acid (HCl). $\mathbf{NaOH_{(aq)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O(l)}$

b. Combustion

c. Respiration

d. The dissolving of sodium hydroxide (NaoH) in water

Exothermic reaction favors products produced, hence, it causes the equation to be switched to the right.

Endothermic reaction.

The reaction that absorbs heat energy from the surrounding. This decreases the surrounding temperature.

Examples.

i. Dissolving Ammonium Nitrate (NH_4NO_3) in water (H_2O)

$NH_4NO_{3(aq)} + H_2O$ ~~(l)~~ $\longrightarrow \quad NH_{4(aq)} + NO_3^-{}_{(aq)}$

ii. The evaporation of water to water vapour

$H_2O_{(l)} + \text{heat} \longrightarrow H_2O_{(g)}$

iii. Photosynthesis

iv. The reaction between nitrogen and oxygen to produce Nitrogen dioxide (NO_2).

$N_{2(s)} + 2O_{2(g)} \longrightarrow 2NO_{2(g)}$

NB: Loss of heat energy in endothermic reaction is a result of interaction between the system and its surroundings.

Heat enerygy level diagrams.

These are the diagrams that plot the heat energy of the reactants and products against the type of reactions (progress of reaction**).** Hence, the graphs are of **heat energy** against **progress of reaction.** The energy change that takes place during a chemical reaction can be shown by an energy level diagram.

Bond energy is the amount of energy in kilojoules (kj) associated with the breaking or making of one mole of chemical bonds in a molecular element or compounds.

Enthalpy.

The enthalpy is the energy stored in bonds. The symbol is "**H** while the **Enthalpy change** is the change in energy going from the reactants to the products. The enthalpy change is also called Energy change of the substrates or reactants and products

The enthalpy change is represented by a symbol "Δ**H**". Therefore, the symbol Δ**H** is called heat of reaction. This is pronounced as "delta H". This enthalpy change (ΔH) is found by subtracting *reactants* from *products*. Mathematically,

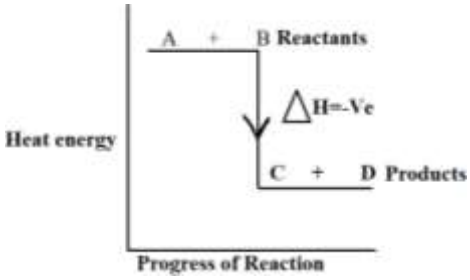
- **Enthalpy change**=enthalpy of products-enthalpy of reactants
- **ΔH = H of products- H of reactant**

In *exothermic reaction*, ΔH is negative, while in *endothermic*, ΔH is positive. The Exothermic reactions are also called **exergonic reaction** while endothermic reaction are called **endergonic reaction.**

Exothermic (Exergonic) reaction is very important in the body system for the action of the enzymes in the metabolisms reaction. Therefore, in biochemistry (medical or veterinary),you will study much of the exothermic (exergonic reactions).

Heat energy diagram in exothermic reaction.

- The reactants are at higher energy level that the products**.**
- All exothermic reaction, ΔH is negative i.e $\Delta H < 0$



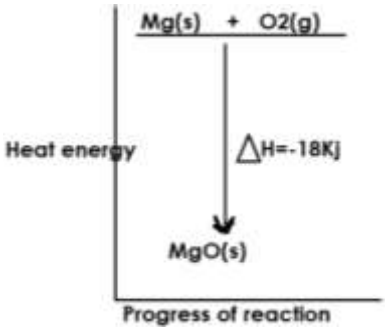
Therefore, ΔH= final heat energy – initial heat energy.

Where **A** and **B** are reactants and **C** and **D** are the products.

Draw heat energy diagram for the following reaction.

$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(l)}$ ΔH = -18kj

The figure 3.19a: Heat energy level diagram



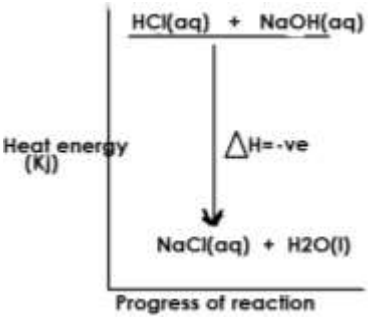
Given heat reaction between HCl and NaCl to produce NaCl and H_2O is exothermic, draw energy level diagram of the reaction.

Solution

Write the balanced equation first

$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(aq)}$

The figure 3.19b: Heat energy level diagram



Ways of writing chemical equation, which include Heat of Reaction (ΔH) in Exothermic Reaction.

The chemical equation is written in full, Heat of Reaction (ΔH) is written separately using comma at the end of the equation.

$2Mg_{(s)} + O_{2(g)} \longrightarrow MgO_{(s)}, \Delta H = -18kj$

The heat of reaction is written as part of the reaction. Exothermic reaction shows ΔH as a part of the products.

The negative sign is not written for ΔH in this case;

$2Mg_{(s)} + O_{(g)} \longrightarrow 2MgO_{(s)} + 18kj$

NB: Make sure that arrows should point down when drawing exothermic diagrams.

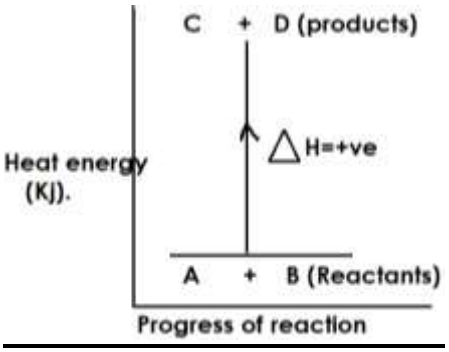
Heat energy level diagram in endothermic reaction.

The reactants are below and the products are above. Heat of reaction (ΔH) is positive for all endothermic reaction. Make sure that arrows should point upward when plotting these graphs.

In general, the graph should appear like this

The figure Heat energy level diagram in endothermic reaction

Where **A** and **B** are **t**he reactants and **C** and **D** are products of reaction.



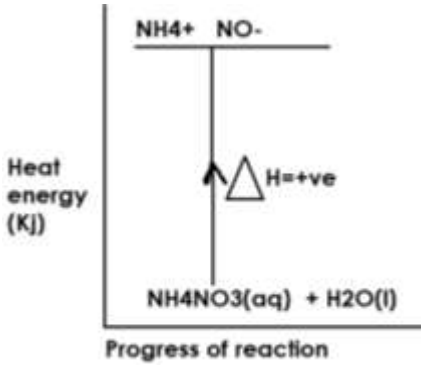
The dissolving of ammonium nitrate in water is endothermic reaction. Draw the energy level diagram.

Solution

Write the balanced equation first

$NH_4NO_{3(g)} + H_2O_{(l)} \quad NH_4^+ + NO_3^-$

The Figure 3.19c. Heat energy level diagram

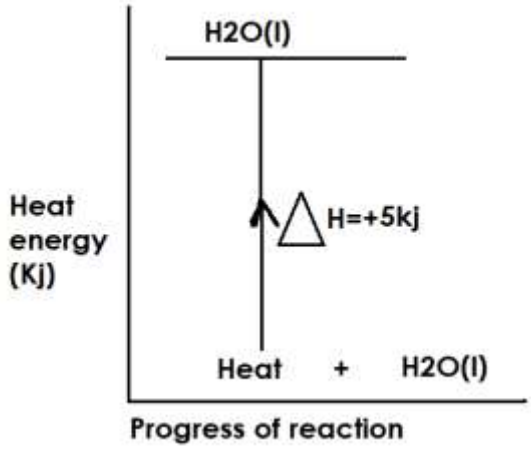


Draw heat energy level diagram for the following reaction.

$H_2O_{(l)}$ + heat ⟶ $H_2O_{(g)}$ ΔH= =5kj.

Solution.

The figure 3.19d: Endothermic Reaction



Experiment 3.3

Aim: To distinguish endothermic reaction from exothermic reaction.

Materials:

- Two test tubes,
- test tube rank, Measuring cylinder, thermometer,
- spatula(spoon)
- tap water, ammonium chloride (NH_4Cl),
- sodium hydroxide (NaOH) pellets.

Procedures

(a) Put $5cm^3$ of tap water in each test tube
(b) Measure the initial temperature of water in each test tube
(c) Records the results in the table.

(d) Add half spatula of the full amount of Ammonium chloride (NH4Cl) in one test tube and shake gently.
(e) Measure the temperature of Ammonium chloride solution
(f) Record the results in the table
(g) Repeat the steps a-f with sodium hydroxide (NaOH) pellets.

Table of Results

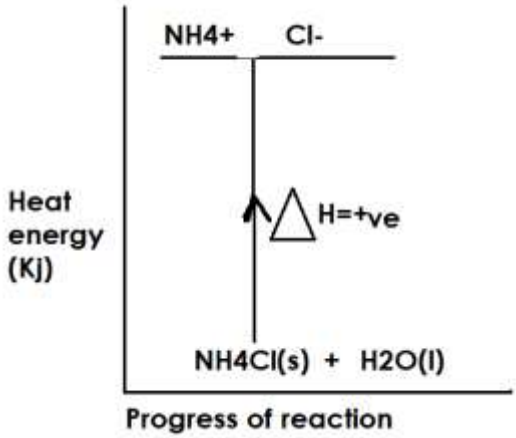
Solution	**Initial (^{o}c) temperature**	**Final temperature(^{o}c)**	**Temperature (f-T) change (ΔT)**	***Other changes)***
Ammonium chloride	25^{o}c	22^{o}c	-3^{0}c	Test tube cold
Sodium hydroxide	25	55	f-t 30	Test tube hot

1.State whether the change in each test tube is endothermic or exothermic.

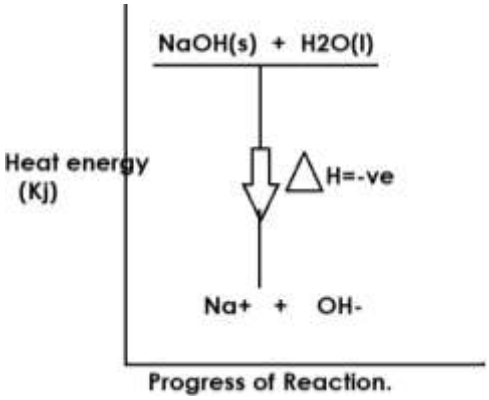
Ammonium chloride (NH_4Cl) = **Endothermic**

Sodium hydroxide (NaOH**) = Exothermic**

a) Draw energy level diagram to illustrate the dissolving of ammonium chloride and sodium hydroxide (NaOH).



The figure 3.19e:Heat energy level diagram in Ammonium chloride (NH_4Cl).



The figure 3.19f: Heat energy level diagram in Sodium hydroxide (NaOH)

Why bond breaking is endothermic while bond making is exothermic reaction?

Bond breaking is endothermic because its requires absorption of heat from the surrounding while bond making is exothermic, because making new bond to from products releases heat energy to the surrounding.

When bond energy change is negative, we always assume that reaction is ***exothermic****, but when it is positive, we conclude that the reaction is* ***endothermic*****.**

3.20 Dertemining exothermic and endothermic using bond energies.

As we have said that bond energy is associated with the breaking or making the other new chemical bonds of the compound; it can be used to determine whether reaction is exothermic or endothermic by calculation using provided data of bond energy.

How possible is this? You draw the structure of covalent bond of each compound in the reaction. Therefore, you identify the type of bonds.

- Each type of bond has got each own amount of energy.
- Count the number of bond of each structure and then multiply with its energy.
- Do this for reactants and products separately
- At the end, find the energy difference (ΔH). Therefore,

Energy change (ΔH) = energy required to break the bond – energy given off to form the bond

The table 3.20 The bond energies.

Bond type	**Energy of the bond ($KJmol^{-1}$)**
C-H	435
O=O	497
C=O	803
H-O	464

C-C	347
C-O	358
C=C	610
C≡C	835
C-F	495
C-CL	339
H-H	160
F-F	156
Cl-Cl	339
-Br	280
Br-Br	193
O-O	166
N-N	390

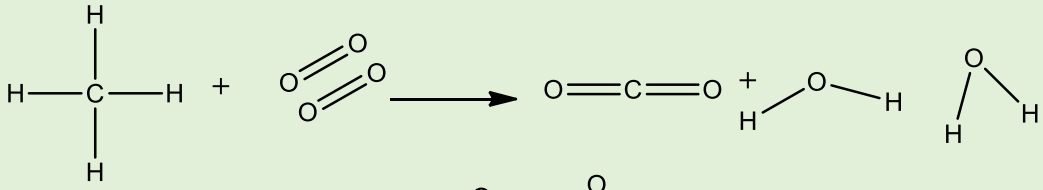
Example 30.

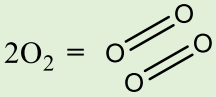
Determine heat change

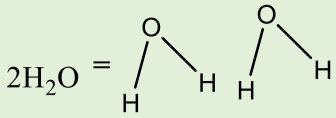
$CH_{4(g)} + 2O_{2(g)}$ $CO_2 + 2H_2O$ and determine if it is endothermic or exothermic.

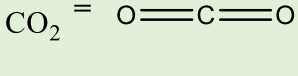
Solution

Draw the structure of reactants and products

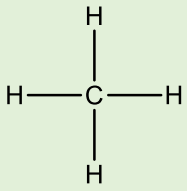




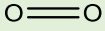




Count number of bond in each compound



Has 

$2O_2$ Has 2 

Then

C- H has bond energy of 435 (table)

4X 435

= 1740kj

O=O has bond energy of 497

2O= O = 2x497

= 994kj

Total energy required; 1740kj + 994kj

= 2734KJ

In products

CO_2 has 2C=O

C=O has bond energy of 803,

= 2x803kj

=1606KJ

$2H_2O$ has 4 O-H

But **O-H** has bond energy of 464KJ.

4O-H

4X464KJ
= 1850KJ

Total energy released = 1606KJ + 1856KJ

= 3462KJ

Total difference (ΔH) = energy required – energy give off

= 2734KJ – 3462

= - 728KJ

The enthalpy change is negative

Therefore, the reaction is **exothermic**

Example 31

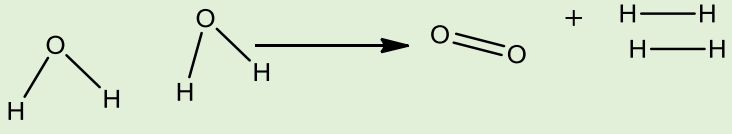
Determine whether the following reaction is exothermic or endothermic.

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

Solution

Draw the structure of reactants and products

$2H_2O$ $O_2 + 2H_2$



$2H_2O$ has 4 O-H

But **O-H** has 464KJ,

4O-H = 4X464KJ
= 1856KJ

O_2 has 1 O=O

O=O has 497KJ

$2H_2$ has 2H-H

H-H has 160KJ

2H-H= 2X160KJ

= 320KJ

Total energy given off = (320 + 497) KJ

= 817KJ

Energy difference = energy required-energy given off

Energy different (ΔH) = (1856-817) KJ

= 1039KJ

Enthalpy change is positive; then the reaction is **exothermic.**

NB: Don't memorize bond energy, you are given in **examination room** or on the **examination paper.** Otherwise, you will waste your time to study organic chemistry in the next chapter.

Self-Assessment 3.21

1. (a) Methane (CH_4) reacts with oxygen (O_2) to produce carbon dioxide CO_2 and water (H_2O).
   i. Write balance equation
   ii. If energy Change is -18KJ, draw energy level diagram
   iii. Given that **O-H** is 646Kj, **O=O** is 435KJ. Determine whether the reaction is exothermic or endothermic.
2. (a) Explain why bond breaking is endothermic and bond breaking exothermic?
   (b) Define "**bond energy**

3.21 Summary

Stoichiometry is a branch of chemistry which is based on the balance equation. This helps us to work out mole. Moles are substances that contain $6.023X10^{23}$ particles. This is found by diving mass of substance by relative molecular mass (molar mass). This theory also helps us to understand concentration which is expressed in three ways; as molarity, percentages and mass per volume (density). This further helps us to understand solution, exothermic and endothermic reaction.

Therefore, the different quantities calculated in analytical Chemistry are based on the balanced equation to make sure that Mass of reactants equals mass of the products. This is very important basic in Chemistry which is commonly used in different fields.

3.22 Revision Exercise

1. Define Moles.
2. Calculate the number of atoms in 20g of sodium chloride (NaCl). RAM: Na=23, Cl=35.5.
3. What is Morality?
4. State three ways of expressing concentration of solution.
5. Calculate the concentrations of 40g of Hydrochloric acid (HCl) which has dissolved in 250cm3 of water.
6. Calculate the percentage by mass of water in the salt hydrate $CaSO_4$. $7H_2o$. (RAM: H= 1, O=16, Ca=40, S=32)
7. Find the empirical formula of the compound that has the following percentage composition by mass of element C=40%, H = 6.67% and O = 53.33%.
8. Calculate relative molecular mass of a gas $10cm^3$ of the gas volume which has a mass of 2g. (At 25^oC, 1 mole occupies $24dm^3$.)
9. How many moles of Oxygen gas (H_2) are in $24cm^3$ at room temperature?
10. Define Standard solution.
11. State two ways of preparing standard solution.
12. Describe how $200cm^3$ of 2M magnesium sulphate solution could be prepared using hydrated copper sulphate ($CuSO_4.SH_2O$). (RAM: Mg=24, S=32, O=16, H=1).
13. Describe how you can prepare the standard solution of 0.2M of sodium hydroxide (NaOH) solution from 5M. The volume of 0.3M of sodium hydroxide (NaOH) is $100cm^3$.
14. Propane (C_3H_2) burns in oxygen (O_2) to produce carbon dioxide (O_2) and water (H_2O). Calculate the mass of carbon dioxide (CO_2) produced if 80g of propane is used.
15. Draw heat energy diagram for the following reaction.
    $2Mg_{(s)} + O_{2(g)}$ $2MgO_{(l)}$ $\Delta H = -18kj$
16. Why bond breaking is endothermic while bond making is exothermic reaction?
17. Methane (CH_4) reacts with oxygen (O_2) to produce carbon dioxide CO_2 and water (H_2O). Given that **O-H** is 646Kj, **O=O** is 435KJ. Determine whether the reaction is exothermic or endothermic.

3.23 Practical Question.....1

You are provided with the following materials; Magnesium Ribbon, tap water, test tube, hydrochloric Acid (HCl), measuring cylinder and thermometer.

Procedures

(a) Put 2cm3 of hydrochloric acid (HCl) test tube using measuring cylinder.
(b) Measure the temperature of the acid
(c) Record the results in the table as initial temperature
(d) Drop a piece of magnesium ribbon in the test tube
(e) Measure the temperature and record of as a final.

Table of results

Substance	Initial temperature (^{o}c)	Substance	Final temperature (^{o}C)	Temperature change
Hydrochloric acid (HCl)		HCl + magnesium ribbon		

(a) Write the balance equation between magnesium ribbon and hydrochloric acid (HCl)

_____

(b) Identify the type of the reaction

_____

(c) Draw energy level diagram of the reaction

-

_____

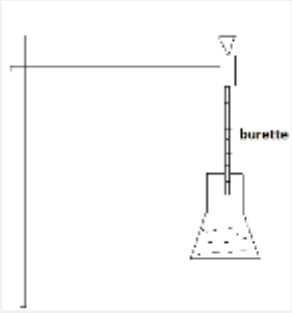
(d) Mention two sources of errors

_____
_____
_____

3.24 Practical queston.....2

1. You are provided with 20ml of 0.1m sodium hydroxide (NaOH) and hydrochloric Acid (HCl) solution, measuring cylinder,10cc syringe, retort stand, funnel and phenolphthalein indicator.

Procedures

a. 

b. Set up the apparatus as shown below
c. Measure 1ocm of the 0.1m sodium hydroxide (NaoH) are put it in conical flask
d. Add two drops of phenolphthalein indicator into the conical flask
e. Pour 20ml of HCl into the burette or syringe
f. Slowly add the hydrochloride Acid from the burette or syringe into the conical flask until colour charge is observed.
g. Record the results in the table

TABLE OF RESULTS

Initial volume (ml)	Final volume (ml) or HCl	Volume of Acid used (ml)

(a) Write balanced equation of the reaction of sodium hydroxide (NaOH) and hydrochloric Acid (HCl).
(b) Calculate the concentration of hydrochloric Acid (HCl) used

---

(c) Mention two sources of errors in this experiment
(d) What is the way of reducing errors in this experiments?
(e) Mention one function of phenolphthalein indicator in this experiment

-

(f) Which one is the standard solution in this experiment

UNIT 4

Alkanols

Learning objectives.

By the end of this unit, you should be able to;

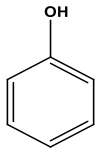
a) Identify the functional group of alkanols.
b) Draw and name the structures of first ten unbranched alkanols
c) Write the molecular of alkanols given the number of carbon atoms.
d) Classify alkanols as primary, secondary and tertiary.
e) Describe the methods of preparing alkanols.
f) Describe the properties of alkanols.
g) Explain the uses of alkanols.

4.10 Introduction.

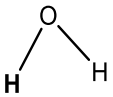
In JC Chemistry you learnt that hydrocarbon compounds include alkanes, alkenes and alkynes. These compounds contains hydrogens and carbon atoms only in their formulas.

Alkanol is another organic compound that contains oxygen atom in the formula. It is also called alcohols. The alkanols contains a functional group called **hydroxyl (-OH)** group. This can be represented as **R-OH**, where **R=alky group** like $CH_3CH_2CH_2^-$. For instance; $CH_3CH_2CH_2$**OH**.

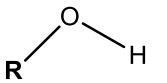
If **R=aryl** (aromatic group, e.g. **phenyl,** $\mathbf{C_6H_5^-}$), the alcohol will be called **Phenol**. For example;



The functional group of alcohol determine its properties. Alcohol can be viewed structurally as organic derivations of water in which one of the hydrogen atom of water is replaced by an **alkyl group (R group)**.



Water



Alcohol

Alcohols (Alkanols) are not hydrocarbons because they contain oxygen atoms in their formula. As such, there are called **oxy-carbons.**

4.11 Nomenclature of alkanols (alcohols).

You have already learnt how to name alkanes and alkenes. In this chapter we will build upon the same knowledge to name alkanols.

Nomenclature of alkanols is just the system of naming alcohols (alkanols). The ending of these compounds is **– ol.** Their parent compound name is derived from the parent name of alkane with the corresponding number of carbon atoms.

This is done by replacing the final – **e** in alkane with –**ol**. For instance; if the final-**e** in methan**e** is replaced by-**ol**, then it will be methan**ol**. This simply means that alkanols have the suffix of **-anol**.

Table 4.11: Names of alkanols.

Prefix	**Suffix**	**Complete name**
Meth-	-anol	Methanol
Eth-	-anol	Ethanol
Prop-	-anol	Propanol
But-	-anol	Butanol
Pent-	-anol	Pentanol
Hex-	-anol	Hexanol
Hept-	-anol	Heptanol
Oct-	-anol	Octanol
Non-	-anol	Nonanol
Dec-	-anol	Decanol

The prefix determine the number of carbon atoms in the compounds while the suffix shows the family (group) of the compound.

4.12 Formula of alkanols.

Alkanols also poses different formulas like alkanes and alkenes. The formulas helps us to understand the properties of these compounds.

Molecular formula.

We can derive the molecular formula of alkanols by using the general formula $\mathbf{C_nH_{2n+1}OH}$ where n is number of carbon atoms, n=1,2,3...

For example, to write molecular formula of methanol, we believe that **meth**-means **one** carbon atom present in compound, therefore, **n= 1**. Substitute **1** for **n** in the formula $\mathbf{C_nH_{2n+1}OH}$

$C_1H_{2(l)}$ +1 OH

Solving mathematically;

$C_1H_{2+1}OH$
$C_1H_{2+1}OH = CH_3OH$

The molecular formula of methanol is $\mathbf{CH_3OH}$.

The molecular formula of the other compounds of alkanols can be found in the same way

Activity 4.0

Write all molecular formulas of alkanols using the formula $C_nH_{2n+1}OH$.

Drawing molecular structure of alkanols.

Here, we go with procedures as follows:

- Draw the carbon skeleton with four bonds each
- Join functional group on one of the carbon atom **(i.e. –OH).**
- Join the hydrogen atoms to the bonds to complete the structure.

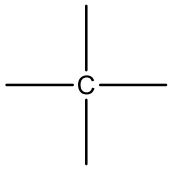
For example.

Write the molecular structure of ***methanol*** *and* ***ethanol*****.**

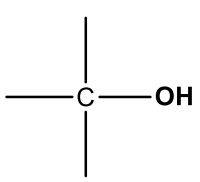
Methanol

i. Molecular formula CH_3OH

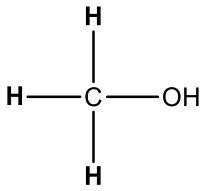
ii. Carbon skeleton



iii. Join functional group **(OH)**



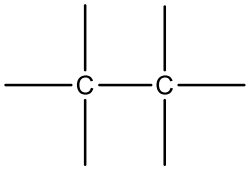
iv. Join hydrogen atom **(H)**



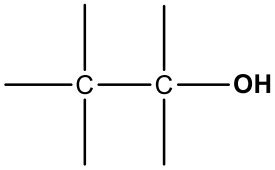
Ethanol

Molecular formula C_2H_5OH

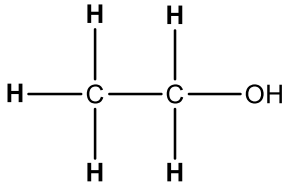
Carbon skeleton



Join functional group **(-OH)**



Join hydrogen atom **(H)**



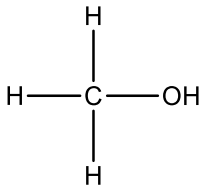
Writing condensed of alkanols.

These are also procedures that are necessary for us to achieve writing good condensed formulas. The procedures are as follows;

i. Draw molecular structure of compound.
ii. Write down the carbon atom with its hydrogen and functional group bonded. For example, write down the condensed formula of methanol and ethanol.

Ethanol

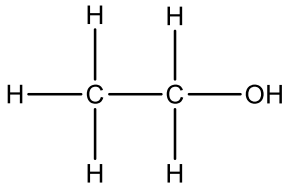
The structure.



Condensed formula **CH_3OH**

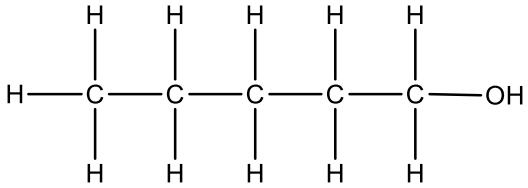
Ethanol

The structure



The condensed formula **CH_3CH_2OH**

Let us write the condensed formula of pentanol which has molecular formula of $C_5H_{11}OH$. Firstly, we need to come up with molecular structure as



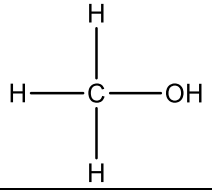
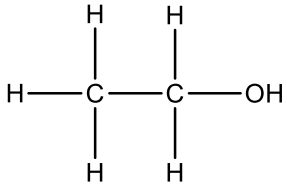
We can write the condensed formula by writing each carbon atoms with the number of hydrogen atoms bonded.

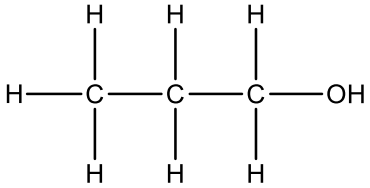
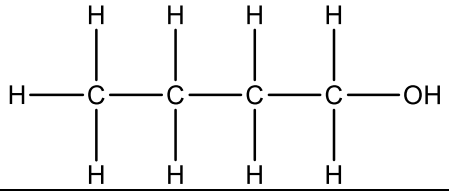
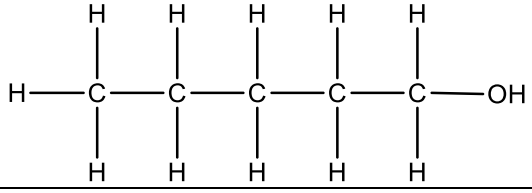
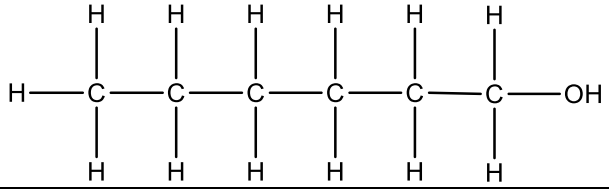
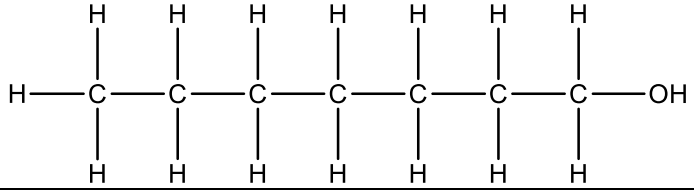
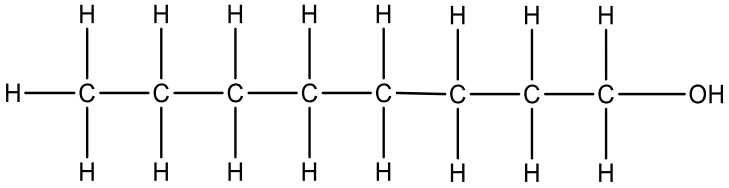
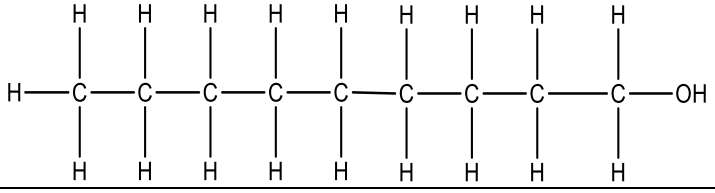
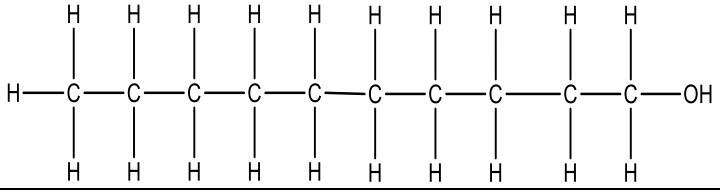
$CH_3CH_2CH_2CH_2CH_2$**OH.**

In the chain above we have four **CH2**, CH_3**$CH_2CH_2CH_2CH_2$**OH. Therefore, we can simplify the condensed formula as

CH_3**$(CH_2)_4$**OH.

Table 4.12: The formula of alkanols.

Alkanols	**Molecular formula (s)**	**Structure formula (s)**	**Condensed formula (s)**
Methanol	CH_3OH		CH3OH
Ethanol	C_2H_5OH		CH3CH2OH

Propanol	C_3H_7OH		CH3(CH2)2OH
Butanol	C_4H_9OH		CH3(CH2)3OH
Pentanol	$C_5H_{11}OH$		CH3(CH2)4OH
Hexanol	$C_6H_{13}OH$		CH3(CH2)5OH
Heptanol	$C_7H_{15}OH$		CH3(CH2)6OH
Octanol	$C_8H_{17}OH$		CH3(CH2)7OH
Nonanol	$C_9H_{19}OH$		CH3(CH2)8OH
Decanol	$C_{10}H_{21}OH$		CH3(CH2)9OH

Self-Assessment 4.1

1 a). Name the functional group of an alkanols.

b). Describe the importance of the functional group.

2 a).Write down the general molecular formula of alkanols.

b). Given the alkanols with 5 carbon atoms in the formula, name the alkanol.

(ii). Draw the molecular structure of the compound.

(iii). What is the condensed formula of the compounds.

4.13 Classification of alkanols (alcohols).

Alcohols are classified into three groups;

- Primary alcohol
- Secondary alcohol
- Tertiary alcohol

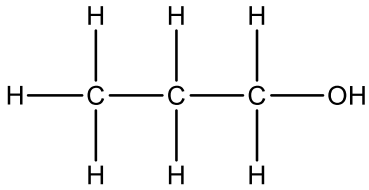
This classification depends on the number of carbon atoms bonded to carbon bearing hydroxyl group **(OH).**

Primary alcohols.

The carbon bearing **– OH** is bonded to one other carbon atom only. For instance;

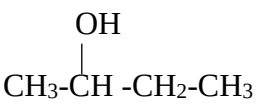
$CH_3CH_2CH_2OH$.

Structurally it can be represented as



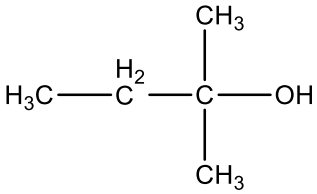
Secondary alcohols.

The carbon bearing – OH is bonded to two other carbons. For instance;



Tertiary alcohol

The carbon with – **OH** is bonded to three other carbons. For instance;



Many of the properties of alcohols are affected by whether the –OH is 1^0, 2^0 or 3^0 carbon.

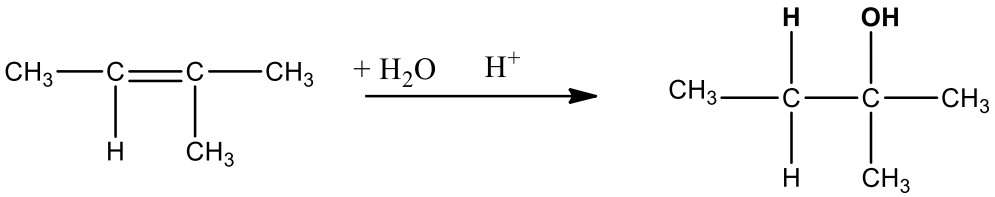
4.14 Preparation of alcohols.

Alcohols are derived from much typical method for preparation of alcohols. These are;'

- Hydration of alkenes
- Use of Grignard reagents
- Fermentation of sugars

(a) Hydration of alkenes.

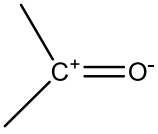
In a hydration of reaction, water is added to a double bond of alkene in the presence of sulphuric acid (H_2SO_4) as a catalyst. For example;



When water dissociate to $\mathbf{H^+}$ and $\mathbf{OH^-}$, $\mathbf{H^+}$ will add to the carbon atom which has hydrogen atom already while **OH-** will get attracted to the carbon atom with less or no hydrogen atom. This follows Markvnikovis rule of which states that *hydrogen atom should add to the carbon atom with more number of hydrogen atom and the rest of group should add to the carbon atom with few number* of hydrogen.

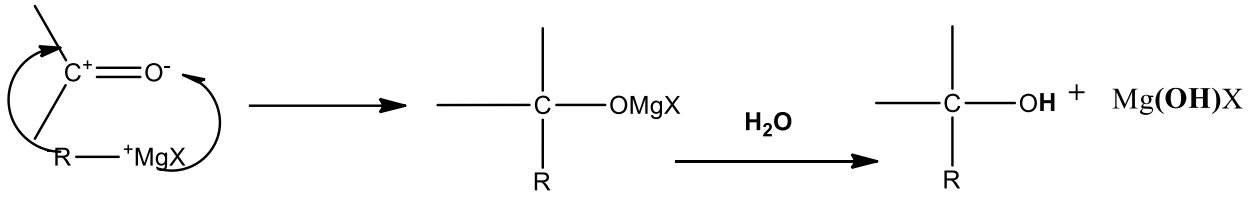
(b) Uses of Grignard reagents.

Grignard reagents are alkyl or aryl magnesium halides represented in the form of **RMgX** where **R=**alky or aryl and **X= HF, Br, I.** The organic part of this reagent is negatively charged while the inorganic part is positively charged. That is to say**, $\mathbf{R^{-\delta}\rightarrow Mg^{+\delta}}$.** The carbonyl carbon is partially positively charged due to electronegative oxygen attached to it.



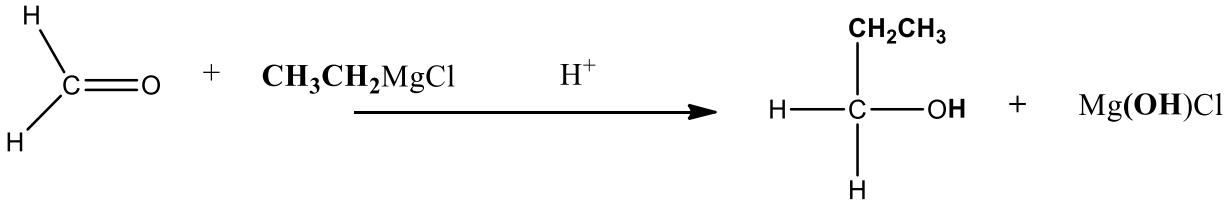
The Grignard reagents react with carbonyl compounds (**Aldehyde and Ketones**) to produce alcohols. Let's see how the reaction happens;

The negatively charged group **adds** to the **positively** charged carbon atom of the carbonyl carbon of the aldehyde or ketone. This produces an intermediate. Adding aqueous acid to the intermediate produces the alcohol. For instance;

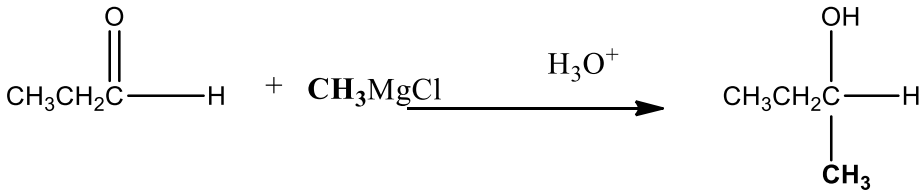


Positive seeking groups are called **nucleophiles**. The negatively charged organic group of the Grignard reagent is the nucleophile. Electron- seeking reagents are called **electrophile**. The partially positively charged carbonyl carbon is the ***electrophile***. Thus, the addition of a Grignard reagent to an aldehyde or ketone is called **nucleophile addition.** Grignard additions to carbonyl compounds are especially useful because they can be used to prepare **primary, secondary or tertiary alcohols.**

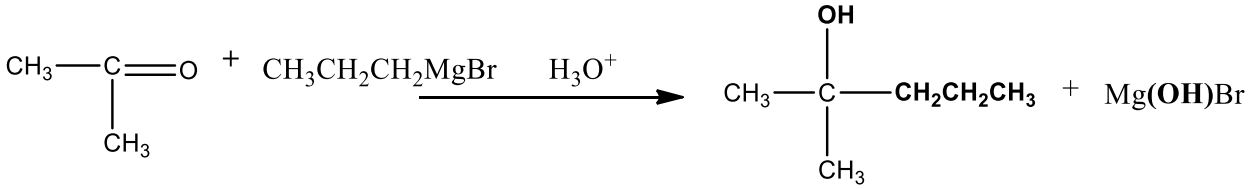
Grignard reagents react with methanal to produce primary alcohols.



Reaction of Grignard reagents with other aldehyde produces secondary alcohols



Reaction of Grignard reacts and ketones produce **tertiary alcohols**.



Formaldehyde produces 1^0 alcohols and all other aldehydes produces 2^0 alcohols. Ketones produce 3^0 alcohols.

(c) Fermentation of sugar.

This is also another methods that is used to prepare alkanols. The word fermentation *refers to the process of breaking down sugar (glucose) by micro-organisms anaerobically like yeast to produce ethanol and carbon dioxide*.

Glucose $\xrightarrow{\text{yeast}}$ ethanol + carbon dioxide

$C_6H_{12}O_6 \xrightarrow{\text{yeast}} 2C_2H_5OH_{(aq)} + 2CO_{2(g)}$.

In the reaction, yeast speeds up the fermentation process.

Preparation of indigenous beer (kachasu).

In your homes you have already seen the methods that women use to prepare alcohol. That is very good basic idea. Based on your JC knowledge about the technique used to separate the mixture, what do you call the process?

In this unit, we will build upon the knowledge of fermentation and distillation technique to prepare alkanols.

Experiment 4.1

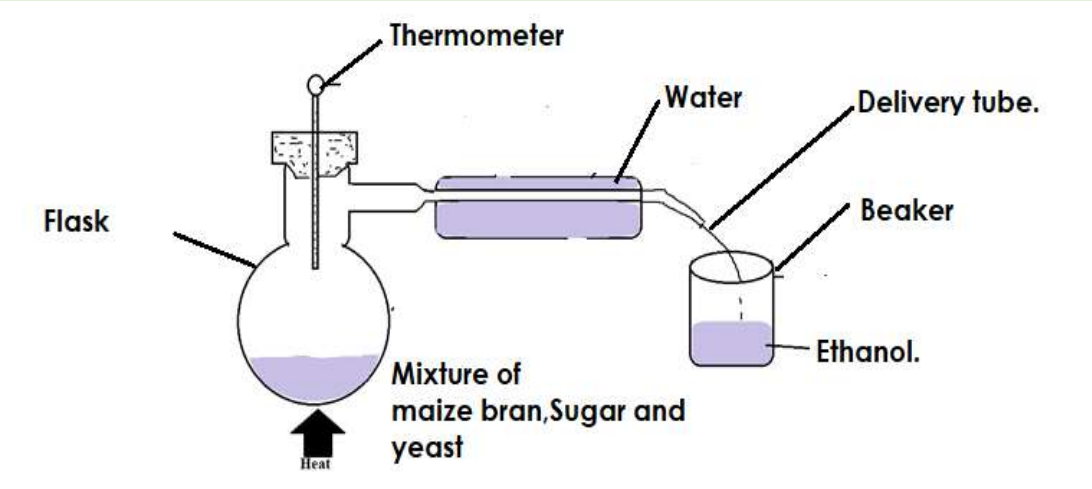
Aim: To prepare the ethnols.

Materials:

- Sugar.
- Yeast.
- Maize bran (Madeya).

Procedures.

(a) Set up the apparatus as shown below



(b) Add yeast to speed up the fermentation process

(c) Keep the mixture for 7 days to allow complete fermentation

(d) Fermentation is complete when the bubbles of gas have stopped coming out from the mixture. The bubbles or CO_2 gas.

(e) Boil the mixture (figure 4.13)

Observation

Liquid is collected in the bottle. This is ethanol

Discussion

When the mixture is heated, ethanol is going to evaporate at 78^0C, its boiling point. This is in form of vapour. When it reaches the cold water in pot it is going to condense. This is called condensation. Now, in liquid form, ethanol will move through delivery tube and get collected in the bottle. This is called distillation of ethanol.

Ethanol is collected first unlike water because it has low boiling point. Water will start evaporating at 100^oC.

The table 4.13a.Difference between local technology and modern technology of distillation.

Local technology	**Modern technology**
Cheap in terms of materials used	Very expensive
Low production of alcohol per unit time	High production of alcohol per unit time

Effects of OH on alkanols.

This affects;

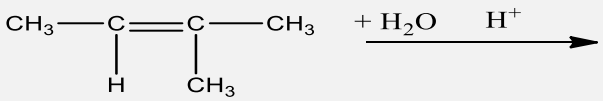
- flame colour
- Speed of reaction with sodium metal
- Solubility in water

Self-Assessment 4.2

1.a. State three classification of alkanols.

.b. Name three ways of preparing alkanols.

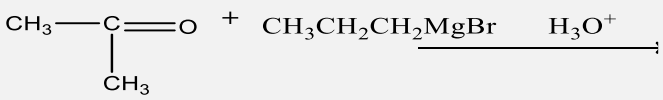
2.a. Complete the equation



b.Name the type of reaction.

C.(i). Define Nucleophiles.

(ii) Complete the equation



Br

3.a. Define “Fermentation”

b. Describe two differences between local technology and modern technology of distillation.

c. Name three effects of OH of alkanols.

4. 14 Properties of alkanols.

Just like hydrocarbons, alkanols also poses both physical and chemical properties.

Physical properties of alcohol (alkanols).

(a) Boiling point and melting points.

This increases as the size of molecules increases. As the molecules increases, the bond becomes much stronger. This requires more heat energy to be broken down.

Alcohols have higher boiling points than corresponding hydrocarbon with same number of carbon atom. This is because, alcohol form hydrogen bond between molecules. So, this is very strong bond that will need more energy to be broken down

In isomeric alcohols, 1^0 alcohols have higher boiling points than 2^0C alcohols. This is due to uneven packing of T- shape of the 2^oC alcohols which decrease the Vander Waals force.

Table 4.14a: Boiling points and melting point of alcohol.

Name	Mps (^{0}C)	Bpt (^{0}C)
Methanol	-98	65
Ethanol	-117	78
Propa-l-ol	-127	97
Propap- l – ol	-128-9^o	82
Butan-l-ol	-89.5	118

(b) They are soluble in water.

The presence of **–OH** group in alcohols makes it soluble in water. Like substances dissolve like substances. The ability to participate in hydrogen bonding enhances the solubility of alcohols in water. Small alcohols are miscible in water in all proportions. Solubility decreases as the size of alky group increases i.e. becomes more hydro carbon like

For isomeric alcohols, solubility increases in order 1^o alcohol $<2^o$ alcohol $<3^o$ alcohols. If there are many –OH in a molecule, solubility increases. For instance, hexan-i-ol has water solubility of 0,59g/100ml while hexane-l, 6-diol is miscible in all proportions.

(c) Density.

All liquid alcohols have densities of approximately 0.8g/cm^3, hence are Low dense than water.

Table 4.14b: Densities of some alcohols.

Alkanol	Density (glmℓ)	Solubility
Methanol	0.781	Soluble
Ethanol	0.789	Soluble
Propan-l-ol	0.803	Soluble
Butan –l-ol	0.810	Partly soluble
Pentan –i-ol	0.814	Party soluble

(d) They are viscous.

The viscosity increases with an increase in the size of molecules.

Chemical properties of alkanol.

(a) Reaction with Oxygen.

They react with oxygen gas (O_2) to produce carbon dioxide (CO_2), water H_2O and energy. Such energy is used as source of fuel in car. For example;

$$C_2H_5OH_{(\ell)} \longrightarrow 2CO_{2\,(g)} + 3H_2O + \text{energy}$$

Small alkanol are more reactive than the bigger one. Therefore, small alkanols produces blue flame colours meaning they have **complete combustion**.

Bigger alkanols produce yellow flame colour when burn. This is **incomplete combustion**. This is because, the proportion of oxygen atom decreases as the size of molecules increases, and however, **OH** remains the same.

(b) Reaction with metals.

They react with metals of both group 1a and 2a of the periodic table with the evolution of hydrogen gas. For instance;

$$\mathbf{R\text{-}OH + Na \rightleftharpoons R{-}ONa^{+} + H_{2(g)}}$$

This impires that

$$\mathbf{2CH_3CH_2OH + Na(s) \longrightarrow 2CH_3CH_2CH_2O^{-}Na+ + H_2(g)}$$

Ethanol + Sodium ⟶ Sodium ethoxide + hydrogen gas

The compound of the type $\mathbf{RO^{-}Na^{+}}$ are called sodium alkoxides and consists of the ions $\mathbf{R^{-}O}$ and Na+.

The $\mathbf{RO^{-}}$ **ions** are conjugate bases of alkonols and are much stronger bases than hydroxide ions $\mathbf{OH^{-}}$.

(c) Reaction with acids and bases.

They have acidic and base properties like water; hence they are called **amphoteric substance.** Alcohols acts as an acid in the presence of strong base and acts like a base in acid environment.

As Bases

Alcohols act as base in the presence of strong acid (HCl, H2So4, and HClo4). They are very weak bases.

$CH_3OH\ (l)\ Cl_{(aq)} \rightleftharpoons CH_3OH_2^+ + Cl^-$

As acids

Alcohols are extremely weak acids and will donate protons only to strongest base to form alkoxide anions **($R\text{-}O^-$).**

$CH_3OH\ (l) + NaOH_{(aq)} \rightleftharpoons CH_3ONa^- + H_2O\ (i)$

(d) Reaction with Sulphuric acids.

They react with concentrated sulphuric acids to form alkenes in hydration reaction. For instance;

$C_2H_5OH_{(aq)} \xrightarrow{\text{conce } H_2SO_4} C_2H_{4(g)} + H_2O$

(e) Reaction with carboxylic acids.

They react with carboxylic acids to form esters in the process called **esterification**. This takes place in the presence of sulphuric acid. In the process, water is eliminated between carboxylic acid and the alcohols.

This implies that;

Alkanols + carboxylic acid $\longleftrightarrow$ ester + Water

The product here is **ethyl ethanoate** which is also called **ester**.

In general

$C_2H_5OH_{(l)}\ CHOOH_{(c)} \xrightleftharpoons{H^+} CH_3COOC_2H_5 + H_2O$

Self-Assessment 4.3

State two physical properties of alkanols.

Explain why alkanols has higher melting and boiling point than hydrocarbons.

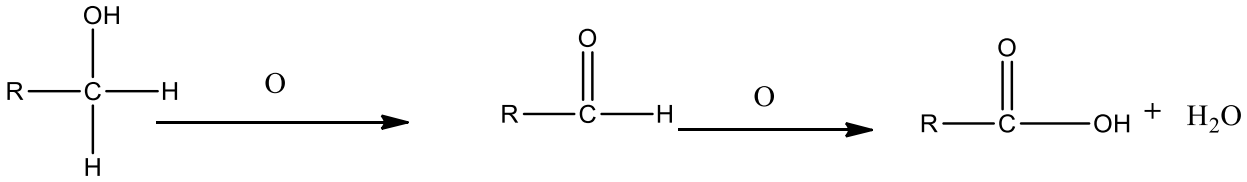
4.15 Oxidation of alcohols.

Oxidation usually manifests in one of two ways or both. In terms of oxygen, oxidation can be defined as the addition of more oxygen to molecule. In terms of hydrogen, oxidation is the loss of hydrogen atoms with their electrons from a molecule. The addition of oxygen usually involves the addition of one atom of oxygen at a time. The loss of hydrogen atoms usually involves the loss of two hydrogen atoms at one time. Two hydrogen atoms are removed from the molecule. One is from the carbon atom bearing the –OH group. The oxidation of alcohols can be accomplished by any of large number of reagents like potassium dichromate (iv) ($k_2cr_2o_7$) and potassium permanganate (vii). When acidified potassium dichromate (vi) is added to ethanol, the colour changes from orange to green. While the addition of acidified manganite's (vii) the purple colour changes to colourless.

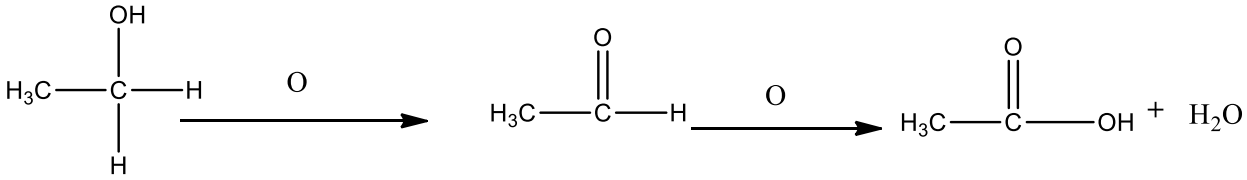
Let's look the oxidation of primary, secondary and tertiary alcohols in details.

Oxidation of primary alcohols.

Primary alcohols can be oxidized to either aldehydes or carboxylic acids depending on the reaction conditions. For instance;

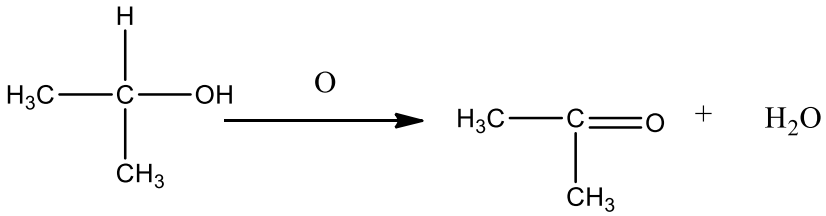


Example



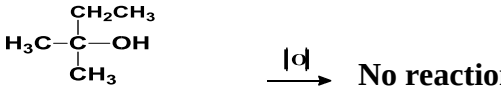
Oxidation of secondary alcohols.

Secondary alcohols are oxidized to ketones. For example;



Oxidation of tertiary alcohols.

Tertiary alcohols do not have a hydrogen atom bonded to carbon bearing the **–OH** group. Therefore, it cannot lose a hydrogen atom. Therefore, it cannot be oxidised, for instance;



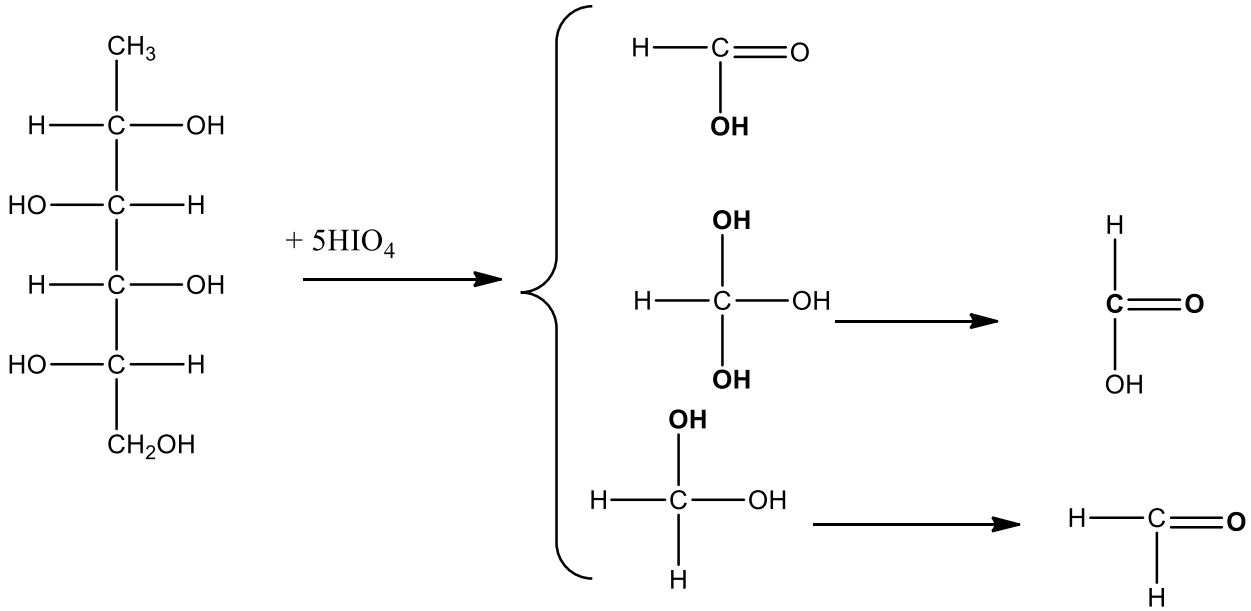
No reaction.

NB: Oxidation reaction can be used to distinguish primary and secondary alcohols from tertiary alcohols.

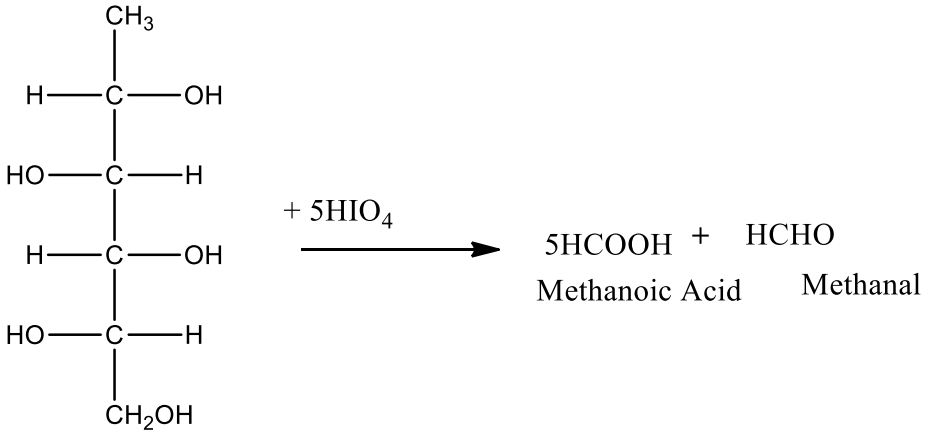
Oxidation of alcohols with many OH group on adjacent carbons.

Periodic acid **(HIO_4)** readily oxidizes alcohols with **OH** on adjacent carbons. This acid breaks the **C-C** bond if both C have **OH** group or if one has OH and the adjacent has aldehyde or ketone group. For each, **C-C** bond broken, that **C** will be oxidized once.

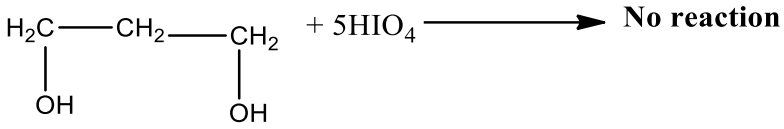
Each broken bond is placed by **OH** group. Any **C** with two **OH** group will lose water and become a carbonyl group. For instance;



This can be summarised as

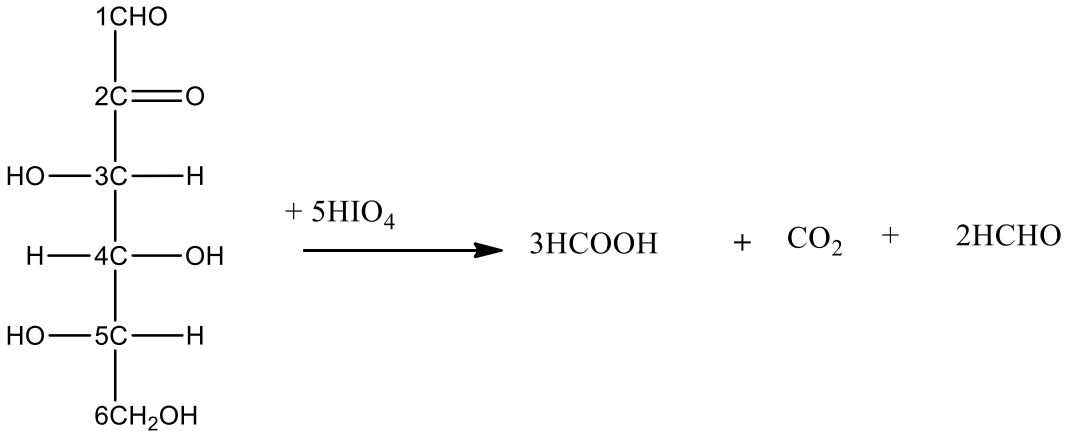


But

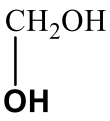


OH are not on adjacent carbon.

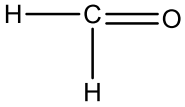
Let us consider this example;



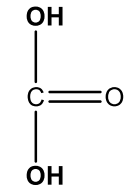
The first carbon **C, CH_2OH** which is



will lose both **OH group** and becomes carbonyl group

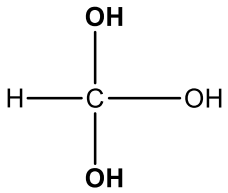


The second carbon **(C)** becomes

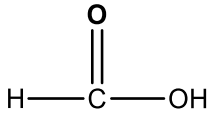


This will lose two **OH** group and becomes **O=C=O,** which is carbon dioxide **(CO_2)**

The third, fourth and fifth carbon have the structure

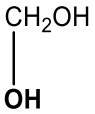


This will lose **two OH** and becomes carboxylic acid which is

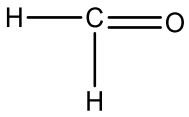


Since there are three carbon element, then they will be three carboxylic acids elements hence; **3HCOOH**

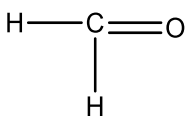
The last one, **CH_2OH** becomes



This will lose **two OH** and becomes carbonyl group which will form aldehyde.

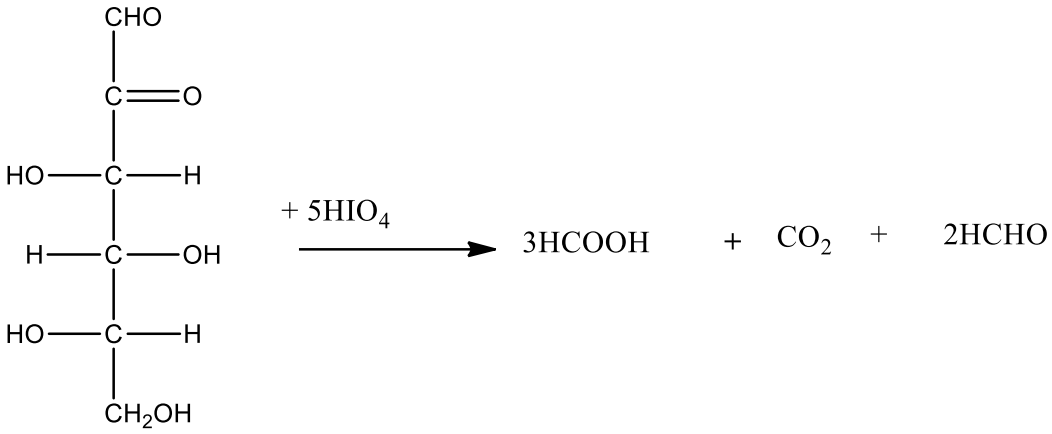


Now we have two groups of



Which **2HCHO.**

This information above can be summarized as;

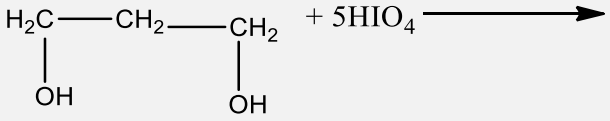


4.16 Uses of alcohols (alkanols).

- Uses as a solvent
- Used as a fuel
- As an antiseptic concentration
- Manufacturing perfumes
- Manufacturing ink, glues and paints
- Production of esters
- Ethanol is used to isolate antibodies from blood plasma
- Ethanol is used for drinking purpose
- It is used in laboratories as reagents.

Self-Assessment 4.4

Complete the equation.



State three uses of alkanols.

4.17 Solubility test.

The idea of solubility can be used to distinguish alkanols from hydrocarbon (alkanes and alkenes) based on fact that hydrocarbons are insoluble in water while alkanols are soluble in water. Consider the experiment.

Experiment 4.2

Aim: To distinguish alkanol from alkane and alkene (hydro carbon).

Materials:

- 10ml of ethanol,
- 10ml butane
- 10ml butene,
- three test tubes,
- measuring cylinder,
- distilled water
- bottle droppers

Procedures.

a) Put 5ml of ethanol in attest A, 5ml of butene in test B and 5ml of Butane in test tube C.
b) To each test add 10ml of distilled water
c) Shake gently
d) Observes and record the result in table

Table of results

Substance(s)	**No. of layers observed**
Ethanol	One layer
Butane	Two layer
Butane	Two layer

Observation.

In test tube A, one layer of liquid is observed while in B and C two layer are observed.

Conclusion:

One layer means substances have dissolved while two layer substance has not dissolved hence it is either butane which can be different y bromine test. Therefore, the dissolved substance is ethanol.

4.18 Summary

Alkanols are also called alcohols. This contains functional group called hydroxyl group, OH. The general formula for alkanols is CnH2n+1OH.

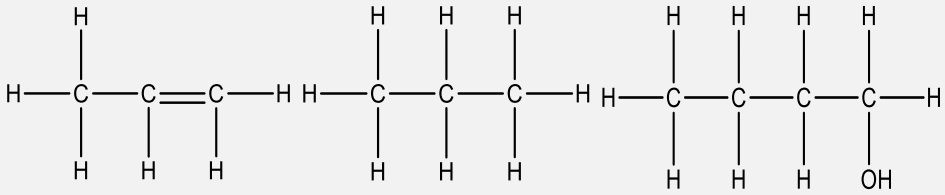
From the molecular formula of alkanols we can draw the structure and condensed formula. Alkanols are classified into primary, secondary and tertiary alkanols. The carbon atom bearing the functional group in primary alcohol is attached to two hydrogen atom. While in secondary alcohols, the carbon is bonded to one hydrogen atom. In tertiary alcohols, the carbon atom bearing the function group is bonded only to carbon atoms.

In laboratory, alkanols can be prepared by fermentation, hydrations and using Grignard reagents. Alkanols are soluble in water, viscous and melting and boiling point which increases as the size of molecule increases. They react with oxygen, metals and carboxylic acids. Alkanols also undergo oxidation reaction.

In industries and laboratories, alkanols are used in solvent, fuel , antiseptics and manufacturing perfumes.

4.19 Revision Exercise.

1. The following are the structure of organic compound. Use them to answer questions given



A **B** **C**

a) I Name organic compound
   A
   B
   C

ii. Which organic compound is soluble in water?

iii. Explain your answer of 2aii

iv. Write down the general molecular formula of the compound C

v. What is the condensed formula of compound A

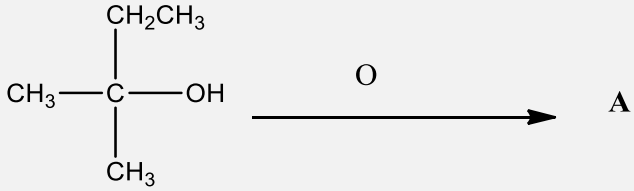
b. i Explain why alcohols have higher melting and boiling points than hydro carbon?

ii. Mention three properties of the compound C

iii. State two uses of alcohols

C. Explain how can you distinguish compound **C** from **B**.

2. The figure is a chemical equation



a. I Name the product A

ii. Explain your answer in 2ai

b. Name the following organic compound

i. CCl_4

ii. CH_3COONa

3. Study the equation below and answer the given question

$C_2H_5OH + CH_3COOH\ (aq) \xrightarrow{H^+} A + H_2O$

a. Name the process above
b. What is the product A

4. From the given equation, answer the questions below

   a. Name the process shown above
   b. Identify product B
   c. Explain why yeast is used

5. With the aid of a well labelled diagram, explain how can you separate the mixture of water and ethanol.

UNIT 5

Alkanals and Alkanones.

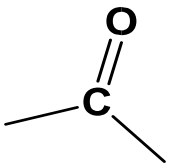
Learning Objectives.

By the end of unit, students must be able to:

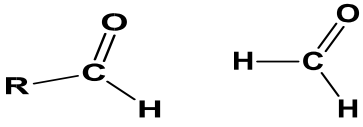
a) Identify the functional groups of alkanals and alkanones
b) Draw and name the structures of the first five alkanals and alkanone
c) Describe the sources and properties of alkanals and alkanones
d) Carry out a test to distinguish alkanals from alkanones
e) Describe the uses of alkanals and alkanones.

5.10 Introduction.

Having looked at alkanols, let look another oxy-carbon compounds called alkanals and alkanones. These two series of compounds both contains the functional group called ***carbonyl group*** **(C=O) I.e**

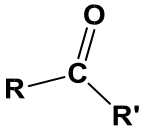


When carbonyl group is attached to *carbon atom and a hydrogen atom or two hydrogen*, the compound is called alkanal where **R** is an alkyl group.



(-CHO)

When the carbonyl group is attached to **two carbon atoms** the compound is called **alkanones.** That's is



The general formula for alkanal is $\mathbf{C_nH_{2n+1}CHO}$ where n is number 0,1,2,3 for **meth, eth, prop** etc. The number start from zero because the formula contains one carbon atom already for **methanol**.

The general formula for alkanones is $\mathbf{C_nH_{2n}O}$ where n is number of carbon 1, 2, 3 etc.

5.11 Nomenclature of alkanal.

The name of Alkanals has the ending of **–anal**. The parent name is derived from the name of alkane with the corresponding number of carbon atoms. The final-**e** in the name of alkan**e** is replaced by

–**al** in the name of alkanal. This means that Alkanals have the suffix of **–anal**. For example, **Methane** an alkane can be named as **Methanal** in alkanal.

Table 5.11 Name of alkanals.

Prefix	**Suffix**	**Complete name**
Meth-k	-anal	Methanal
Eth-	-anal	Ethanol
Prop-	-anal	Propanal
But-	-anal	Butanal
Pent-	-anal	Pentanal

5.12 Formulas of Alkanals and Alkanones.

The molecular formula of alkanals can be derived using the general formula $C_nH_{2n+1}CHO$, where **n** are numbers of carbon. For example, write molecular formula of the following compounds.

- Methanal
- Ethanal

Methanal.

Meth- means one carbon atom, then n= 0 **Why?** The formula has one carbon already.

Sub 0 for n in C_nH_{2n+1} CHO will be $C_0H_{2\,(0)}$ + 1 CHO

Solving mathematically, it will be **HCHO**

Therefore, the molecular formula of methanol is **HCHO**.

Ethanal.

Eth-1, then n=1, **why?** One carbon is already in the formula

Substitute **1** for **n** in the formula

C_1H_2 (1) +1CHO

Solving mathematically,

CH_3CHO

Table 5.12a: Molecular formula of alkanals

Alkanal	**Molecular formula**
Methanal	HCHO
Ethanal	CH_3CHO
Propanal	C_2H_5CHO

Butanal	C_3H_9CHO
Pentanal	C_4H_9CHO

Molecular structure of Alkanals.

To draw corrent structure, we follow procedures as follows;

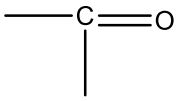
- Draw carbon skeleton, each carbon has four bonds
- Join include functional group **(C=O) or -CHO**
- Join hydrogen atom **(H)** to complete the structure

For example, draw structure of methanal and ethanal.

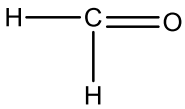
Methanal.

Molecular formula **HCHO**

Carbon skeleton with functional group atom



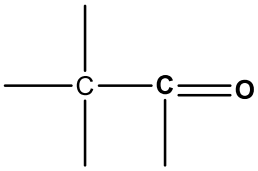
Join hydrogen atom to complete the structure



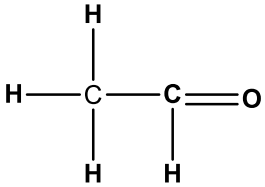
Ethanals;

Molecular formula **CH_3CHO**

Carbon skeleton with functional group



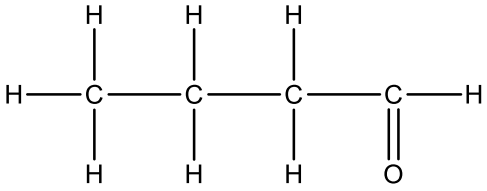
Join hydrogen atom to complete the structure



NB: **-CHO** is always at first position

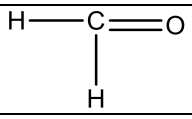
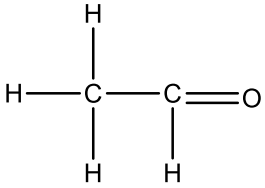
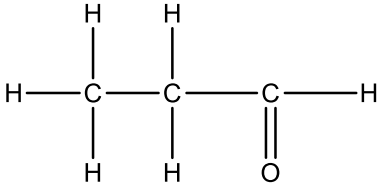
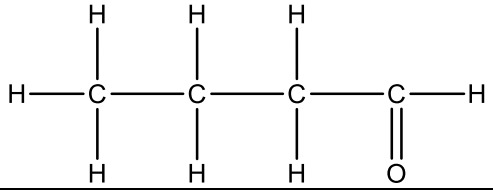
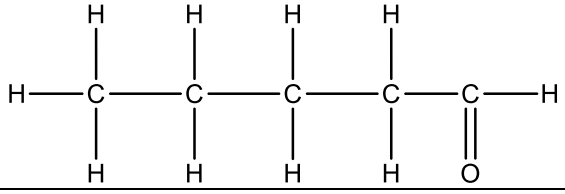
Condensed formulas of Alkanals.

To do this, we use same procedures we have been using with alkane, alkene and alkanols. For example, the condensed formula of butanal which has the structure of;



$CH_3CH_2CH_2CHO$. This can still be simplified to **$CH_3 (CH_2)_2CHO$**, since we have **two CH_2** in the structure

Table 5.12b: Formuala alkanals.

Alkanals	**Molecular formula (s)**	**Structure (s)**	**Condensed formula (s)**
Methanal	HCHO		HCHO
Ethanal	CH_3CHO		CH_3CHO
Propanal	C_2H_5CHO		CH_3CH_2CHO
Butanal	C_3H_7CHO		$CH_3(CH_2)_2$ CHO
Pentanal	C_4H_9CHO		$CH_3(CH_2)_3CHO$

Self-Assessment 5.1

1. a).Write down the functional group of alkanals.
   b).What is general formula for alkanals.
2. The alkanal has five carbon atoms in the chain.
   a).Write down each molecular formula.
   b).Name the compound.
   c).Draw the structure of the compounds.
   d).What is its condensed formula.

5.13 Nomenclature of alkanones.

The name of alkanones have the ending of **–one**. The name is derived from the name of an alkane with the same number of carbon atoms. The ending **–e** in alkane is replaced by "**-one** "in alkanones. This means that the suffix for alkanone is **–one**. For example, we can replace **–e** in methan**e** (alkan**e**) by **-one** to make **methanone**, **(alkanone).** Methanone and ethanone does not exist because they have one carbon atoms.

Table 5.13:Names of alkanones.

Prefix	**Suffix**	**Complete**
Methan-	**-one**	Methan**one** (does not exist)
Ethan-	**-one**	Ethan**one** (does not exist)
Propa an-	**-one**	Propan**one**
Butan-	**-one**	Butan**one**
Pentan-	**-one**	Pentan**one**

5.14 Molecular formula of alkanones.

This can be derived using the general formula of $\mathbf{C_nH_{2n}O}$ where n are numbers of carbon. For example; methanone. We believe that **meth-** means one carbon atom, therefore, **n=1,** sub. 1 for n in $C_nH_{2n}O$ will produce $C_1H_{2(l)}$ O.

Solving mathematically, it will be CH_2O.

Then, molecular of methanal is CH_2O

Table 5.14a: Molecular formula of alkanones

Name of alkanone	**Molecular formula**
Methanone	CH_2O does not exist
Ethanone	C_2H_4O does not exist
Propanone	C_3H_6O
Butanone	C_4H_8O
Pentanone	$C_5H_{10}O$

Molecular structure of alkanones.

the procedures are as follows;

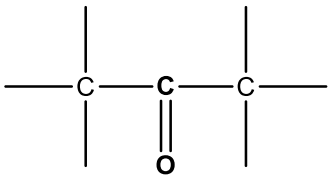
- Write down carbon skeleton. Make sure each carbon has four bonds.
- Join functional group**(C=O).** This group should be like carbon atom which is attached to two carbons.
- Join hydrogen atom **(H)** to complete the structure;

Draw structure of propan-2-one and butan-2-one.

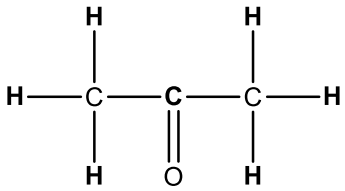
Propanone

Molecular formula; C_3H_6O

Carbon skeleton with functional group **(O)**



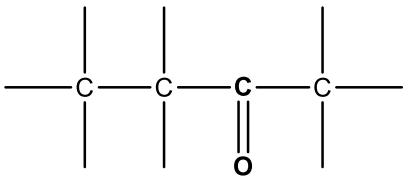
Join Hydrogen atoms to complete the structure



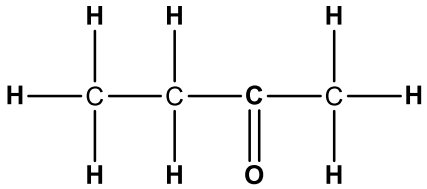
Butanone;

Molecular formula C_4H_8O

Carbon skeleton with functional group

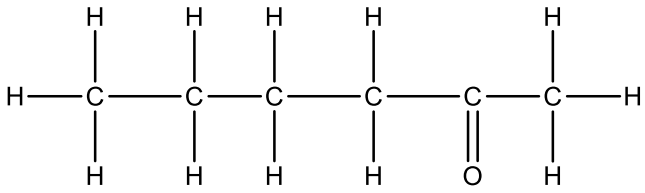


Join hydrogen atom to complete the structure



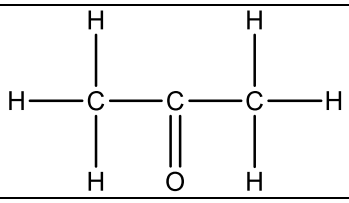
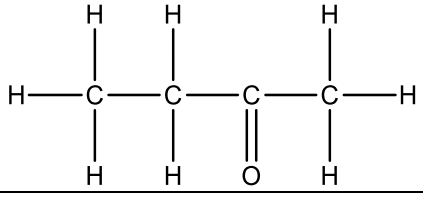
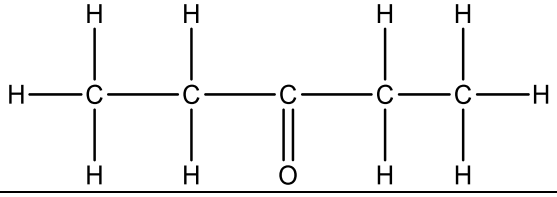
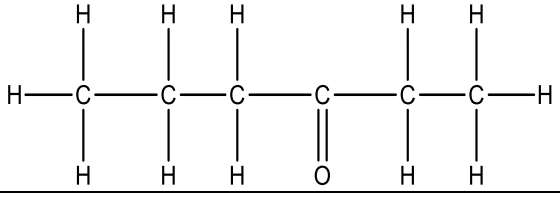
Condensed formula of alkanones.

This is written in the same way as we have been doing with hydrocarbons (alkanes and alkenes) and alkanols. For example; condensed formula of Hexan-2-one which has the structure of



This will be $CH_3CH_2CH_2CH_2COCH_3$. This can be simplified to **$CH_3(CH_2)3COCH3$.**In examination it is better to report your answer in form of **$CH_3(CH_2)_3COCH_3$** that in its simpliest form

Table 5.14b: Formula(s) of alkanones.

Alkanone	**Molecular formula**	**Structure**	**Condensed formula**
Propan-2-one	C_3H_6O		CH_3COCH_3
Butan-2-one	C_4H_8O		$CH_3CH_2COCH_3$
Penta-3-one	$C_5H_{10}O$		$CH_3CH_2COCH_2CH_3$
Hexan-3 one	$C_6H_{12}O$		$CH_3(CH_2)_2COCH_2CH_3$

Self-Assessment 5.2

a).Write down the functional group of alkanones.

b).What is general formula for alkanones.

1. The alkanal has six carbon atoms in the chain.
   a).Write down each molecular formula.
   b). Name the compound.
   c).Draw the structure of the compounds.
   d).What is its condensed formula.

5.15 Sources of alkanals and alkanones.

Sources of alkanones.

- Oxidation of secondary alcohols
- Body salts
- Sugars

Sources of Alkanals.

- Sugars
- Traces of air
- Essential oil

5.16 Properties of Alkanals and alkanones.

These compound have both physical and chemical properties

Physical Properties of Alkanals and Alkanones.

(a) Soluble in water.

The C=O group is polar, hence like substances dissolves like. Low molecular weight of propanone and ethanal are miscible in water in all proportions.

(b) Melting and boiling point.

This Increases as the size of molecules increases. **Why?** As the size of molecules increases, the hydrogen bond becomes much stronger. This will require more heat energy to be broken down in reaction to form products.

Alkanals and alkanones do not form hydrogen bond with each other. They form hydrogen bond with water.

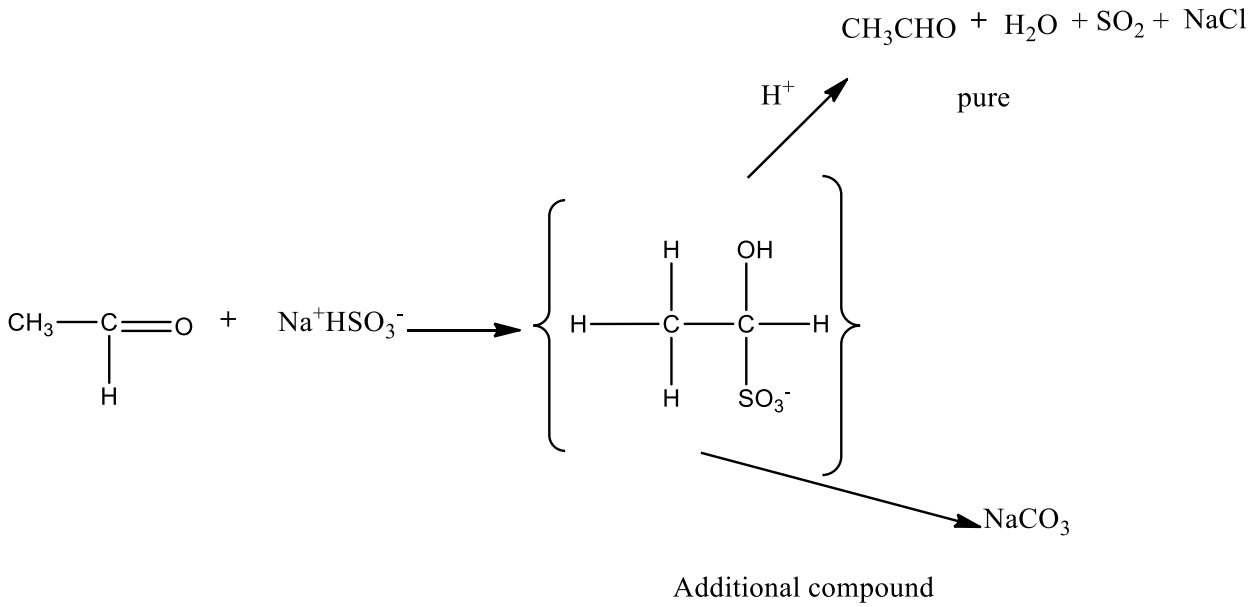
Chemical properties of Alkanals and alkanones.

Alkanals and alkanones contains a double bond (**C=O)** just like alkenes. Therefore, they also undergo **addition reaction**. However, there is different between **C=O** and **C=C.** This implies that substance(s) that add to a **C=O** are not the same as those that add to **C=C.**

*In order to appreciate the addition reactions to **C=O** group*, let us consider the polar nature of this group. The **C=O** being polar, the carbon atom is partially positively charged while the oxygen has a partial negative charge **$-C^{+\delta}=O^{-\delta}$.** Anything that is negatively charged will be attracted to the carbon. Such a "**nuclear lover** "is called a ***nucleophile***. Alkanals and alkanones undergo nucleophilic addition. Several substances add to the carbonyl carbon (**C=O).**

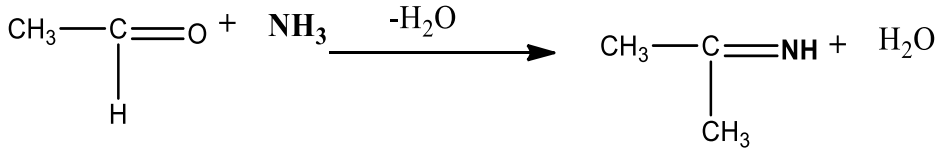
a) Reaction with sodium bisulphite ($NaHSO_4$)

Alkanals and alkanones react with saturated aqueous solution of $NaHSO_3$ to give a "bisulphite addition" compound. The product is salt and so insoluble in organic solution.It is precipitated out of solution. Once precipitated, it is filtered and on addition of acid or base a pure alkanals or alkanones is obtained.

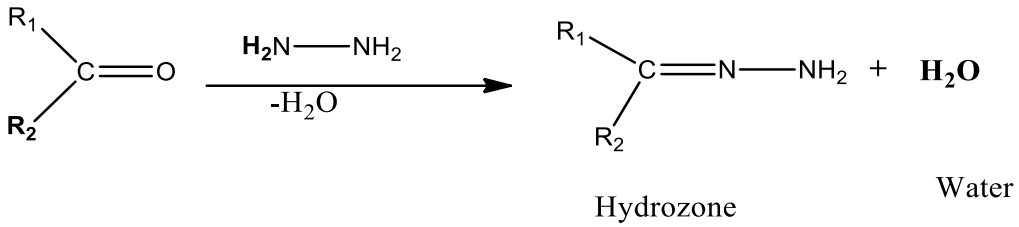


(b) Reaction with ammonia and related compound.

Ammonia **(NH_3)** add to alkanal or alkanones to form an unstable product which then loses a molecule of water. The product is called *an imine (if alkanal*) or *Schiff's base (if alkanone)* for example;



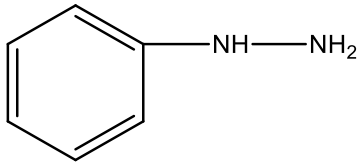
Replacing one **H** in **NH_3** with **NH_2** gives hydrazine, **(H_2N-NH_2)**. Hydrazine reacts with Alkanals and alkanones in the same way as **NH3** giving a product called (**Hydrazone).** For example;



Hydrozone

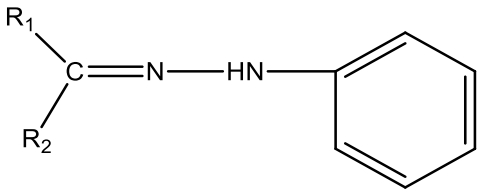
Water

Replacing one **H** in **H_2N-NH_2** by a phenyl group **(-C_6H_5)** gives *phenyl hydrazine*. **(H_2N-NHC_6H_5).** The structure of phenyl hydrazine is



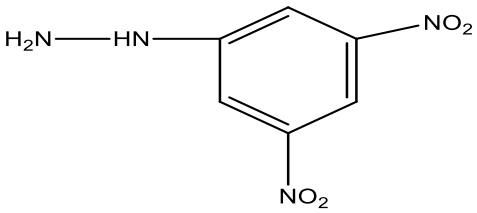
Phenylhydrazine

Phenyl hydrazine reacts with Alkanals or alkanones in the same way as hydrazine or ammonia giving products called "**phenyl hydrazone**"



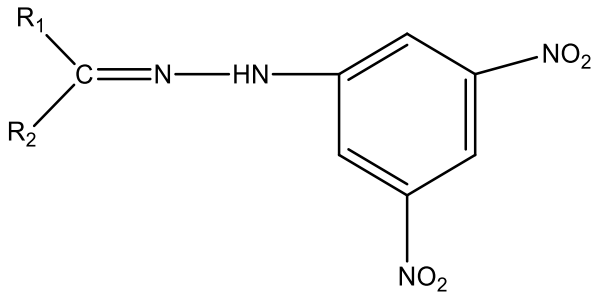
phenylhyrazone

Replacing **H** in hydrazine **(H2N-NH2)** by **a 2, 4- ditrophenyl** group gives a compound called **2, 4- dinitrophenyl hydrazine (2, 4-DNPH)**. The structure of **(2, 4-DNPH)**



2,4-dinitrophenylhydrazine

2,4-DNPH (2,4-diitrophenyl hydrazine) reacts with Alkanals and alkanones in the same way giving products (**solids**) called **2,4-dinitrophenyl hydrazon.** The structure of 2, 4-dinitrophenyl hydrazones.



2,4-dinitrophenylhydrazone

Experiment 5.1

Aim: To perfume **2,4-DNPH TEST** on alkanals and alkanones

Materials:

- 10ml ethanol,
- 10ml of propa-2-one,
- 10ml of propanal
- three test tubes
- measuring cylinder
- 15ml of 2,4 –DNPH solution.
- Stop watch

Procedures

a. Put 5ml of ethanol, propan-2-one and propanal in different test tubes labelled A, B and C represctively.
b. Add 5drops of 2, 4- dinitrophenyl hydrazine (2,4-DNPH) solution to each test tube
c. Heat the solution for 1o minutes
d. Observe the colour charge.

Observation.

Test tube A no colour change. In B and C yellow and brown colours are observed.

Conclusion.

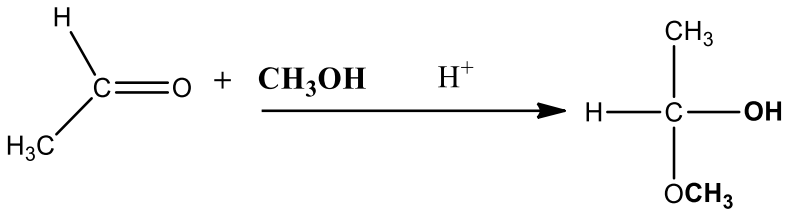
Test tube A contains ethanol. B and C are carbonyl compounds (Either aldehydes or ketone).

(c) Reaction with alcohols

Hemicental/ acetal / hemiketal/ ketal formation.

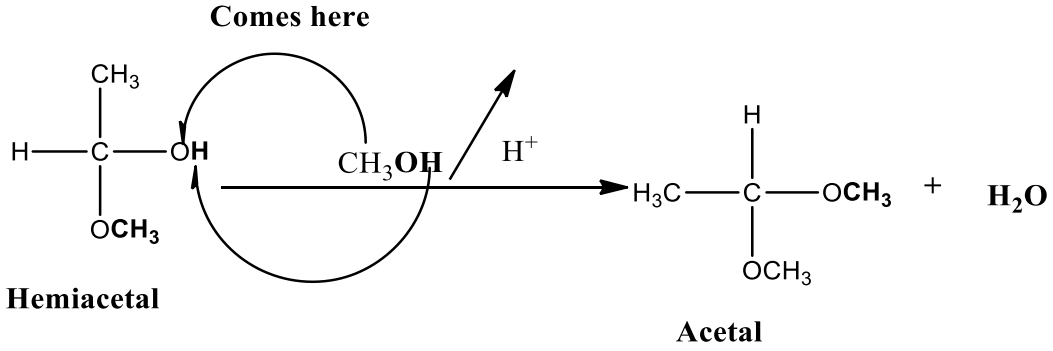
Addition of an alcohol to an aldehyde in the presence of an acid forms hemiacetal. The hemiacetal reacts with a second molar equivalent of the alcohols to produce an acetal.

Consider the reaction of **ethanal** and methanol **(CH_3OH)**



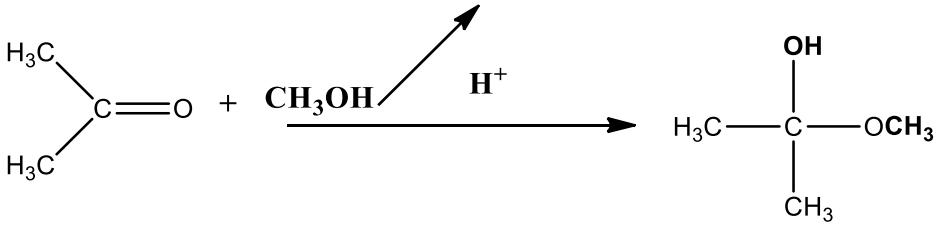
Hemiacetal

NB: In hemiacetal; **OH** and **OR** groups are bonded to the same carbon. An acetal has **two OR** groups attached to the same carbon atom.



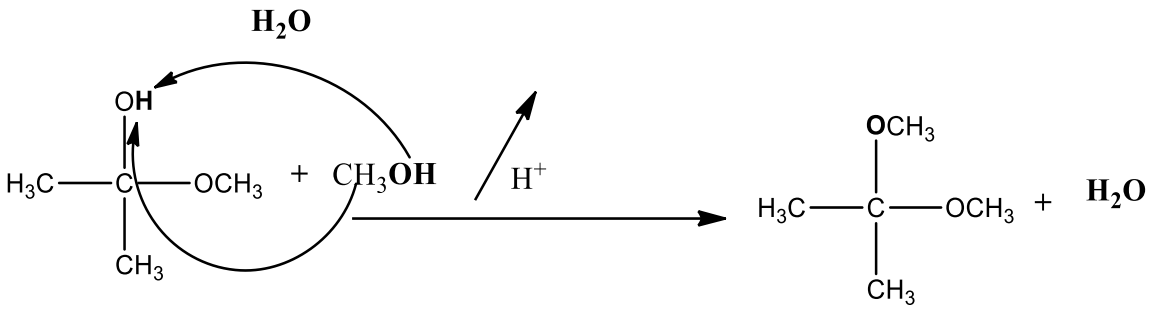
Alkanones also react with alcohols in the same way giving **hemiketals** and then **ketal.**

Let's consider the following reactions;



Hemiketal

Then, hemiacetal react with ethanol again to produce cetal.



The reaction is important in sugars which react internally to produce rings in a process called **Mutarotation.**

Self-Assessment 5.3

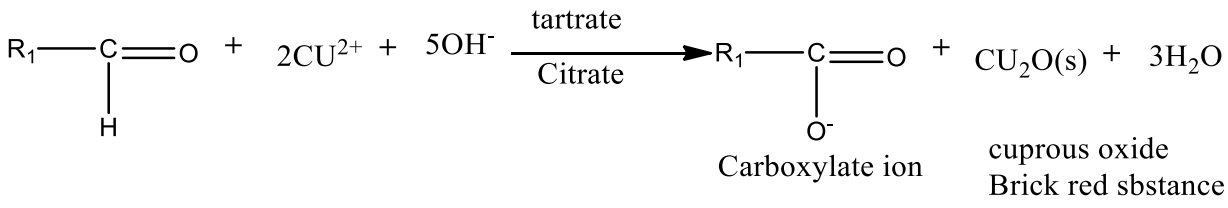
1. State two sources of the followings.
   Alkanals.
   Alkanones.
2. Name three properties of the carbonyl compounds.
3. Describe the boiling point of carbonyl compounds.

5.17 Oxidation of Alkanals and alkanones.

The reaction for **2, 4-DNPH** is positive for both alkanals and alkanones. **How do we distinguish between these two**? *Use oxidation reaction*. Alkanals are readily oxidized while alkanones are resistance to oxidation; **why?** The carbonyl carbon is not attached to any hydrogen atom. Therefore, *there is no loss of two hydrogen atoms*. Two solutions are used in the oxidation of alkanals. Alkaline solution of cupric salt (**Cu^{2+}**) and alkaline solution of silver nitrate **$AgNO_3$.**

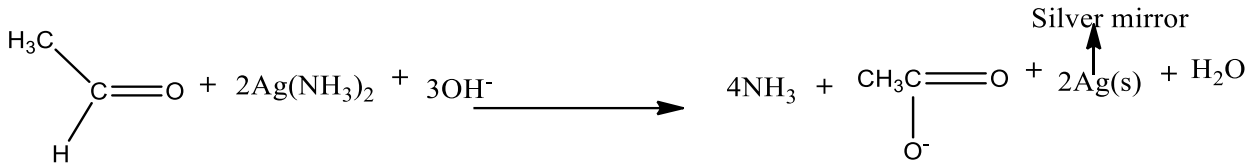
The solution of cupric **CU^{2+}** ions, two exist depending on the counter ions used. **Cu^{2+}** complexed with **tartrate** is called fehlings solution (used in chemistry) called **fehling's test**.

Cu^{2+} complexed with citrate is called **benedict solution** which is used in biochemistry to test the reducing sugars. The products are cuprous oxide (a **brick red solid**) and a carboxylate ion (conjugate base of the acid). (conjugate base of the acid)



This reaction is important in sugars. Sugars are classified as reducing or non-reducing.

Alkaline solution of silver nitrate is called **Tollens solution (Tollen's test)**. The product of the ***reaction is a mirror or grey black precipitate.***



The two reactions give positive results with alkanals and **NOT alkanones**

Important note.

a-hydroxyl ketone does react with fettling's or toile's solution. They are first converted to alkanals in alkanalic solution and it's this alkanal that react.

Fehlings test

Aim: to distinguish propanal from propan-2-one

Materials:

- 10ml of pranan-2-one
- propanal
- 15ml of fehling solution
- two test tube
- measuring cylinder.

Procedures

(a) Put $2cm^3$ of propan-2-one and propanal in test tube
(b) Label them A and B respectively
(c) Put 2 drops of fehling solution in each test tube
(d) Heat the content for five minutes.
(e) Observe the contents (mixture).

Observations.

In tube B, red precipitates are observed while in A no change.

Conclusion.

Propanal is an alkanal which has been oxidized by fehling solution. Propan-2-one is alkanone.

Tollens test.

Aim. To distinguish butanal from butanone.

Materials.

- 20ml Tollen solution.
- 50ml of butanal.
- 50ml of butanone.
- Two test tubes.
- 1001ml measuring cylinder.

Procedures.

a) Put 15ml of butanal in test tube A and 10ml of butanone in test tube B.
b) Add 5ml of Tollens solution into each test tube.
c) Heat the mixture for about ten minutes.
d) Shake gently.
e) Observe what happens.
f) Record the results in the table.

Table of results.

Substances.	**Observations.**
Butanal	
Butanone	

Observations.

Gey precipitates are observed in the test tube A while in test tube B, no colour change.

Discussion.

Tollen reagent react with alkanals to produce grey precipitates and silver mirror. After heating the mixture, the grey precipitates was observed.

5.18 Uses of alkanals and alkanones.

- They are used as solvent
- Manufacturing of perfumes
- Raw material in production of cellos acetates
- Manufacturing plastic and synthetic rubber
- Used as a denaturant for alcohols
- As reducing agent in mirror production
- As reagent in both clinical and school laboratories

Self-Assessment 5.4

1. Describe the oxidation of alkanals and alkanones.
2. State two test that can be used to distinguish alkanal and alkanones.
3. Explain why alkanones produce negative results with Tollen and fehling reagents?
4. Name three uses of alkanal and alkanones.

5.19 Summary

Alkanals and alkanones are carbonyl compounds that has functional group called carbonyl group of (C=O). This determine the properties of the compounds. when carbonyl compounds is bonded to two hydrogen atoms or one hydrogen atom, it is called alkanal. When bonded to carbon atoms only, it is called alkanones.

We have discussed that from molecular formula we can write different formulas of both families. Both compounds are soluble in water, they react with alcohols and sodium bisulphide.

We can distinguish Carbonyl compounds from alcohols using 2,4-DNPH test. However, we can distinguish Carbonyl compounds from each other by carrying out Tollen and fehling test. Tollen test react with Alkanals leaving grey precipitates and silver colour while fehling solution produce brick precipitates in reaction with alkanals. Alkanones do not react with both reagent, tollen and fehling because the carbonyl carbon is not attached to any hydrogen atom. Therefore, there is no loss of two hydrogen atoms. Remember for oxidation to take place there must be loss of hydrogen atoms.

These carbonyl compounds are used as solvent, manufacturing perfumes and plastics.

5.20 Revision Exercise.

1. (I) Name the following organic compounds
   a. C_2H_5OH
   b. C_4H_8O
   c. C_2H_5CHO

   (II). Draw the molecular structure of compound oil

   (III). Mention two physical properties of alkanal and aldehydes.

   (IV). Explain how can you distinguish the following compounds

   a From b

   c From b

2. Mention two sources of the following compounds.
   i. Alkanals
   ii. Alkanones
3. (a) Draw the molecular structure of the pentan-2-one.

   (b) Write down the condensed structure of the 3a

4. State three uses of the aldehydes and ketones.
5. (a) Explain why solubility of aldehyde and ketone increases as the size of molecular increase.

6.a) Name the families to which the following general formula belongs to,

i. C_nH_2n
ii. C_nH_{2n+1} CHO
iii. $C_nH_{2n}O$
iv. C_nH_{2n+2}
v. $C_nH_{2n+1}OH$

(b). Draw the molecular structure of the compound 5b ii which has 6 carbon atoms.

(d). Classify them as soluble and insoluble in water.

7. Mention three properties of carbonyl compounds
8. Explain why ketones do not react with Tollen's reagent.
9. Write down the general representation of aldehydes and ketones.

UNIT 6

Alkanoic Acids.

Learning Objectives.

By the end of unit, students should be able to:

a) Identify the functional group of alkanoic acids.
b) Name the first ten unbranched alkanoic acids.
c) Draw the structures of first ten unbranched alkanoic acids.
d) Write the molecular formula of alkanoic acids given the number of carbon atoms.
e) Describe the sources of alkanoic acids.
f) Describe the properties of alkanoic acids.
g) Describe the uses of alkanoic acids.

6.10 Introduction.

Alkanoic acids are also called **carboxylic acids**. They have a functional group **(-COOH),** which is called **carboxyl group**. This determines the chemical and physical properties of the compounds. The general presentation of carboxylic acids is **R=COOH,** where R is an alkyl group or hydrogen atom (H). Carboxylic acids are oxy-carbon. **Why?** The compound contains oxygen atoms.

An alkyl group is the hydrocarbon part **$CH_3CH_2CH_2$….)**. The general formula for carboxylic acid is **C_nH_{2n+1} COOH** where n are values of carbons.

6.11 Nomenclature of carboxylic acids.

This is the system of naming **carboxylic acid**. The name of these compound is derived from the name of alkanes with the corresponding number of carbon atoms. The final **– e** in the name of alkane is replaced by **–oic acid** in the carboxylic acids. For example; we can remove **e** in **methane** and put **– oic acid** to form **methanoic acid.** This means that all carboxylic acids have the same suffix **–oic acid.**

Table 6.11: Names of carboxylic acids.

Prefix	**Suffix**	**Name (s)**
Methan	**-oic acid**	Methan**oic acid**
Ethan	**-oic acid**	Ethan**oic acid**
Propan	**-oic acid**	Propaoic **acid**
Butan	**-oic acid**	Butanoic **acid**
Hexan	**-oic acid**	Hexanoic **acid**
Heptan	**-oic acid**	Heptanoic **acid**
Octan	**-oic acid**	Octanoic **acid**
Nonan	**-oic acid**	Nonanoic **acid**
Decan	**-oic acid**	Decanoic **acid**

Self-Assessment 6.1

1. What is the functional group of carboxylic acids?
2. Write down the general formula for carboxylic acids.
3. Name three carboxylic acids.

6.12 Formula of carboxylic acid.

Carboxylic acids poses both formulas like what we have already appreciated in other families of organic compounds previously.

Molecular formulas of Carboxylic acids.

This is derived by using the general formula $C_nH_{n+}COOH$, where the values of **n** starts from zero. **Why?** The formula has already one carbon atom. Remember, **meth-1, eth-2**….So to maintain these numbers of carbon we must start using values of 0, 1, 2, etc for meth-, eth- and propanoic acid respectively.

For example,

Write down the molecular formula of methanoic acid.

Since meth-1, then n=0 **why?** There is already one carbon atom in the formula.

Sub. 0 for **n** in the $C_nH_{2n+1}COOH$

$=C_0H_{2(0)} +_1 COOH$

=HCOOH
the molecular formula **is HCOOH** for methanoic acid.

For ethanoic acid; eth-2, then n=1 **why?** Same reason with methanoic acid.

Sub. 1 for n in $\mathbf{C_nH_{2n+1}COOH}$

$\mathbf{C_1H_{2(1)} +_1COOH}$

Solving mathematically it will give us;

$\mathbf{CH_3COOH}$

The molecular formula for ethanoic acid is $\mathbf{CH_3COOH.}$

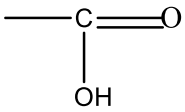
Molecular structure of carboxylic acids.

Here we use the same procedures we have been using to draw the structure of alkanes, alkenes, alkanols and carbonyl groups. For example; draw the structure of methanoic acid, ethanoic acids and propanoic acids.

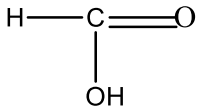
Methanoic acids.

Molecular formula**: HCOOH**

Carbon skeleton with functional group



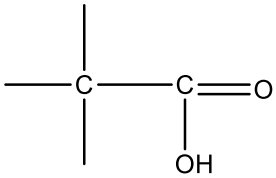
Join hydrogen atom to complete the structure



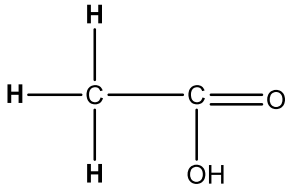
Ethanoic acid.

Molecular formula is **CH_3COOH**

Carbon skeleton with functional group



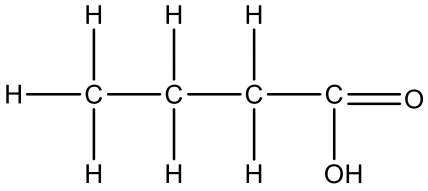
Join hydrogen atom to complete the structure



Condensed formula of alkanoic acid.

To write down the condensed formula of alkanoic acid, we apply the same theory we used in alkanols, alkanone and alkanals. This is achieved by first drawing structure and then carbon atom with its functional group or hydrogen atoms attached;

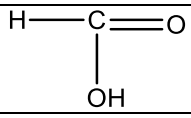
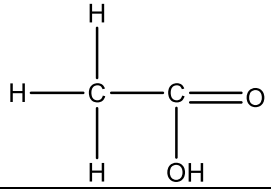
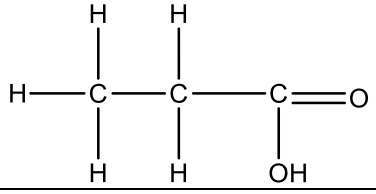
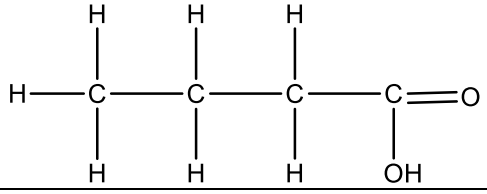
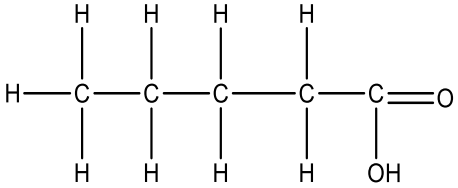
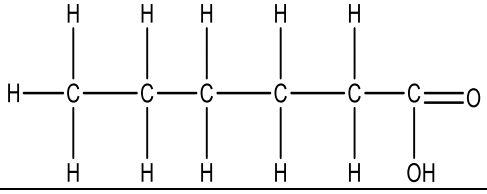
For instance, condensed formula of butanoic acid which has the structure of



This will be $CH_3CH_2CH_2COOH$. This is written as **$(CH_3)(CH_2)_2COOH$.**

Table 6.12: Names and formula of alkanoic acids.

Alkanoic acid	Value of n	Molecular formula	Structure formula	Condensed formula

Methanoic acid	0	HCOOH		HCOOH
Ethanoic acid	1	CH_3COOH		CH_3COOH
Propanoic acid	2	C_3H_7COOH		CH_3CH_2COOH
Butanoic acid	3	C_2H_5COOH		$CH_3(CH_2)_2COOH$
Pentanoic acid	4	C_3H_7COOH		$CH_3(CH_2)_3COOH$
Hexanoic acid	5	$C_5H_{11}COOH$		$CH_3(CH_2)_4COOH$

Self-Assessment 6.2

1. Given the general formula CnH2n+1COOH, Write down the molecular formula of carboxylic acid which has six carbon atom.
2. Name the compound.
3. Draw the structure of the compounds.
4. Write down its condensed formula.

6.13 Nature sources of alkanoic acids.

- Citrus fruit e.g citric acid
- Sour milk e.g lactic acid
- Vinegar e.g citric acid
- Ants, bees sting e.g methanoic acid

- Cheese e.g batanoic acid
- Grapes e.g tartaric acid
- Tea e.g tannic acid

Reasons why carboxylic acids are classified as the acids

- They change the colour of indicator from blue to red e.g. litmus paper
- They react with base in neutralization reaction

Carboxylic Acid + sodium hydroxide ⟶ sodium ethanoic + water

$CH_3COOH_{(l)} + NaOH_{(aq)} \longrightarrow CH_3COONa + H_2O$

- They have a sour taste
- They have a PH at less than 7
- They have a property or conductivity

6.14 Properties of Carboxylic acids.

Carboxylic acids have both physical and chemical properties as discussed below.

Physical properties of carboxylic acids.

(a) They are liquid at room temperature.

(b) *They are soluble in water*.

The carboxyl group **(-COOH**) is a polar. Therefore, like substance dissolve like small carboxylic acids are misable in water in all proportions. The solubility decreases as the size of molecule increases because the proportion of oxygen molecules also decreases. This reduces its polarity nature. The solubility is possible because they from hydrogen bond with water.

(c) Carboxylic acids conduct electricity.

When they dissociate, the positive ions (cations) and negative ion (**anions**) are free to move hence current can easily flow. This is the best property that we can use to distinguish alkanols from carboxylic acids in the process called **conductivity test.**

(d) The viscosity, increases as the size of molecules increase.

Increase in size of molecules increases the number of carbon atoms per unit area. Hence; the substance has more viscosity. They look like glycine oil

(e) The melting and boiling point increases as the size of molecules increases.

This is because that they have strong bond. Therefore, increases in size also increases this intermolecular forces of which will require more heat energy to be broken down to separate molec
ules.

Explain why carboxylic acids have higher melting and boiling points than alkanols with the corresponding number at carbon atom.

Carboxylic acids have stronger hydrogen bond than alkanols. This intermolecular force requires more heat energy to breakdown for the reaction to take place to form products.

Table 6.14: Comparison of melting and boiling points in alkanols and carboxylic acids.

Alkanols	**formula**	**Mt (^{o}c)**	**Bt (^{0}c)**	**Carboxylic acids**	**Formula**	**Mt(^{0}c)**	**Bt (^{0}c)**
Methanol	CH_3OH	-98	64	Methanoic acid	HCOOH	9	101
Ethanol	C_2H_5OH	-117	78	Ethanoic acid	CH_3COOH	17	118
Propanol	C_3H_7OH	-126	97	Propanoic acid	C_2H_5COOH	-21	141

Chemical properties of carboxylic acids

- They change the colour of indicators. e.g turn blue litmus paper to red

They react with alkanols to form esters.

Ethanoic acid + ethanol $\rightleftharpoons$ ***ethylethanoute + water***

$$CH_3COOH + CH_3CH_2OH \overset{H+}{\rightleftharpoons} CH_3COOHC_2H_5 + H_2O_{(l)}$$

- They react with metal higher than hydrogen in the reactivity series e.g.

Magnesium metal + ethanoic acid $\rightleftharpoons$ **magnesium ethanoate hydrogen gas**

$$Mg_{(s)} + 2CH_3COOH\ (l) \rightleftharpoons (CH_3COO)_2Mg + H_{2(g)}$$

- They partially ionize in water

$$CH_3COOH + H_2O_{(l)} \rightleftharpoons CH_3COO_{(aq)} + H_3O^+$$

They reaction is reversible. The product can react to produce reactants back.

- They react with base in neutralization reactions

$$CH_3COOH + NaOH \rightleftharpoons CH_3COONa + H_2O$$

Ethanoic + sodium $\rightleftharpoons$ **sodium + water**

Acid hydroxide ethanoate

- They react with carbonates to produce salt, carbon dioxide and water

  Calcium carbonate + ethanoic acid $\rightleftharpoons$ Calcium ethanoate + carbon dioxide+ water

$$CaCO_{3(l)} + 2CH_3COOH_{(aq)} \rightleftharpoons (CH_3COO)_2Ca + Co_2 + H_2O$$

Self-Assessment 6.3

1. Name three sources of carboxylic acids.
2. Explain why carboxylic acids are classified as acids.
3. State three properties of carboxylic acids.
4. Explain why alkanoic acids have high melting and boiling points than alkanols?

Tests for carboxylic acids.

Generally, we can use conductivity test and acids test to differentiate carboxylic acid from the other organic compounds like alkanols, alkanals and alkanones.

Acids test.

Experiment

Aim. To demonstrate the acid test.

Materials

- Two test tube
- 10ml of sodium hydroxide
- Phenolphthalein indicator
- Acetic acid

Procedures

a. Put 15 drops of sodium hydroxide (NaOH) in test tube
b. Add 2 drops of phenolphthalein indicator. This gives a pink colour
c. Add about 6 drops of acid (test liquid)
d. Shake the mixture gently
e. Observe what happens

Observation of results.

The mixture changes from pink to colourless.

If the test liquid is not acid the mixture stays pink. This test can be used to distinguish carboxylic acids and alkanols.

Conductivity test.

Experiment

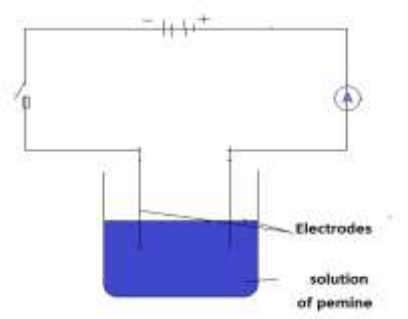
Aim: to investigate the conductivity of alkanols and carboxylic acids

Materials:

- cells, cell holder,
- ethanol (alkanol), acetic acid (carboxylic acid),
- a bulb, conducting wire,
- beaker, ammeter,
- carbon rods, switch

Procedures

a. Set apparatus as shown below(figure).



b. Put 25ml of acetic acid in a beaker.
c. Close switch.
d. Note and record reading of ammeter.
e. Repeat steps b and d with ethanol.

Table of results.

Organic compound	Current
Acetic acid	0.8
Ethanol	0

Observation of results.

Acetic acids conduct electricity while ethanol does not according to ammeter readings.

Discussion.

Acetic acid conducts electricity because it ionizes to produce ions. These ions allows current to flow while ethanol does not conduct electricity because it does not ionize.

Reaction of carboxylic acids with alkanols.

The ethanoic acid will react with ethanol in the presence of a few drops of concentrated sulphuric acid to produce ethyl ethanoate (esters) and water.

- Ethanoic Acid+ ethanol $\rightleftharpoons^{H_2SO_4}$ ethyl ethanoate + water

- $CH_3COOH\ (l) + CH_3CH_2OH \rightleftharpoons[H_2SO_4]{} CH_3COOC_2H_{5(q)} + H_2O\ (l)$

This reaction is called ***esterification***. The products of this reaction are **esters** and **water**. Sulphuric acids act as catalyst to speed up the reaction process. The first part of ester is derived from **the acid** while second part is derived from the **alcohols.** sIn writing the name, we start with alcohol part and end with acid part.

Importance (uses) of esters.

- Food flavouring
- Making perfumes
- Energy storage e.g fats and oils

Table 6.14: Differences between esterification and neutralization.

Esterification.	**Neutralisation.**
Reaction is between alkanols and alkanoic acid	Reaction is between acid and base
Products are water and ester	Production are salt and water
The reaction is catalysed by sulphuric acid	Reaction is not catalyzed
The reaction is reversible	The reaction is not reversible (irreversible)

6.15 Uses of carboxlic acids.

- Production of esters
- As solvent
- Production of salts
- Production of soap and detergents
- Manufacturing drugs e.g aspirin
- Used in protein analysis
- In textile process e.g formic acid
- Additive in bread making
- Additive in butter
- Food preservative e.g benzoic acid

Self-Assessment 6.4

1. Study the reaction given.

   $CH3COOH(l) + CH3CH2OH(l) \xrightarrow{H_2SO_4} \mathbf{A} + H2O(l).$

   (a) complete the equation.
   (b) Name the type of reaction.
   (c) State three uses of product A.
2. Describe three difference between esterification and neutralization.
3. State three uses of carboxylic acids.

6.16 Summary.

Alkanoic acids are also called carboxylic acids. The compounds have the functional group which is called carboxyl group (R-COOH) where R is hydrocarbon part. The functional group determines the properties of carboxylic acids.

The general formula for alkanoic acids is CnH2n+1COOH where n is value of carbon atom in the formual. From these formula we can deduce molecular, structure and condensed formula.

Alkanoic acids are found in vinegar, sour milk and citrus fruits. They are soluble in water, viscous and react with alkanols. The reaction of carboxylic acids and alkanols in the presence of concentrated Sulphuric acid is called esterification. This produces esters and water. Esters are used in perfumes and food flavouring.

All carboxylic acids are classified as acids because they have acidic properties like pH less than 7 and sour taste. They are weak acids.

In industries carboxylic acids are used to manufacture soap, salt, esters, drug and detergents.

6.17 Revision Exercise.

1. (a) I What is the general formula of alkanoic acid.
   ii. Using the general formula, write the carboxylic that contains 5 carbon atoms
   iii. Draw structure of carboxylic acid name in ii

   (b). Mention two sources of carboxylic acids.

2. Use the molecular formula below to answer the given questions:
   a. CH_3OH
   b. HCOOH
   c. C_2H_4
   d. C_2H_6

   (a) Name the compound that are soluble in water
   ii. Explain the reason to your answer in 2ai
   iii. Explain why compound B has high melting and boiling point than A
   (b) Explain how can you distinguish compound B and C
   (c) Write down the three properties of compound B
   (d) What is the general molecular formula of compound D

3. Use the chemical equation below to answer the question given

$$CH_3COOH_{(aq)} + C_2H_5OH_{(l)} \underset{}{\overset{H_2SO_4}{\rightleftharpoons}} \mathbf{A} + H_2O$$

(a). I Identify the process shown above

ii. Name the product A
iii. State three uses of product named in 2aii

(b) Mention two differences between reaction shown above and neutralisation.

(C) Explain why the reaction shown is called "**reversible reaction"**

4. (a) Mention three uses of carboxylic acids.

   (b) With the aid of a well labelled diagram, prove that ethanoic acids conduct electricity.

5. Mention three reasons why carboxylic acids are classified as acids.

UNIT 7

Alkanoates (Esters).

Learning Objectives.

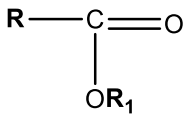
By the end of this unit, Students must be able to:

a) Identify the functional group of alkanoates (esters).
b) Name and draw the structures of alkanoates.
c) Describe the sources of alkanoates.
d) Describe the properties of alkanoates.
e) State uses of alkanoates.
f) Describe the process of soap making (saponification).

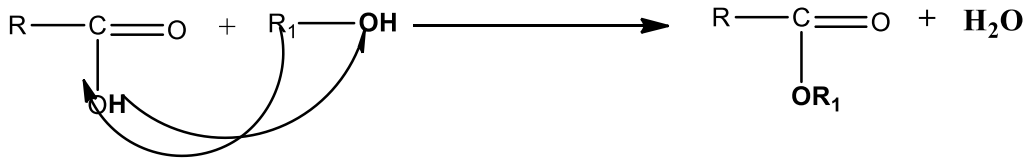
7.10 Introduction.

This is a family of organic compounds formed by the reaction of an alcohol with a carboxylic acid in the presence of concentrated sulphuric acid (H_2SO_4). This type of reaction is known as **esterification**. In esterification, ester and water are the products.

Due to loss of water molecules in the process, the reaction is also called **condensation reaction**. One of the simplest ester is ethyl ethanoate. The alkanoates contain the functional group of **RCOOR^1** where **R** and **R^1**are alkyl group (**Hydrocarbon part**). This can be presented structurally as



The general way in which alkanoates are formed.



The functional group is responsible for the properties of esters. Many esters occur in nature. In the functional group (**R COOR^1), R^1** is an **alkyl group** from alcohol while **RCOO$^-$** is called **carboxylate group** from carboxylic acids.

7.11 Nomenclature of alkanoates.

In alcohols, when the **OH** group is removed, the remaining hydrocarbon part is called **alkyl group.** The **alkyl group** is also found by removing one hydrogen from alkane. If it contains, one carbon atom it takes the **root of methane.**

When we remove the suffix- **ane** in methane and replace it with **– y1,** this makes a new compound called **methyl**. This theory applies to **alkyl group** that have more than one carbon atom.

Table: 7.11a. Alkyl group.

Formula	**Name**
CH_3-	Methyl
C_2H_5-	Ethyl
C_3H_7-	Propyl
C_4H_9-	Butlyl
C_5H_{11}-	Pentyl

In short**, alkyl group** are derived when *alkanes lose their one hydrogen atom and alkanols lose OH group.*

In carboxylic acids, they lose one hydrogen atom in functional group **(R-COOH)** to produce carboxylate ion **(R-COO^-).** The carboxylate group are also known as **alkanoates**. This means that alkanoates have the suffix of **–oate** in naming them.

The **–OH** group from alcohols (alkanols) combine with **H** from carboxylic acid (alkanoic acid) to form water **(H_2O).**

Alkanoates are represented as **$RCOOR^1$** where **$RCOO^-$** is **alkanoate** (carboxylate) and **R^1 is alkyl group.** When naming, we begin with **alky group (R^1)** from alkanols and end with **alkanoate group** (carboxylate group) **($RCOO^-$).** This forms a name called **alkyl alkanoate, where RCOO-** is derived from carboxylic acid (alkanoic acids).

For example, name this alkanoate.

$CH_3COOC_2H_5$

Firstly, divide it into two, **alkyl group** and **alkanoate group** (**$RCOOR^1$**).

CH_3COO C_2H_5

C_2H_5 is an alkyl group. Since it contains two carbon atoms, it is called **ethyl.** The **CH_3COO-** is alkanoate group. This has two carbon atoms; hence it is called **ethanoate.**

Its name follows **alkyl alkanoate,** therefore it is called **ethyl ethanoate**. This theory applies to the rest of alkanoates.

Table 7.11b Names and formulas of alkanoates.

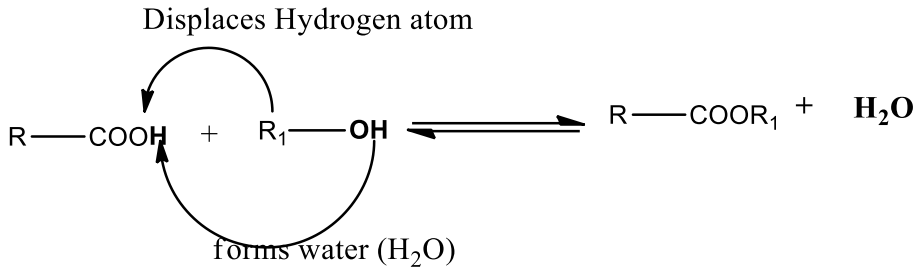
Molecular formula	**Names(S)**
$HCOOCH_3$	Methyl methanoate
$H_3CCOOCH_2CH_3$	Ethyl ethanoate
$HCOOCH_2CH_2CH_3$	Propyl methanoate
$CH_3COOCH_2CH_2CH_2CH_3$	Butyl ethanoate

$CH_3CH_2COOCH_3$	Methyl propanoate

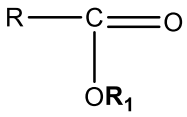
The structure of alkanoates.

Given that the structure of carboxylic acid is R-COO**H** and that of the alcohol is R^1 –**OH**. In the structure of alkanoate, the atoms in bold combine to form water **(H_2O).** The **R^1** from alcohol displaces **H (hydrogen atom)** in carboxylic acid forming R-COO**R^1.**

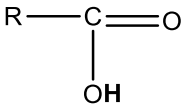
This can be summarised as



The structure of **R- C=OOR^1** can be drawn



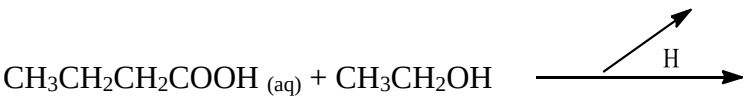
This is just like that of carboxylic acids.



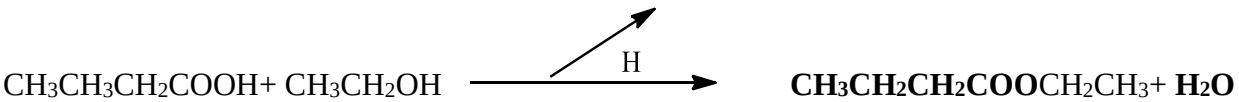
Where **H** is displaced by **R^1** from alcohol.

Example

Complete the chemical equation below and draw the structure of its product.

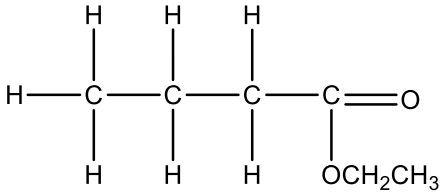


Solution



The product is **$CH_3CH_2CH_2COOCH_2CH_3$**

Structurally



Given the product (**CH_3COOCH_3**), you can easily identify the reactants. You count the carbon atom **before two oxygen to identify alkanoic acid** and count number of carbon atoms a**fter two oxygen to identify alkanol.**

For example, the alkanoate **$CH_3CH_2COOCH_3$**, has three carbon atoms before two oxygen atoms hence it is **propanoic acid a reactants** and has one carbon atom after two oxygen, and then it is **methanol**.

Therefore, **methyl ethanoate** is made from propanoic acid **(C_2H_5COOH)** and methanol **(CH_3OH).**

Self-Assessment 7.1

1. What is the functional group of esters?
2. Explain why the process of esterification is called condensation reaction?
3. Name three examples of alkanoates.
4. Draw the structures of Butyl ethanoate.
5. Name the reactants of CH_3COOCH_3

7.12 Sources of Alkanoates.

- Fruits
- Sunflower
- Animals fast
- Plants oil
- Saponification
- Making perfumes
- As a solvent

Uses of ethyl ethanoates.

Ethyl ethanoate are used as follows:

- Saponification
- Making perfumes
- As solvent

7.13 Properties of alkanoates.

The alkanoates group have both physical and chemical properties.

Physical properties of esters (Alkanoates).

- Slighting soluble in water
- Colourless
- Pleasure fruit smell
- Volatile
- Low melting and boiling point

Low melting and boiling points.

Alkanoates have low melting and boiling point than other oxy-carbon because they do not have **free-OH group**. Therefore, they cannot form hydrogen bond, this means that the molecules are held together with weak intermolecular forces of which will requires less amount of heat energy to be broken down. This constitutes to low melting and boiling points.

Chemical properties of esters (Alkanoates).

(a) Reaction with water.

They react with water to form carboxylic acids and alcohol. This process is called **hydrolysis**. For example;

Ethyl propanoate + water $\xrightarrow{H}$ **propanoic acid + ethanol**

$C_2H_5COOC_2H_{(aq)} + H_2O \xrightarrow{H} C_2H_5COOH + C_2H_5OH$

(b) Reaction with basic solution.

They react with basic solution to form metal ethanoate and alcohols. This process is called saponification. It is used in soap making whereby ester react with sodium hydroxide solution. For example,

Ethyl ethanoate + sodium hydroxide $\longrightarrow$ **Sodium ethanoate + ethanol**

$CH_3COOC_2H_{5\,(aq)}\ NaOH_{(aq)} \longrightarrow CH_3COONa + C_2H_5OH$

(c) Burns with bright flame colours.

Self-Assessment 7.2

1. Name three sources of alkanoates.
2. State three uses of ethyl ethanoate.
3. Describe two properties of alkanoates.
4. Explain why alkanoates have low melting and boiling points than other oxy-carbon?

7.15 Fats and oils.

You have already used fats and oils at your homes. How do they differ? In your further studies of chemistry called biochemistry, you will learn that lipids are classified into four major classes;Triacylglycerol's, Phosphoacylglycerols, Sphingolipids and Non-saponification (cholesterol steroid). we can find fats in animals and oils in plants.

Fats and oil belong to triacylglycerol group which has major function of energy storage. Triacylglycerol is made up of three **fatty acids (carboxylic acids)** and **one glycerol (alkanol)** as fat backbone.

Fast are usually saturated while oils are usually unsaturated. Therefore, at room temperature, they exhibit different properties.

Properties of fats at room temperature.

- Particles are closely packed together
- Solid state
- High melting and boiling point

Properties of oil at room temperature.

- Particles are not closely packed together
- Low melting and boiling point
- Liquid state

The melting and boiling point is affected by fatty acid chain length and degree of saturation. Saturated fats have no double bond while unsaturated have double bond.

The table 7.15 Fats and oils.

FATS	**OIL**
Saturated	Unsaturated
High melting and boiling point	Low melting and boiling point
Solid at room temperature	Liquid at room temperature

Self-Assessment 7.3

1. Name one source of:
   i).Fats
   ii).Oil.
2. Describe three properties of fats and oil.
3. Name the class to which fats and oil belong.
4. State three differences between fats and oil.

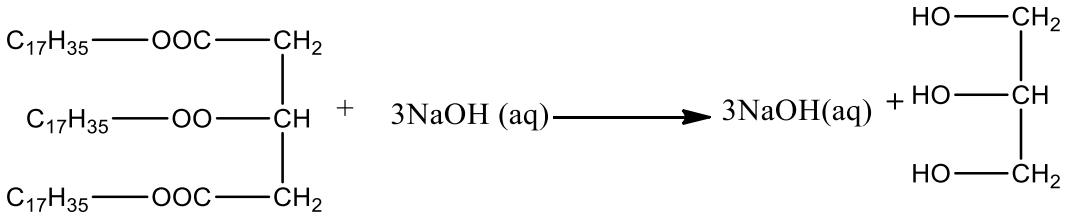
7.16 Saponification.

Soap and dertgents are used for cleaning. Millions of tonnes of soap and soapless detergents are manufactured worldwide every year. Soap is manufactured by heating natural fats and oils of either plants or animals with a strong alkali.

The product of soap is produced when esters react with sodium hydroxide solution. These fats and oil (triacylglycerol) are complicated ester molecules. Fat is boiled with aqueous sodium hydroxide (NaOH) to form soap. The esters are broken down in the presence of water-hydrolysed. This type of reaction is what we call **saponification.**

The equation below is that for the saponification of glyceryl stearate (a fat)

Glycerl stearate + sodium hydroxide ⟶ Sodium stearate + glyceral



The cleaning properties of the soap depend on **structure of soap** and **bonding of soap**. The Sodium stearate ($C_{17}H_{35}CO$-Na) consist of long hydro-carbon chain which is hydrophobic **(water hating)** and Ionic head which is hydrophilic (**water loving**).

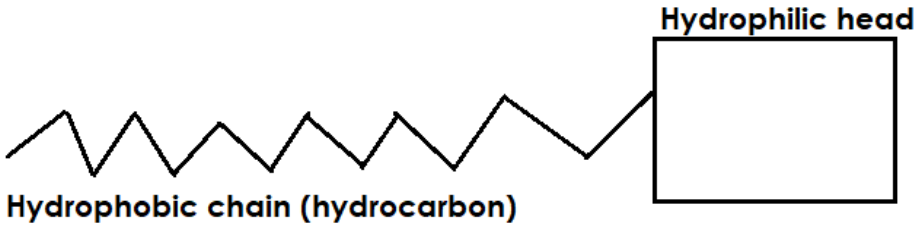


Figure: 7.16a. Simplifies the diagram of soap molecules.

Explain how soap dissolves grease?

Covalent compounds are generally insoluble in water but they are more soluble in organic solvents. Ionic compounds are generally water soluble but tend to be insoluble in organic solvents.

When a soap is put into water which has a greasy cloth in it, the hydrophobic hydrocarbon chain on each soap molecule becomes attracted to the grease and becomes embedded in it.

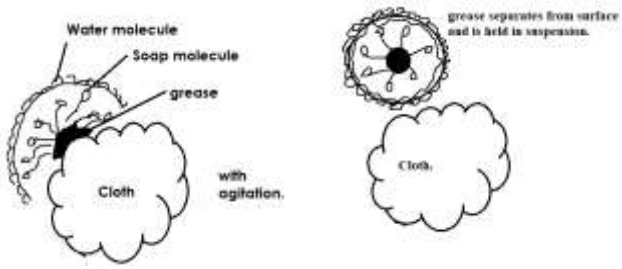


Figure 7. 16b. Soap dissolves grease.

On the other hand, the hydrophilic ionic head group is not attracted to the grease but is strongly attracted to the water molecules. When the water is stirred, the grease is slowly released and is completely surrounded by the soap molecules. The grease is therefore solubilized and removed from the cloth. The soap is able to remove grease *because of its dual nature of the soap molecule. It has a hydrophobic part (hydro carbon), covalent and a hydrophilic part (ionic head), so it will involve itself with both grease and water molecules*. **However!** Soap can form a scum with hard water by reacting the **Ca^{2+}= or Mg^{2+}** present.

Soapless detergent do not for scum with water since they do not react with Ca+ and Mg+ present in water. They are also designed that they are biodegradable.

7.17 Uses of Alkanoates.

- A solvent
- As cooking oil
- Food flavouring
- Making cloth e.g terelyne
- Making perfumes

Self-Assessment 7.4

1. Define **saponification**.
2. Describe how soap is manufactured?
3. Draw and label the structure of soap molecules.
4. State **two** cleaning properties of soap molecules.
5. Explain how soap removes grease when washing plates?
6. State **three** uses of alkanoates.

7.18 Summary.

Alkanoates are called esters which are produced when carboxylic acids react with alkanols in the presence of concentrated Sulphuric acids. This process is called esterification. The products are esters and water. Since, the process involves water loss then it is also called condensation reaction.

Alkanoates (esters) are used to manufacture cloth, perfumes, cooking oil, as a solvent and in food flavouring. The properties of alkanoates are stated as; soluble in water, fruity smell and react with water in hydrolysis reaction.

We have also looked that fats are solid while oil are liquid at room temperature. Fats are obtained from animals while oils are obtained from plants.

Soap are substances formed by saponification. In this reaction, the oil or fat (glycerol ester) is hydrolyzed by aqueous sodium hydroxide to produce the sodium salt of the fatty acid, particularly sodium stearate (from stearic acid). Soap will dissolve grease because of the dual nature of the soap molecule. It has a hydrophobic part (the ionic head) and so will involve itself with both grease and water molecule. However, it forms a scum with hard water by reacting with Ca+ (or Mg+) present.

7.19 Revision Exercise.

1. (a) Name the following compounds
   - i. $CH_3CH_2CH_2COOCH_3$
   - ii. $CH_3COOCH_2CH_2CH_2CH_3$
   - iii. $CH_3COOC_2H_5$

(b) Write down the structure of the following compounds

- i. $CH_3CH_2COOH_3C$
- ii. $CH_3COOC_2H_5$
- iii. $CH_3CH_2CH_2$

(C).i Define **"esterification"**

- ii. Name two products of esterification
- iii. Explain why sulphuric acid is used in the process of esterification?

(d). Name the reactants to which the following products have been made

- i. CH_3COOCH_3
- ii. $CH_3CH_2CH_2COOC_2H_5$
- iii. $HCOOCH_2CH_3$

1. (a) I Mention three uses of alkanoates.
   - ii. State two sources of alkanoate.

(b) Use the following molecular formulas below to answer the questions

A. CH_3COOH

B. CH_3OH

C. $CH_3COOC_2H_5$

i. Name the compound that are soluble in water

ii. Which compound has low melting and boiling point

iii. Explain your answer in 2bii

(c) I explain why saturated salts are solids at room temperature while unsaturated are liquids.

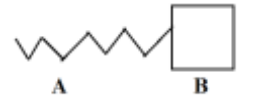
ii. Explain why fats have higher melting and boiling point than oils.

(d) I Define 'saponification"

ii. Complete the equation given below

$CH_3COOC_2H_{5\,(aq)} + NaOH_{(aq)} \longrightarrow$ **A** + ethanol

a detergent molecule may be represented by the following simplified diagram.



Name part labeled A and B.

Explain how the detergent can remove grease from dish?

Explain why soapless detergent do not form scum with water?

UNIT 8

Identifying unknown organic compound.

Learning Objectives.

By the end of this unit, students must be able to:

a) Deduce the family and structural formula of an unknown organic compound.
b) Distinguish organic compounds basing on their properties.

8.10 Introduction.

All the organic compounds we have stated; Alkanes, alkenes, alkanols, alkanals, alkanones, carboxylic acids and alkanoates can be identified using their properties. Therefore, in chemistry laboratory the technician should be aware of all properties of each compounds. If the labels fell off from the bottles containing these organic compounds, the series of experiments are conducted to identify each compound in which it is called flow diagram.

8.11 Deducing class and formulae of unknown organic compounds.

In chapter 4, we looked at empirical formula. In unit, we can build on such knowledge to deduce molecular formula of organic compounds.

Example 1

A certain alkanol contains 73.85% carbon, 13.85% hydrogen and 12.31% oxygen. Its relative molecular mass is 130. Work out the molecular formula of this alkanol and draw its structure. (RAM: C=12, H=1, O=16)

Solution

Element	Number of mole	Simplest ratio
C	$\frac{73.85}{12}$ =6.15mol	$\frac{6.15}{0.77}$ =8
H	$\frac{13.85}{1}$ =13.85mol	$\frac{13.85}{0.77}$ =17.99 =18
O	$\frac{12.31}{10}$ = 0.77	$\frac{0.77}{0.77}$ =1

The empirical formula is $C_8H_{18}O$

Molar mass (RFM) = 18+1+12+8+1+16

= 130

Number of empirical formula=$\frac{molecular\ mass}{mass\ of\ empirical\ formula}$

n=$\frac{130}{130}$

n=1

Molecular formula= **n** (empirical formula)

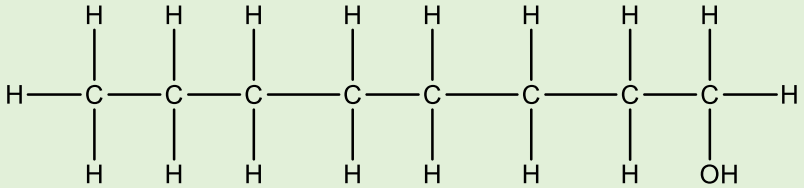
= 1 ($C_8H_{18}O$)

=$C_8H_{18}O$

In alkanol form is **$C_8H_{17}OH$**

This is octanol

The structure of the derived compound, alkanols



Example 2

An organic compound contains 58.8% carbon, 9.8% hydrogen and 31.4% oxygen. In relative molecular mass is 102. RAM: C=12, H1, O=16)

i. Write down molecular formula
ii. Identify the organic compound

Solution

Element	Number of moles	Simplest ratuo of moles
C	$\frac{58.8}{12}$ =4.9	$\frac{4.9}{1.96}$ =3
H	$\frac{9.8}{1}$ = 9.8	$\frac{9.8}{1.96}$ =5
O	$\frac{21.4}{16}$ =1.96	$\frac{1.9}{1.9}$ =1

Empirical formula =C_3H_5O

Mass empirical formula = 3C+5H +O

= 3+12+5+1+1+16

= 57

Number of empirical formula = $\frac{\text{Molecular mass}}{\text{Mass of empirical formula}}$

$n=\frac{102}{57}$

n=2

Molecular formula= n (empirical formula)

= 2(C_3H_5O)

= $C_6H_{10}O_2$

- This has a general formula $C_nH_{2n+1}COOH$
- Hence it is C_5H_9COOH
- This is hexanoic acid
- The organic compound is ***alkanoic acid***

Self-Assessment 8.1

A certain compound contains 73.85% carbon, 13.85% hydrogen and 12.13% oxygen. Its relative molecular mass is 130. Work out the molecular formula and name the family to which it belongs.

8.12 Flow diagram.

This can be used to identify unknown organic substance. In order to do this properly, one must examine the nature and structure by separating it into parts and then identify it.

In this book, we have included; bromine test, solubility test, 2,4-DNPH test, Fehling test, acidic test and conductivity test and acidic test.

Bromine test.

This is used to differentiate alkenes from alkanes. It produces positive result with alkenes by changing brown colour to colourless.

Solubility test.

This is used to differentiate oxy carbon from hydro carbons. Hydro carbons are insoluble in water; hence they form two layers when mixed with water. Oxy carbons are polar substance so they dissolve when mixed with water and they form one layer.

2, 4-DNPH.

This can differentiate carbonyl group (carbonyl compound) like aldehyde (alkanals) and ketones (alkanones) from other organic compounds.

The carbonyl compounds produce positive results with 2,4-DNPH solution by producing **yellow/ orange/ red solids (precipitates**)

Fehling's test.

Fehling solution reacts with alkanals (aldehydes) in oxidation reaction. This produces cuprous oxide (brick red precipices). Ketone (aldehydes) do not react with this solution, hence no colour changed.

Tollens test.

This also reacts with aldehyde (alkanals) producing black or grey precipitates. This is negative with ketones (alkanones).

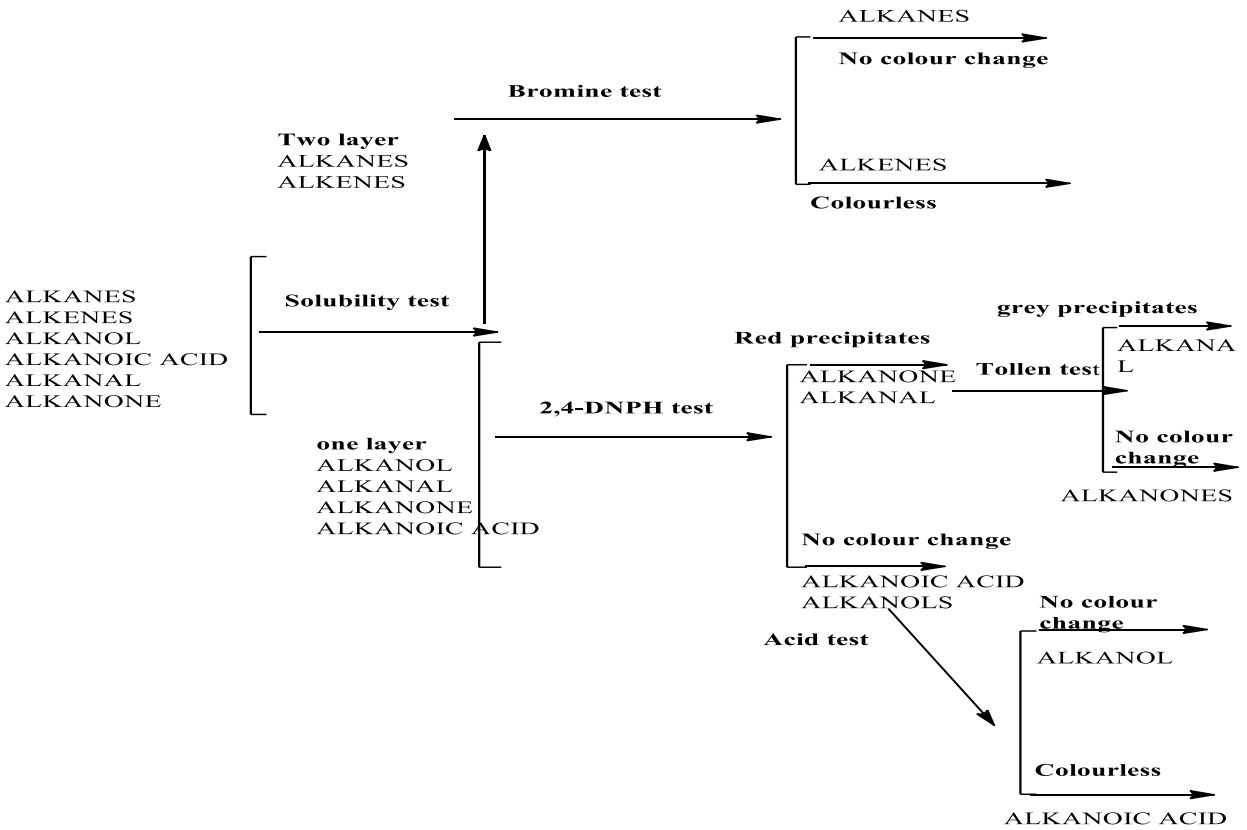
Therefore, both **Tollens test** and **Fehling's test** are used to differentiate aldehydes (alkanals) from ketones.

Acid test.

Both alkanols and carboxylic acid are polar substance hence they dissolve in water .However, we have learnt that carboxylic acid have acidic properties. As such, alkanoic acids can be differentiated from alkanols using **acid test** through neutralization reaction, where by carboxylic acid decolourise pink solution to colourless. **The minor is conductivity methods**. Alkanols do not conduct electricity.

All the test, can be summarised used the **flow diagram**. **Mind you! Procedures of all tests have been described already in all compounds.**

The key of flow diagramm used to identify unknown organic compounds.



8.14 Summary.

In this chapter we have looked at how we can deduce molecular formula and construct flow diagram. The basic studies of different test on organic compounds is very important in constructing flow diagram.

8.15 Practical question.

You are provide with dropper bottles labelled A, B, C, D and E which contains pentane, propanone, ethenal, acetic acid and pent-2-ene (not necessary in that order), burner, spatula, phenolphthalein indicator, hydrochloride acid (HCl), distilled water, sodium hydroxide (NaOH) solution, five test tubes,

Procedures

On each unknown compound perform the test in the table below and record your observations in the table below and record your observations in the appropriate space. Remember to wash the table with distilled water ester each test.

• Put 5cm3 of water to each test tube (substance) • Shake it gently • Observe	• Add 5ml of bromine solution to each substance • observe	• put 5ml of 2, 4-DNPH to each substance • heat the mixture • observe after 10 minutes	• add the mixture of 5ml NaoH and phenolphthalein indicator to each tube • shake gently • observe	• put 5cm of fehlin solution t each tub (substance • heat th mixture fo 5minutes • observe
Substance	Results	Results	Results	Results
A				
B				
C				
D				
E				

On the basis of your results, identify the unknown

A_____ B_____ C_____ D_____ E_____

Mention two sources of errors in this experiment

--

_____

_____

_____

Explain how these errors can be minimized

UNIT 9

Isomerism.

Learning objectives.

By the end of this unit, you should be able to:

- Define isomer
- Draw the structure of isomers of alkane, alkenes, alkanols, Alkanals, alkanones and alkanols acid.
- Name the isomer of alkane, alkene, alkanols, Alkanals, alkanones and alkanoic acid.
- State the effects of branching on physical properties of organic compounds.

9.10 Introduction.

Having looked at different organic compounds and how we can identify them, let us study Isomer**.** Are you familiar with this word? What is it?

The word isomer refers to *the organic compound with the same chemical formula but different arrangement of atoms*. The word of isomer is a Greek word in which **Iso-** means **same (*equal*)** and **meros** is ***part.*** The occurrence of these forms in the compound is called i**somerism**.

The study of Isomers is a great focus of many drug companies- wants to develop compounds to prevent/ control certain disease due to **isomerism.** This is based on the facts that isomers are ***life sources*** while other isomers are **poisonous.** This means that the compound can be used as drug to heal but same compound with different arrangement of atoms can be used as a **poison**. Therefore, isomers of the same compound behave different due to different arrangement of atoms.

In this unit, you are going to draw and name the structure isomers of alkane, alkenes, alkanols, alkanones, alkanals and carboxylic acids.

Types of isomers.

Generally, we have two different types of isomers which are;

- **Chain or branched isomers**
- **Positional isomers**

However, in your further studies, you will study different types of isomers.

(a) Chain/Branched isomers.

In this type, the position of substituents (**groups of atom e.g *alkyl group*, attached to longest continuous chain of carbon atoms**) are shifted to different carbon atoms within the same compound. The position of **functional group** is maintained.

(b) Positional isomers.

The position of functional group in this type of isomer, is shifted to different position of the same carbon molecules. Remember the **functional groups;** Hydroxyl group **(-OH**), carbonyl group (**C=O**), carbon to carbon double bond in alkene (**C=C**), and carboxyl group **(-COOH).** In this kind of isomer, we mainly look at position of these functional groups.

9.11 Nomenclature of the branched isomers.

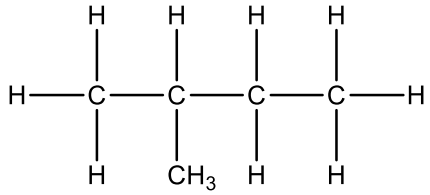
In naming isomers, we follow International Union of Pure and Applied Chemistry (**IUPAC)** rules. This is called **systematic naming.**

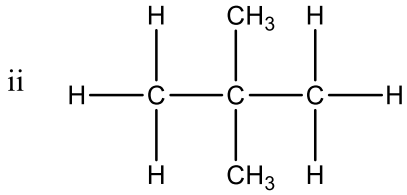
- Choose the longest continuous chain in the compound which is taken as the parent compound (or **base compound**) and given its name as appropriate (i.e., whether alkane or alkene, etc)
- Number the parent chain starting from the end nearest to the **functional group** so that *the position of the group should have the lowest possible number*. In substituted alkanes, numbering starts from near **the substituents**.
- The functional group is located by the number of the carbon atom to which it is attached and this number is placed either before or inserted in the parent name before the ending
- The name of the substituent is written before the parent name, preceded by the number of the carbon atom to which it is attached. Names of the substituent are written in alphabetical order. Every substituent group must have its own number. Numbers are separated from each other by comma and from word by hyphen
- If two or more substituents are similar, they are indicated by prefixes, di (for two), tri (for three), tetra (for four), penta (for five), etc, Each substituent must have a number locating its position on the carbon atom to which it attached
- In the alkenes or alkynes, since the double or triple bond is between two carbon atoms, the lower number is used to locate the position of the bond.
- In the cyclic hydrocarbons, the prefix cyclo- is added to the parent name and the rules above then apply. Numbering starts from one carbon atom and proceeds round the ring in such a way that the numbers locating the substituents are as small as possible

9.12 Isomers of alkanes.

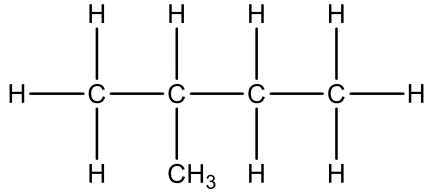
We have already looked at the alkanes. In this unit, we are only interested to look into their isomers. To do this successfully, let us apply the rules given.

Example 1

i 

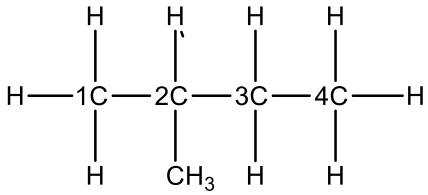
ii 

Solution 1

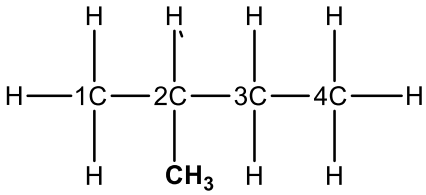
i 

i. Identify longest chain of carbon atoms to give its parent name.
   - *The longest chain has four carbons; therefore, it is **butane**.*

ii. Number the chain starting near the side group (**substituents).**



iii. Identify substituents **(side or branched**) groups

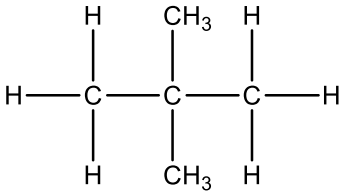


The side group is **(CH_3).** This is **methyl** and it is only one at ***positon 2***.

iv. Then, to give the full name; start with **position of side group (substituents)-side group (substituents) and finally with parent name (Root)**

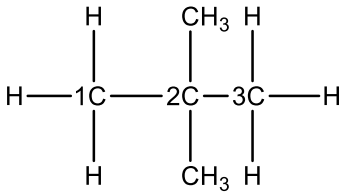
Hence; ***2-methyl butane***

Solution 2

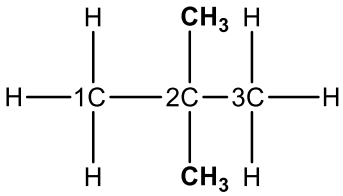
ii 

i. The longest carbon chain has 3 carbons; hence it is **propane.**

ii. Number the longest carbon chain.



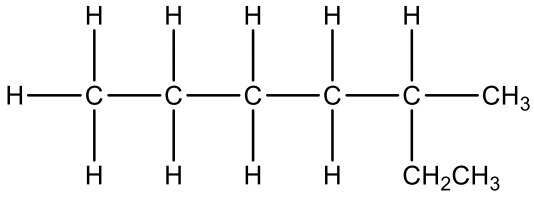
iii. Identify the Substituents **(side groups)**.



The substituents are **two CH_3** on position 2**. CH_3** is (methyl group). Since they are two, then it will be **dimethyl group.**

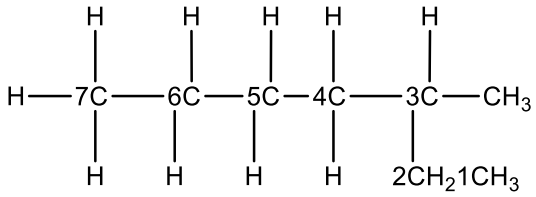
Therefore, the name is ***2, 2-dimethyl propane.***

Solution 3.

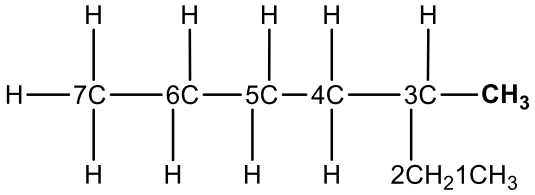


i. Longest chain has seven carbon atoms, it is called **heptane.**

ii. Number the longest chain of carbon atoms.



Substituents (side group) is **CH_3** which is **(methyl).**



Since **CH_3** is only one at position 3, the name will be.

3- Methyl heptane.

Drawing the structure isomer of alkane.

To do this successfully, you must understand the name in full. Here are some useful procedures;

- Identify the root of the name of isomers
- From the root, draw the carbon chain
- Identify ***the substituents (side group or branched) and their positions in a root (parent name)***
- Having knowing their **position** and **name**, attach them to the bond of carbon chain.
- Join hydrogen atoms, to successfully complete the structure.

Example 2.

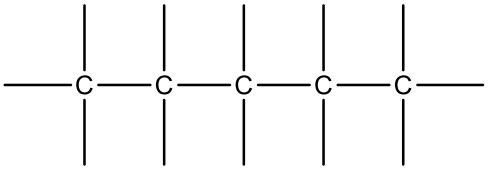
Draw the structure isomers, which have the following name;

i. 2- methyl pentane
ii. 2, 3- dimethyl butane

Solution i

2- Methyl pentane

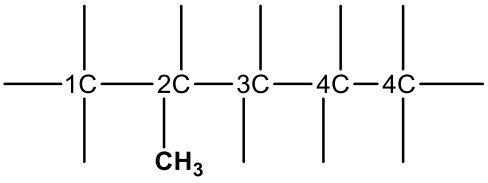
The root name is pentane; this means it has 5- carbon atoms in the longest chain.



The side group (**substituents**) is methyl group **(CH_3)** which is at position 2

Number the carbon chain and attach this at position 2.

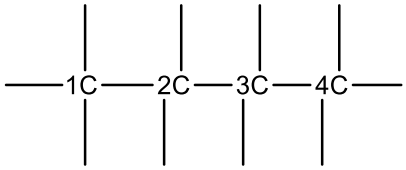
The structure of isomer is



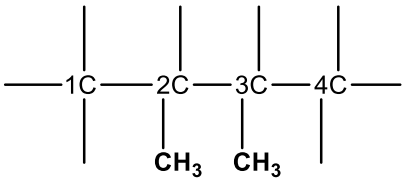
Solution i.

2,3-dimethyl butane

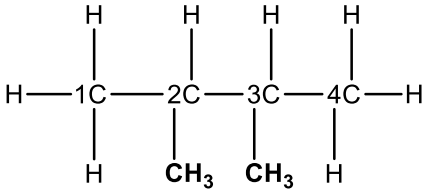
The parent name (root) is butane. Therefore, it has 4- carbon atom in its longest chain



The side group is **dimethyl** which means that it has two **methyl group**. *2, 3- dimethyl* means that one methyl group is at p**osition2**, and the other methyl group at **position 3**. Remember methyl is **(CH_3^-).** Number the carbon chain and put these **Methyl's** in their respective position.



Join hydrogen atom to complete the structure.



Example 3.

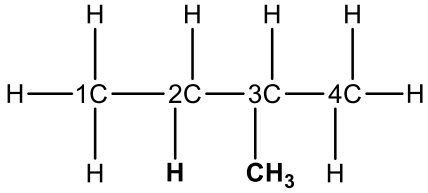
Write down and name all isomers of pentane.

Solution

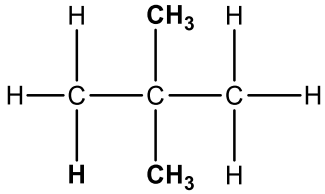
Pentane **C-C-C-C-C**

First isomer; n- pentane **C-C-C-C-C**

Second isomer.



Third isomer;

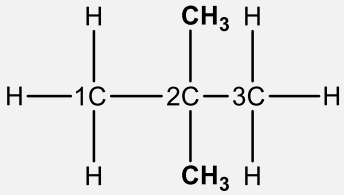


2,2-dimethyl propane.

Self-Assessment 9.1

Define Isomer.

Name the following isomers.



Draw and name all isomers of hexane.

9.13 Isomers of alkenes.

Unlike alkanes, alkenes show two types of isomerism. These are;

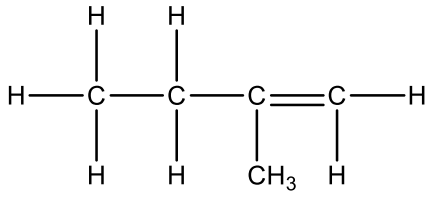
- **Chain isomerism.**
- **Position isomerism.**

In chain isomerism, we look at the **side group (branched or substituents**) attached to a long chain of carbon. We change or shift the position of **carbon- carbon double bond**, a functional group (**C=C).** In naming, we use the procedures as follows;

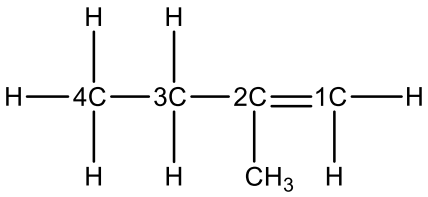
- Count up the number of carbon atoms in the parent chain
- In naming parent chain, include the position of functional group (C=C) bearing in mind that it should take smallest position.
- State the number and position of function group using carbon numbers and type of alkyl group attached ().**substituent.**

Example 1

Name the following isomers.



Number the longest chain of carbon



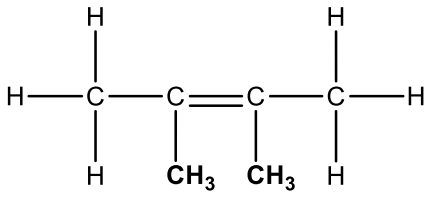
The position of functional group **(C=C) is** 1 while the substituents **(CH_3)** is at position 2 (**remember we take smallest position)**

Then functional group **(C=C),** meaning it is an alkene. Since it has 4-carbon atom, then it is **butene.**

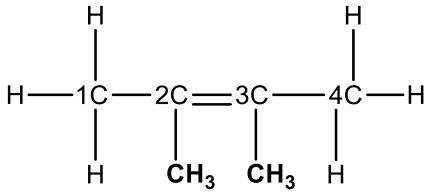
Therefore, following IUPAC rules, the name will be

2-Methylbutene

Example 1b



Number the longest chain of Carbon atoms



The position of functional group (C=C) is at position 2 while **Methyl (CH_3)** is at position 2 and 3. Therefore, following IUPAC rules the name will follow the formula stated below.

Position of side group (branched group)-Side group (branched) + parent name which includes position of functional group.

This gives.

2,3-dimethylbut-2-ene.

Drawing structure of isomers of alkenes.

Given a name, to draw the isomer you identify the following.

- The root of the name
- Position and number of double bond
- The group attached.

Draw the isomer of the following name

Example 2

i. 2- methyl prop-l- ene
ii. 2-methylbut-2-ene

Solution 1

2-methylprop-l-ene. The root is ***propene (pro-1-ene***). This means that the carbon chain will have **three carbon** atoms and **one double bond** at position 1.

C1=C2-C3

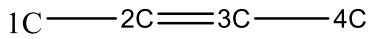
2-methyl, means that CH3 is at position **2**

C1=**C2**-C3

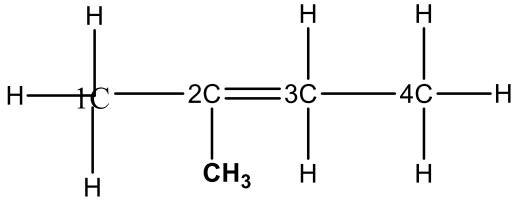
The complete structure is H-C=C-C-H

Solution 2b

3-Methylbut-2-ene. In this case, the root name is **but-2-ene.** This means that it has Four Carbon atoms and and double bond at second position of carbon chain.

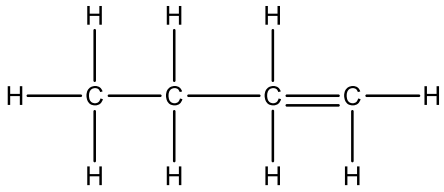


2-Methyl. This means, at position 3, the Carbon chain has **Methyl group (CH_3)**

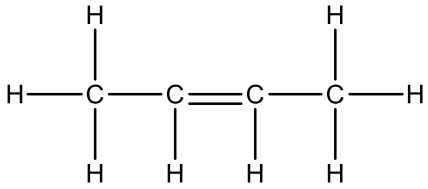


Example 3

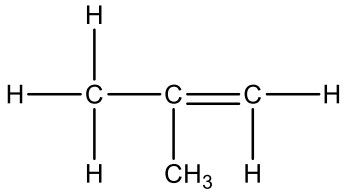
Write down and name all isomers of **Butene.**



But-1-ene



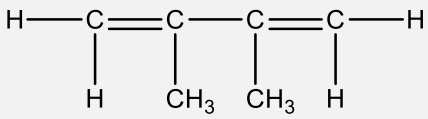
But-2-ene



2-Methylprop-1-ene.

Self-Assessment 9.2

Name the isomer.



Draw the structure isomer of 3-dimethylbutene.

Draw and name all isomers of pentene.

9.14 Isomers of alkanols.

Just like alkenes, alkanols have two isomers;

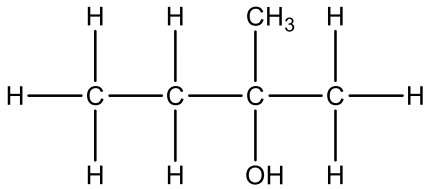
- Position isomers
- Chain isomers

In position isomers, we shift **OH** from one carbon atom to another in the same compound while in chain isomer we shift **the side group (substituents)** e.g. alky group to different position of carbon atoms in the same chain..

To name theses, we follow procedures as those of IUPAC rules.

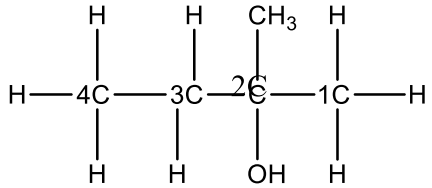
Example

Name the following isomers



Solution 1

Number the longest carbon chain. In alcohols numbering should always start near **OH group**. This is to make sure that **functional group (OH**) should smallest position. Therefore;



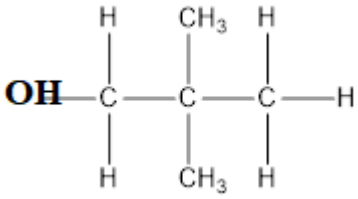
The long chain has 4 carbon atoms and **OH** group at position 2, then it is **butan-2-ol**

At the same position, it has methyl group **(2- methyl)**

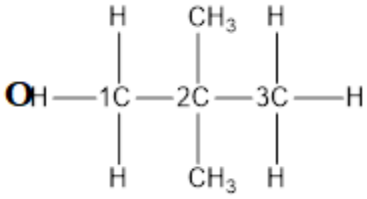
The complete name is

2-methylbutane-2-ol

Solution 2.



Number the longest chain from OH group.



At position 1, there is OH group but the longest chain contains 3 carbon atoms. This means that the name is **propanol**. At second position, it has two methyl groups **(CH_3)** which is written **as 2,2-dimethyl**

Now the complete name is **2,2- dimethyl propanol.**

Drawing structure of isomers given a name.

Draw the structure of the following isomers

i. 3,3-dimethylbutanol
   2-dimethyl propanol

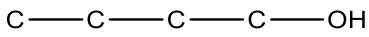
Solution 1

3,3-dimethylbutanol

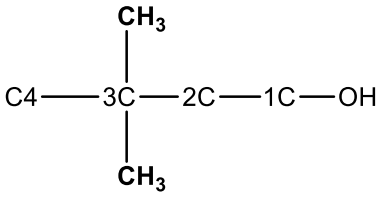
Firstly, identify the root name

Butanol.

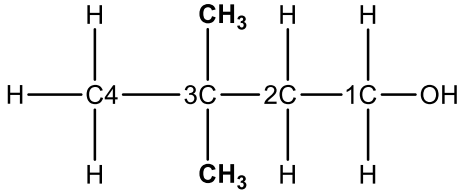
The root name has 4 carbon atoms in the long chain and **OH group** is attached to the carbon at position 1



Two methyl groups are attached the third position**.**



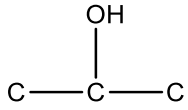
Join hydrogen atom to complete the structure



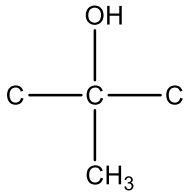
Solution 2

2-methylpropan-2-ol

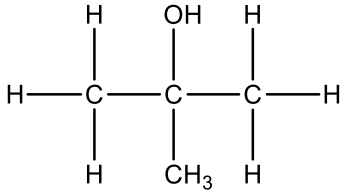
The root name is **(propan-2-ol).** This means that the longest chain has 3carbon atoms and OH group at second position



2-methyl means it has **methyl** group at same position 2.



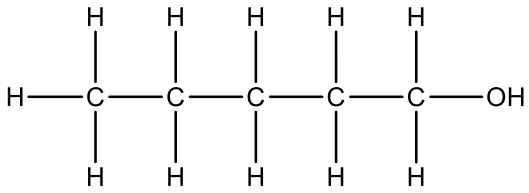
Join hydrogen bond to complete the structures.



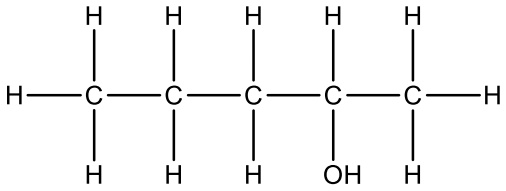
Example 3

Write down and name all isomers of pentanol.

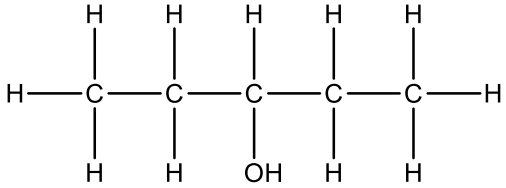
Solution



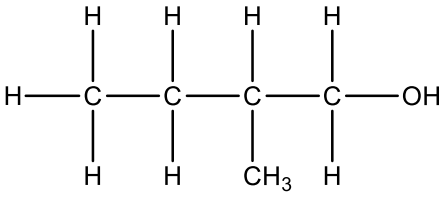
n-pentanol



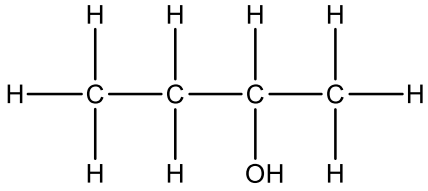
Pentan-2-ol



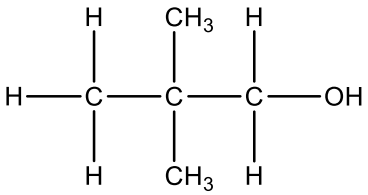
Pentan-3-ol



2-Methylbutanol



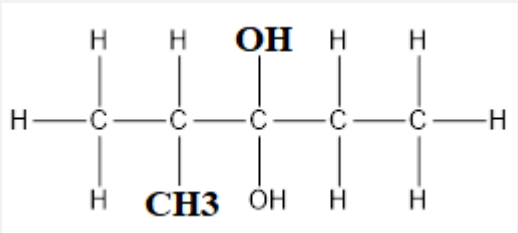
2-Methylbutan-2-ol



2,2-dimethylpropanol

Self-Assessment 9.3

1. Name the isomer below.



2. Draw the structure isomer of 3-dimethly butan-2-ol.
3. Draw and name all isomer of Heptanol.

9.15 Isomers of alkanals.

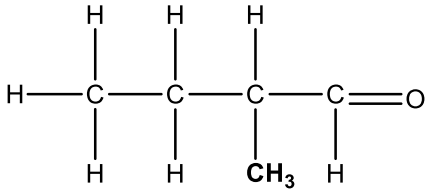
Alkanals have only chain isomers because its functional group, carbonyl group **(-CHO)** is always located at the end or position 1. Therefore, we mainly focus on the position of **side chain** within the molecule from one carbon atom to another.

When naming this compound, we use the same IUPAC rules, as follows;

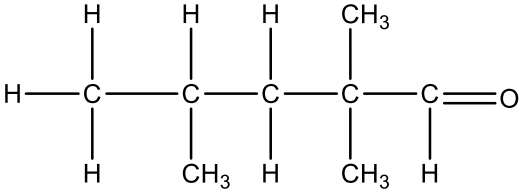
- Number the carbon chain starting from CHO (C=O) group to make sure that it takes smallest position
- Identify the carbon chain to give basic name
- Identify the position of alkyl group or any other substituent.

Example 1

Name the following isomers of Alkanals.



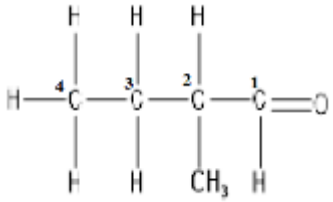
A



B

Solution 1

Number the carbon atoms starting from **CHO**



The carbon chain has **4 carbon atoms,** hence it is **butanal**. At position **2**, there is **methyl group (CH3**). This can be written as **2-methyl**.

The complete name follows**; position- substituent and rootname**

Position is **2**

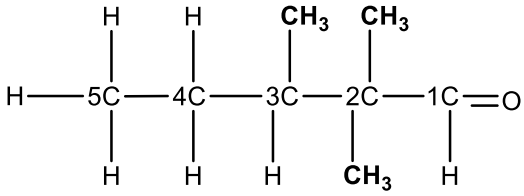
Substituent is **methyl**

Name is **butanal**

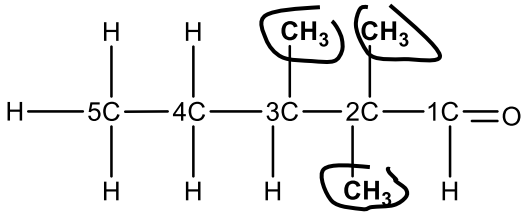
Therefore, the complete name will be: **2-methylbutanal**

Solution 2

Number the carbon skeleton starting from **CHO**



The carbon chain has **5- carbon** hence it is **pentanal.** At position 2 of carbon chain**, two methyl group** are attached and at position **3 one methyl group** is also attached



These **three methyl group (CH3)** can be written as;

2, 2, 3- trimethyl parent name

Since parent name is **Pentanal**, then the full name is **2, 2, 3- trimethyl Pentanal**

Drawing structures of isomers of alkanals.

To draw the structure of isomers, identify the **root name**, **side group** attached and **their positions**

Procedures

- After identifying root name, then draw carbon skeleton
- Indicate the position and attach functional group **(CHO**)
- Put the side group **(substituents**) in their respective groups.

Example 1

Draw the structure isomers of the following compounds.

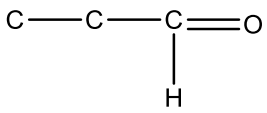
2,2- dimethyl propanal

Solution 1

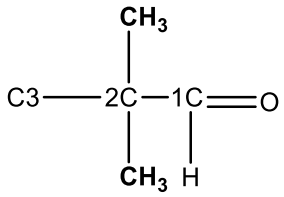
2,2- dimethyl propanal.

Remember! **Position- side group- root name**

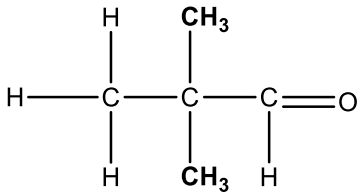
Then, here root name is **propanal**. This means, it has 3 carbon atom and –**CHO** at position 1



2, 2-dimethyl means that it has **two** methyl group **(CH3)** both attached at second position of carbon atom.

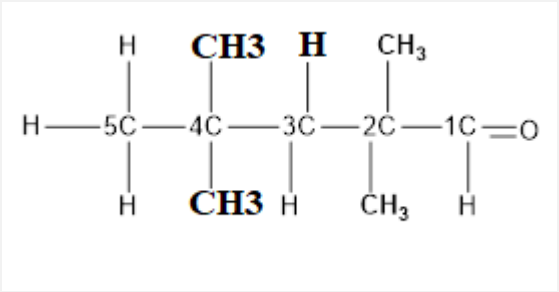


Join hydrogen atoms to complete the structure



Self-Assessment 9.4

1. Name the isomer.



2. Draw the structure isomer of 2,3-dimethylpentanal.
3. Draw and name all isomer of Pentanal.

9.16 Isomers of alkanones.

These have two isomers. Namely;

- *Position isomers*
- *Chain isomers*

In position isomers, we shift the position of carbonyl group **(C=O)** from one carbon atoms to another within the same molecules.

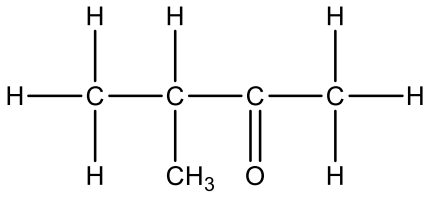
In chain isomerism, only side group (**substituents**) e.g. alkyl group like ***Methyl group*** is changed from one carbon atoms to another.

When naming this isomer, we follow procedures started below;

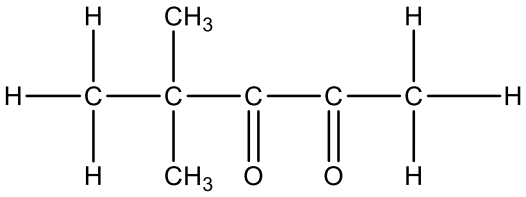
- Number the carbon skeleton (chain) to give the basic/root name. Numbering start from the functional group (C=O)
- Make sure you indicate the position of functional group or side groups attached. If more than two groups are used the word **di, tri, tetra**…… before functional group name. This also applies to **side groups**
- Then, give a full name.

Example 1

Name the following isomers;



A



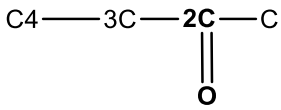
B

Solution a

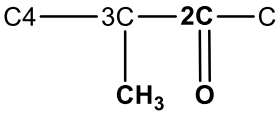
Number the carbon skeleton;

C^4-C^3-$\mathbf{C^2}$-C^1

The functional group is at position **2 (C=O)** and the chain has 4- carbon atoms. The root name will be Butan-2-one.



At the third position, **we have methyl group (CH3**) attached.

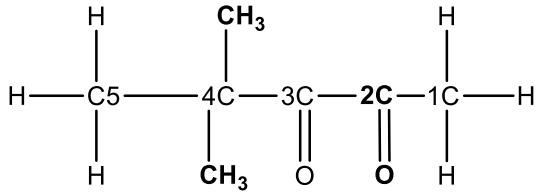


This can be written as; **3- Methyl**.

Then, the full name will be **3- methyl butan-2-one**

Solution b

Number the carbon skeleton of the structure.

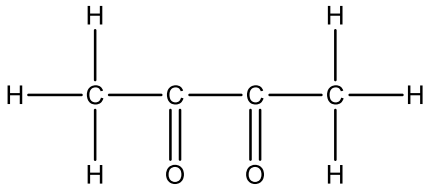


The carbon chain has **5- carbon atoms;** hence the functional groups are at **position 2 and 3.** However, the two side groups **(Methyl) are at position 4.** Since it has 5 carbon atoms, the name will be Pentanone. The full name/ correct name **is 4, 4-dimethylpentan-2,3,-dione** *why dione?*

It has **two (C=O)** groups.

Example c

,



The compound below has two **C=O** group .The isomer name will be**, Butan-2,3-dione**

Drawing structure isomers of alkanones.

Follow procedures below;

- Identify parent name. In the parent name, make sure you know the functional group and their position.
- Draw carbon skeleton which include functional group.
- Identify the **side group** and attach them to the carbon skeleton in their respective position.

Example 2

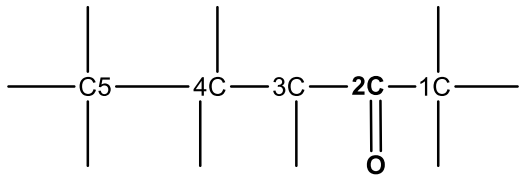
Draw the isomer of the following name.

- 4- dimethyl pentan-2-one
- Butan-2,3-dione

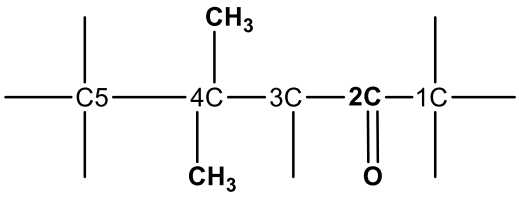
Solution 1

4- Methyl pentan-2-one

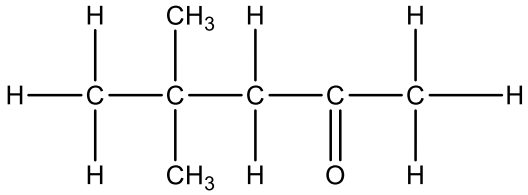
The parent name is **pentan-2-one**. This means it has 5 carbon atoms with its functional group at **(C=O) at** position 2. Then, draw the carbon skeleton using this information.



4,4-dimethyl-means it has two Methyl group, (**CH3**) group at position-**4**.



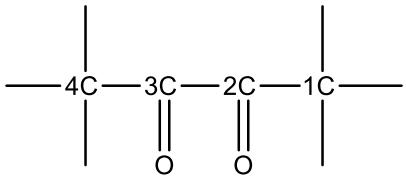
The complete structure will be



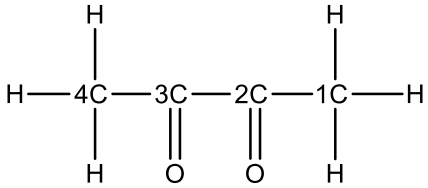
Solution 2

Butan-2,3-dione

The parent name **is Butan-2,3-dione**. This means it has 4- carbon atom and functional group at position 2 and 3.

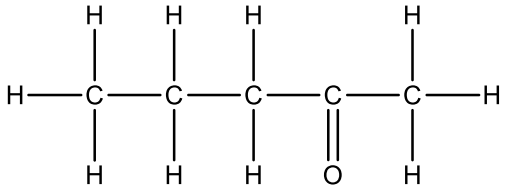


The hydrogen atoms are joined to complete the structure.

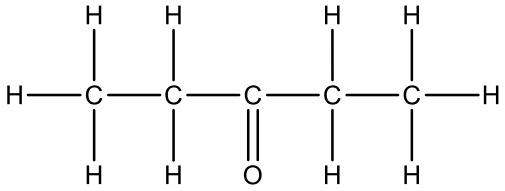


Example 3

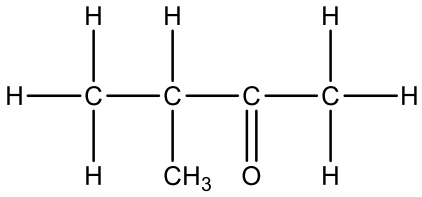
Write down and name all isomers of pentane-2-one



n-pentan-2-one



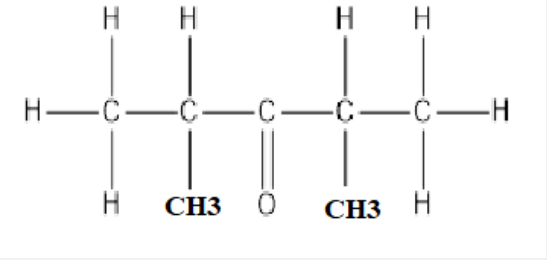
Pentan-3-one



3-Methylbutan-2-one

Self-Assessment 9.5

1. Name the isomer.



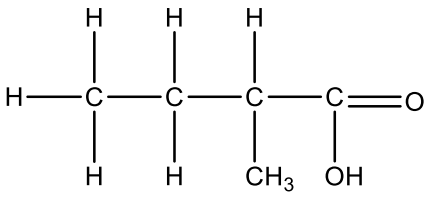
2. Draw the structure isomer of 4-Methylpentane-2-one.
3. Draw and name all isomers of butanone.

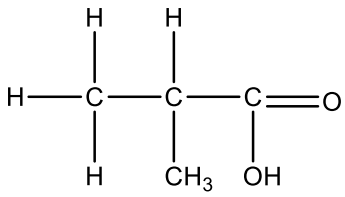
9.17 Isomers of carboxylic acids.

This has only **chain isomerism since –COOH is always at position 1**. Here, we look at the position of **alkyl group** or any other **substituents**. When naming, we follow some procedures as we did with the other compounds.

Example 1

Name the isomer of the following compounds

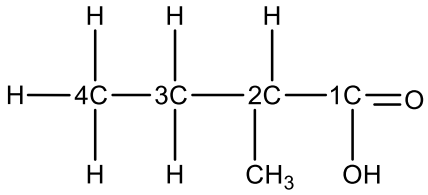




A B

Solution a

Number the carbon chain.



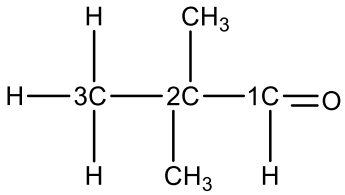
This has **methyl group (CH3**) at position 2 **(2-Methyl-).**

The carbon chain has 4 carbon atom, with its **R-COOH** group at first position. Hence, it is **Butanoic acid**

The complete name will be **2-methyl butanoic acid**

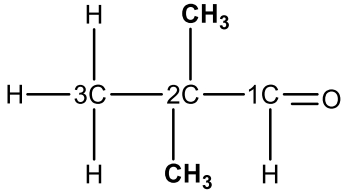
Solution b

Number the carbon skeleton



The chain has 3- carbon atoms hence it is **propanoic acid.**

The carbon chain has **two methyl group (CH3**) at second position.



This can be written as **2, 2- dimethyl**

Therefore, the complete name will be **2,2- dimethyl propanoic acid**

Drawing structure isomers of carboxylic acid.

Given a name, we can draw the structure isomers of carboxylic acid. This is done as follows;

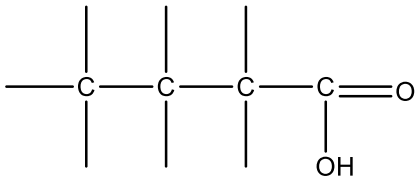
- Identify root name
- Drawing the carbon skeleton which includes **functional group**
- Identify **side group (substituents)** and attach them to carbon chain

Example 1

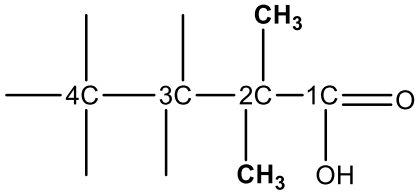
Draw the structure isomer **of 2,2 –dimethyl butanoic acid**

Solution 1

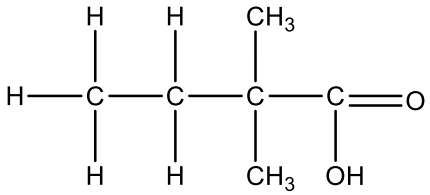
The root name is **butanoic acid**. This means it has **4- carbon** atoms which include (R-C=O) *functional group*



2, 2-dimethyl means that it has the (**CH_3)** group both at position 2.

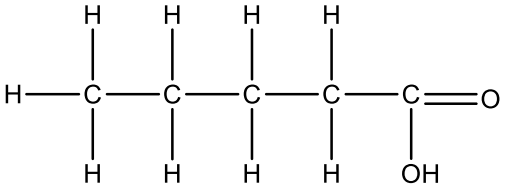


The complete structure will be

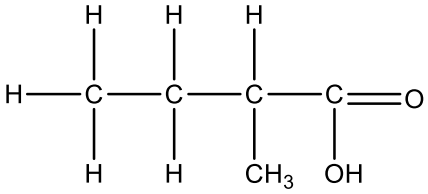


Example 3

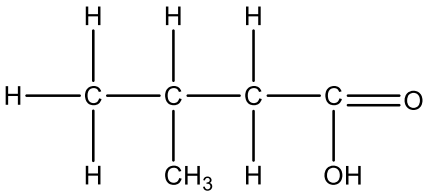
Write down and name all isomer of pentanoic acid



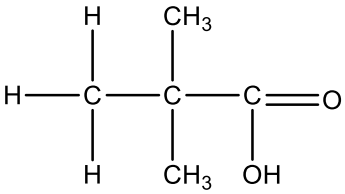
n-Pentanoic acid



2-Methylbutanoic acid



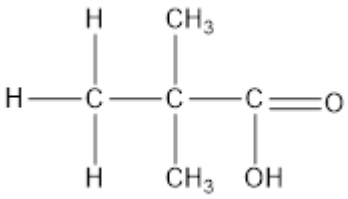
3-Methylbutanoic acid



2,2-dimethylpropanoic acid

Self-Assessment 9.6

1. Name the isomer of the given compound.



2. Draw the structure isomer of 2,3-dimethylbutanoic acid.
3. Draw and name all isomers of Hexanoic acid.

9.18 Effects of braching on physical properties of organic compound.

(a) Melting and boiling point

Boiling point

Branched isomer have low boiling point. This is due to closely packed together of molecules which decreases surface area. This reduces intermolecular force between molecules; hence they are easily broken down by less amount of heat energy unlike in unbranched isomers.

Melting point

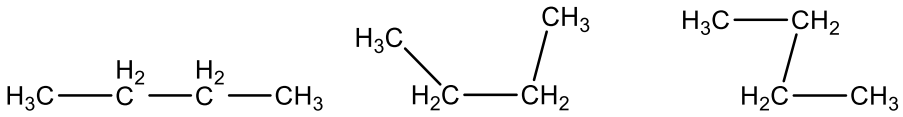
Higher branched isomer has **high melting point** because of its solid state arrangement of molecules. This increase more force of attraction that hold molecules together hence more heat energy is needed to break forces

9.19 Conformations.

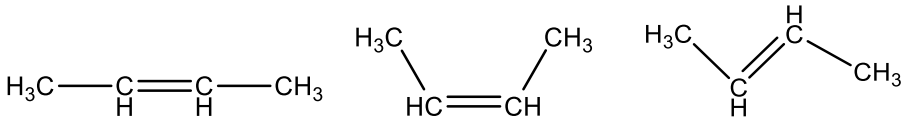
Having looked at isomers, let us look the molecules having the same molecular and structure formula but different shapes. These molecules are called **conformations**.

Examples

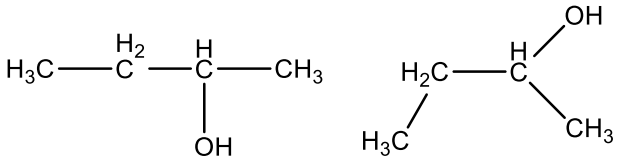
Conformations of Butane



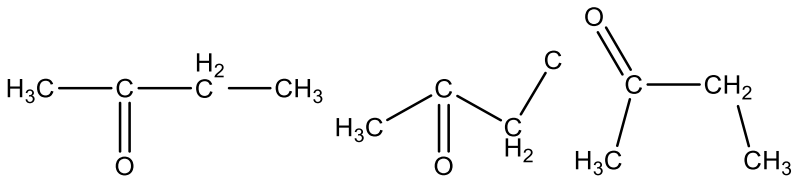
Conformations of butene



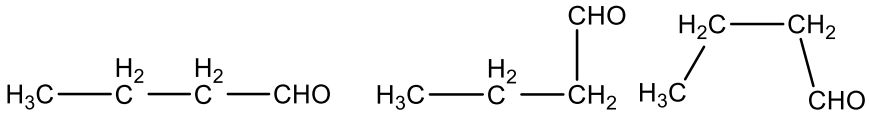
Conformations of butanol



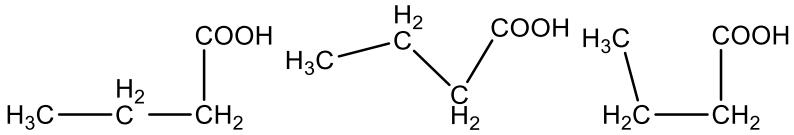
Conformations of butanone



Conformations of Butanal



Conformations of Butanoic Acid



9.20 Summary.

Isomers are organic compounds which have the same chemical formula but different structural formula due to different arrangement of atoms. In this book, you have studied the isomers of different compounds like alkanes, alkenes, alkanols, alkanols and carboxylic acids.

Some of the compounds like alkanols, alkanones and alkenes have both chain and position isomers because of the functional group present in the carbon chain length can be shifted to different position.The process of naming isomers follows International Union of Pure and Applied Chemistry **(IUPAC)** rules.

Branched isomers, have low boiling than unbranched because less energy is required to break less forces of attraction between the molecules.

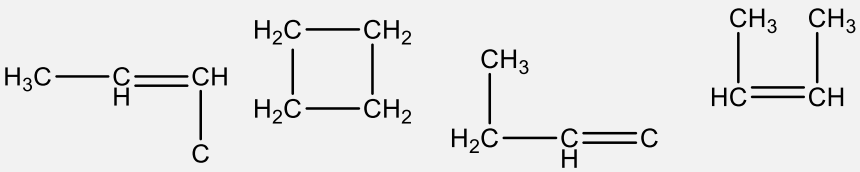
9.21 Revision Exercise.

1. (a) Write down and name all isomers of the following compounds
   - i. Butane
   - ii. Pentene
   - iii. Butanal
   - iv. Pentanone
   - v. Pentaoic acid

(b). Draw the structure formula of the following compound.

- i. 3, 3- dimethyl pentan-2-one
- ii. 2, 3- dimethyl pent-2-ene
- iii. 3- methyl pent-2-ene
- iv. 4- methyl hexan-2,3-dione

2. The following are structure formulae of four molecules. Use them to answer given questions



A **B** **C** **D**

I. Define "**isomer"**
II. Describe two types of isomers.
III. Name the molecules **A** and **B**
IV. Which two structures are
    a. Corfomations of each other
    b. Isomers of each other

IV Explain your answer to 2III.

2 Explain why branched isomers have low boiling but high melting point

UNIT 10

Polymerization.

Learning objectives.

By the end of this unit, you should be able to:

a) Define polymer
b) State the types of polymers
c) Explain the difference between **addition** and **condensation** types of polymerization
d) State the properties and uses of the polymers
e) Explain the difference between thermos softening and thermos setting plastics
f) State the ways of mouldling plastics.

10.10 Introduction.

In polymer chemistry, **polymerization** is a process of reacting *monomer molecules* together in a chain of chemical reaction to form polymer chains (three-dimensional networks)**.**For example**,** In living system, small molecules (***amino acids***) join up to form large molecules (***proteins***).

In chemical compounds, polymerization can occur via a variety of reaction mechanisms that vary in complexity due to the **functional groups** present in the reactants and their inherent steric effects. In more straightforward polymerizations, **alkenes** form polymers through relatively simple *radical reactions*.

In contrast, reactions involving substitution at a **carbonyl group** require more complex synthesis due to the way in which reactants polymerize. **Alkanes** can also be polymerized, but only with the help of strong acids. There are many forms of polymerization and different systems exist to categorize them.

Polymer is the substance possessing very large molecules consisting of repeated units or monomers linked together.These are also called ***Macromolecule.*** On the other hand, the **monomer is** a simple molecule from which polymers are built. The monomers are also called micro molecules. Therefore, **polymerisation** can be simply defined as ***the chemical reaction in which monomers, join together to form a polymer***.

The difference between monomers and polymers.

Monomers are small molecules that join together to form long chain of atoms or molecules while polymers are molecules which are formed when small molecules join together to form long chain.

Self-Assessment 10.1

1. Define polymerization.
2. Describe between a polymer and monomer.

10.11 Methods of polymerisation.

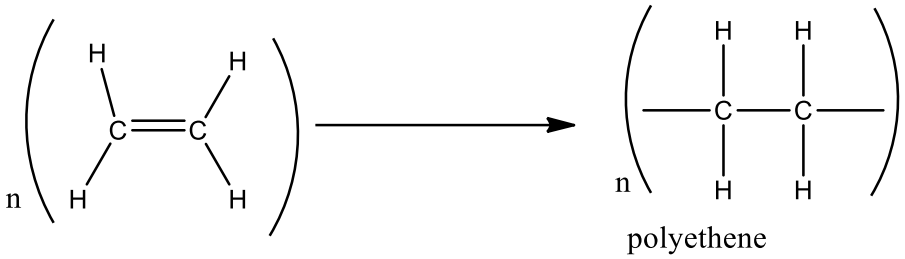
There are two methods of polymerization.

- Additional polymerization.
- Condensation polymerization.

(a) Additional polymerization.

This is process in which monomer units of the same type join together to form long chain of one product. For example, alkene molecules add into each other to form long chain molecules. Alkene may do this, because the weak bond in Carbon to Carbon double bond **(C=C)** is easily broken in chemical reaction.

For example, if **Ethene** is needed and put under pressure. Its molecules join up to form **polyethene** or **polythene.**

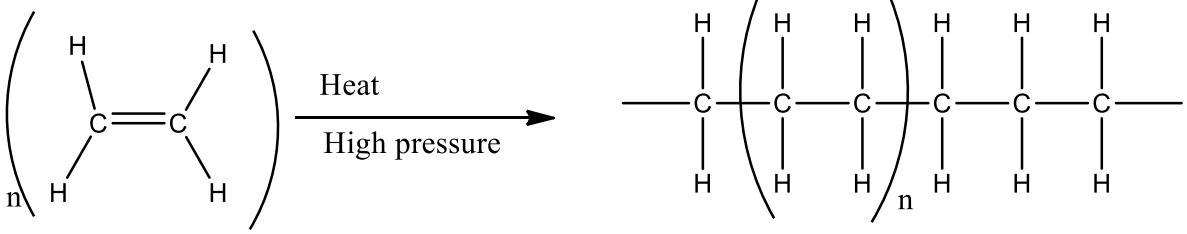


Ethene monomer

polyethene polymer

Where **n** is very largest number. In this case, **ethene** is monomer while products formed **poly ethane** is polymer.

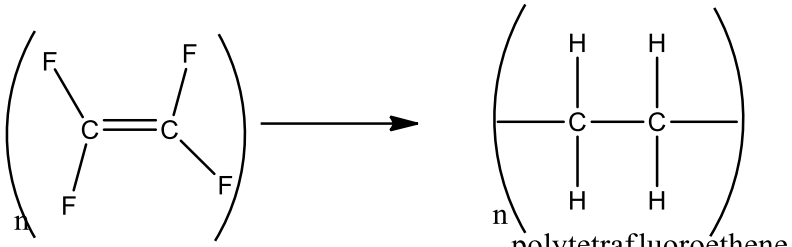
The process above can be also be written in this way



Ethene monomer

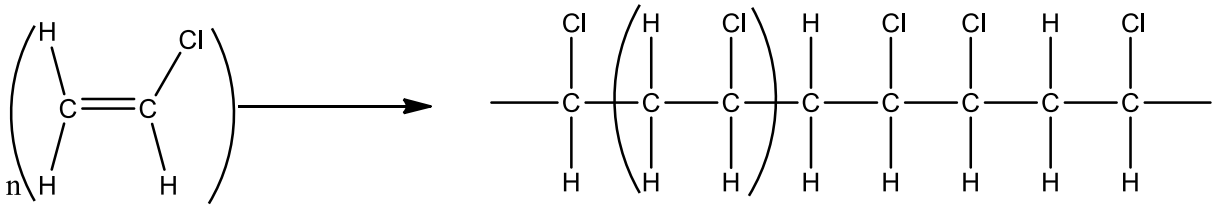
Polyethene polymer

The other example of addition polymensation.



Tetrafluoroethene
(monomer)

polytetrafluoroethene
(polymer)



Monochloroethene
(monomer)

polychloroethane or Poly vinyl chloride(PVC)
(polymer)

The polymers from group of compounds called plastics with different properties because of the different in the molecular units.

The table 10.11a. Some addition polymers

Plastic addition polymer	**Monomers**	**Properties**	**Uses**
Poly ethen	$CH_3=CH_2$ ethen	Tough durable corrode easily moulded not affected	• Making bouds • Making baskets • Packing • Making caries bags
Polypropen	$CH_3CH=CH_2$ propene	• Tough • Durable	• Making ropes • Packing
Poly chloride (pvc)	$CH_2=CHCl$ vinyl chloride	• Strong • Hard but flexible	• Electrical insulation • Making pipes • gattering
Poly tetra fluo ethane	$CF=CF_2$ tetraflouro ethane	• hydrophobic • non-stick surface • resistance to chemical attack • wind high pressure	• non-stick prying pans • soles of iron
Polystyrene	$CH2=CHC_6H_5$ styrene	• lighter • poor conductor of heat	• insulation • packaging

Perspex	$CH_2=C(CO_2CH_3)\ CH_3$ methyl propanoate	• Transparent	Used as a glass substitute

Forms of poly ethene

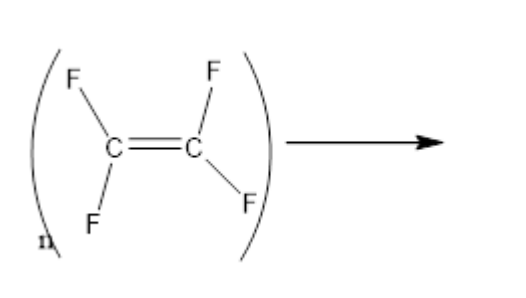
Poly ethene is produced in three main forms;

- low density poly ethane (LDPE)
- Linear low density poly ethene (LLDPE)
- High density poly ethane (HDPE)

Therefore, poly ethane formed by addition polymerisation can be represented by **A-A-A-A-A**; where **A** is monomer.

Self-Assessment 10.2

1. Describe addition polymerization.
2. Name three uses of poly vinyl chloride.
3. Complete the structure.



4. State three forms of poly ethene.

(b) Condensation polymerization.

This is the process in which different monomers react with each other to form a long chain molecule with the loss of small molecules of water. Each monomer contains two function groups.

Condensation polymers.

- Nylon
- Terylene (poly ester)

Uses of Nylon.

Used as fiber in sewing

- thread,
- clothes,
- computer ribbons

Nylon is formed from two monomers **1, 6- diamino hexane** and **hexane dioic acid**

$$H_2N(CH_2)_6NH_2 + HOOC(CH_2)_4COOH \longrightarrow H_2N(CH_2)_6NHOC(CH_2)_4COOH + H_2O$$

Each time a reaction takes place, the **water** is produced, and then it is called **condensation reaction**. In condensation polymerisation, different kind of monomers react which result in a chain of this type;

A-B-A-B-A-B-A-B-A-B-A-B where **A** is a monomer and **B** is also a monomer of different type.

Nylon is known as polymide because the amide link is formed during polymerisation.

Uses tylene

- Clothing

Different between addition polymerisation and condensation polymerisation

- In addition, polymerisation there is no loss of water molecules while in condensation polymerisation there is loss of water molecules
- In addition, polymerisation one type of monomers join to from polymer while in condensation polymerisation two different type of monomers are involved in the reaction

Self-Assessment 10.3

1. Describe condensation polymer.
2. Name three forms of condensation polymer.
3. State three uses of Nylon.
4. Describe the difference between condensation and addition polymerization.

10.12 Types of polymers.

- Biopolymers (natural)
- Artificial polymer (synthetic)

(a) Natural/biopolymers.

These are not man made polymers. They occur naturally in plants and animals.

Table:10.12a Natural polymers.

Natural polymer	Monomer	Uses
Proteins	Amino acids	Grown, immunity
Starch	Glucose	Provide energy
DNA	Nucleotide	Control proteins synthesis
Cellulose	Glucose	Making paper/clothing
Rubber	Isoprene	Making tyres

(b) Synthetic (artificial) polymers.

These are man-made polymers formed by reaction of monomers.

Ways of making synthetic polymers.

- *By addition polymerisation of alkenes* e.g. poly ethene, poly propenes, poly vinyl chloride (PVC)
- *By condensation polymerisation* e.g. terylene and nylon

In general, examples of synthetic polymers are;

- Poly ethene
- Poly vinyl chloride
- Polytetra flaon ethene
- Poly propene
- Nylon
- Terylene

Advantages of natural (biopolymers) polymers.

- They do not produce poisonous gases when they burn
- They are bio degradable. Bio gradable means that they can be decomposed by macro-organisms easily.
- Made from renewable resources.

Disadvantages of natural polymers.

- Expensive
- They cannot be moulded into complicated shape
- They cannot last for very long.

Properties of synthetic polymers.

- Do not decompose easily
- Electrical insulator
- Do not corrode in air or water
- Relatively light
- Transparent

Advantages of synthetic/arterial polymers.

- They are cheap to produce
- Good insulator
- Portable

- Resistant to acids, alkali, air and water
- Can be moulded into different shape

Disadvantages of synthetic polymers

- Expensive to recycle
- Not readily bio gradable
- After burnt, they produce poisonous gase.
- They have comparative low strength
- Poor heat resistance.

10. 13 Factors that affect properties of polymers.

Chain strength.

As chain strength increases, the strength and melting of polymers also increases.

Intermolecular forces.

The higher the intermolecular forces, the stronger the polymers and it becomes more difficult to meet.

Branching

A higher branched polymer has lower tensile strength. Therefore, they melt more easily.

Crossing linking

The more the crossings, the more the rigid and brittle the polymer becomes.

Self-Assessment 10.4

1. List two examples of the following polymer.
   Natural
   Synthetic
2. Describe two ways of making synthetic polymers.
3. Describe:
   i. Advantages of natural polymer
   ii. Disadvantages of natural polymer.
   iii. Advantages of synthetic polymers.
   iv. Disadvantages of synthetic polyemers.
4. Describe three factors that affect properties of polymer.

10.14 Plastics.

There are two types of plastic.

- Thermoplastic (thermos softening plastics)
- Thermosetting plastic/ thermos set.

(a) Thermoplastics/thermosoftening plastics.

These are kinds of plastics that melt or soften when heated.

Examples.

- Poly ethene
- Polystyrene
- Poly vinyl chloride (PVC)
- Poly tetra flour ethane

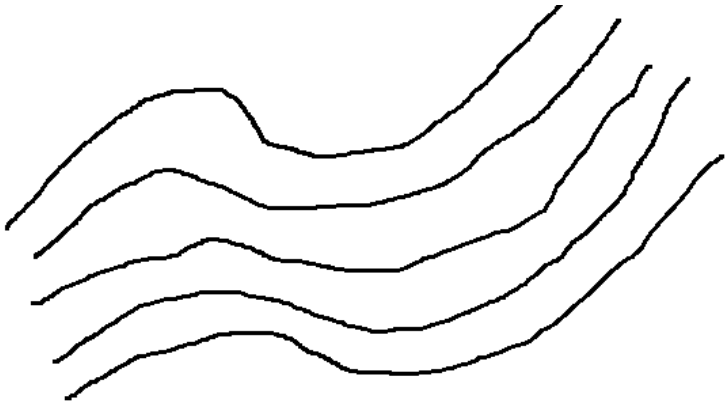
Advantages of thermoplastics.

- They are easily moulded
- They are easily soften or melt when heated

Reason why thermoplastic melt when heated.

The polymer chains are not cross linked. So when they are subjected to heat, their polymer chains flow over one another and the plastics melt.

Figure10.13a. Thermoplastics/ thermos softening plastics.



Characteristics of thermoplastics.

- They are flexible and therefore, do not break easily
- The soften hence they melt at low temperature
- They stretch under tension because molecules are able to slider over each other

(b) Thermosetting plastics.

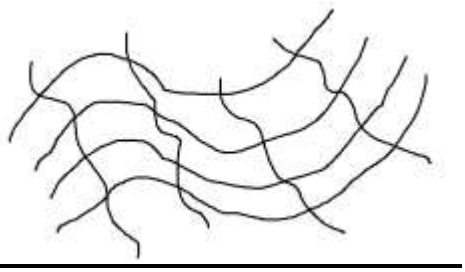
The kind of plastic which do not melt or soften on heating. They only Chur (reduced to charcoal) and decompose on further heating.

Examples:

- Melamine
- phenolic

Thermosets (thermosetting) can be heated and be moulded only once usually by compression moulding.

The figure 10.13 Thermo-setting plastics.



Reasons why thermosetting do not soften on heating.

Polymer chain which are linked or bounded to each other to give a cross linked structure and the chains are firmly in space. When they are heated they char and decompose. Therefore, no softening takes place.

NB. Thermosetting plastics can be reconverted into thermoplastic by opening the cross link. They can be made soft by the addition of certain compound, unknown as **plasticisers.**

Advantages of plastics over metal substances.

- They are light
- They are relative unreactive
- They can easily be moulded
- They can be dyed into bright colour
- They do not corrode
- They do not conduct electricity hence are used as insulators
- They are water proof

Disadvantages of plastics.

- They do not decay hence they cause environmental and sanitation problems.

10.15 Waste management of plastics.

Recycling.

Large quantities of plastics are produced for resale

Companies should produce biodegradable and photodegradable plastics.

Bio degradable plastics are broken down by bacteria while photodegradable plastics are broken down by sunlight

Incineration or burning the plastics.

Schemes have been developed to use the heat generated for healing process.

The table 10.15 Differences between thermoplastics and thermosetting plastic.

Thermo softening plastics	Thermos setting plastics
Soften when heated and easily remoulded	• Do not soften on heating and cannot be remoulded
Have no cross link chain hence held together by van-der waals forces	• They have cross link chains which held then together by permanent chemical bonding
Highly branched hence low tensile strength and low melting points	Not branched hence high tensile strength and high melting points

Self-Assessment 10.5

1. Describe the differences between thermosetting and thermosoftening plastics.
2. State three advantages of thermosoftening plastics.
3. Explain why thermosetting plastics do not melt when heated?
4. State two examples of:
   Thermosoftening plastics.
   Thermosetting plastics.
5. Describe three advantages of plastics over metal substances.
6. Describe three waste management of plastics.

10.16 Summary.

In this unit, you learnt that polymers are made from monomers which join up by chemical process called **polymerisation**. The products are plastics which are categorized as **thermoplastics** and **thermos setting plastics** depending on their ability to melt upon heated.

However, polymers can be natural **(biopolymers**) and **synthetic (man-made)** of which are produced from the two methods of polymerisation **Addition** and **Condensation.**

10.17 Revision Exercise.

1 (a) Define a polymer

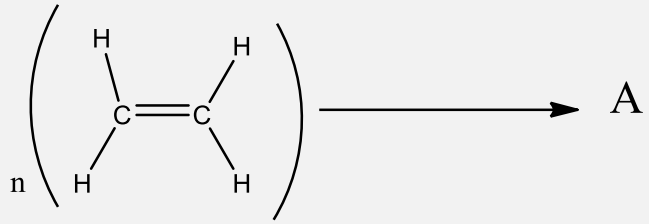
Explain the difference between synthetic and natural polymer

(a) State four advantaged and disadvantages of the following polymers;

i. Synthetic polymer

ii. Natural polymers

1. (a) Study the equation below and use it to answer the questions.



Ethene
monomer

i. Complete the equation

ii. Identify the process above

(b) Define "condensation polymerisation"

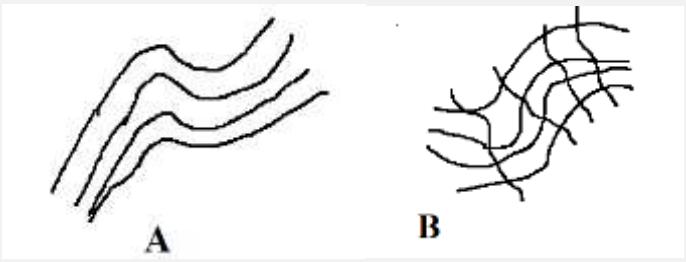
2. (a) Mention three characteristics of thermoplastics

(b) Explain why thermos setting plastics do not melt when heated?

(c) State three differences between thermoplastics and thermos setting plastics.

(d) State the difference between **polymer** and **monomer**

Use the figure to answer the questions



Identify the plastic labelled

A

B

UNIT 11

Rates of reaction.

Learning objectives

By the end of this unit, you should be able to:

- Define the rate of reaction
- Describe factors that affect rate of reaction
- Describe reversible and irreversible reaction
- Explain chemical equilibrium in reversible reactions

11.10 Introduction.

In unit 4, you looked at chemical reaction takes place to produce products. A measure of the change which happens during a reaction in a single unit of time is called **reaction rate**. In chemical reaction, old bonds in the reactants break to form new bond in the products. This requires energy to occur. Therefore, to be able to know how fast a reaction is taking place, we measure the rate at which a reactant is consumed or a product is formed over a given period of time. i.e.

$$\textbf{Reaction rate} = \frac{\textbf{change in amount of substance}}{\textbf{Time taken}}$$

Therefore, chemist and chemical engineers have greatly looked for ways to control the rate at which chemical reactions take place. In doing so, they have discovered that the rate of a chemical reaction can be regulated. The ideas are not only incredibly useful to industry but can also be applied to reactions which occur in the body, school and clinical laboratories.

Collision theory.

This is based on the kinetic theory of matter which states that substances are made up of a very smallest particle which are in constant motion. This motion produces collision between particles of which is a basic fundamental principle on chemical reaction. Therefore, the theory is based on the three main points for a chemical reaction;

- Reactions take place due to collisions of reacting particles like atoms, ions and molecules.
- Collisions which produce a reaction are fruitful collisions. Therefore, the rate at which particles collide is called ***collision frequency***. The greater of this collision frequency, the faster the reaction rate.
- Colliding particles need certain amount of energy before they can bring about a chemical reaction. This energy is called ***activation energy***.

Activation energy.

The excess energy that a reaction must acquire per unit for the reaction to occur.For a chemical reaction to occur very quickly, the activation energy must be very small i.e. less than zero.

Self-Assessment 11.1

1. Define Reaction rate.
2. Describe:
   Activation energy.
   Collision theory.
3. Describe three points of collision theory for the chemical reaction to take place.

11.11 Factors that affect the rate of a reaction.

Surface area

Small particles increase surface area for a chemical reaction to occur more quickly. **Why?** Large surface area increases the chances of the successful collisions between particles. This leads to greater frequency of these collisions. Therefore, greater frequency collision leads to an increase in the rate of chemical reaction.

Concentration of the reactions

The products of the reaction are formed as a result of the collisions between reactant particles. When there more particles in a more concentrated solution, collisions occur more often. The more often they collide, the greater the chance of high frequency collisions. This leads to an increase in the rate of a chemical reaction to occur very quickly.

Temperature

When the temperature at which the reaction is carried out is increased, the particles gain more kinetic energy. This causes the particles to move with high velocity.This increases the number of collisions between particles and collisions which occur are more energetic of which lead to higher frequency collisions. This causes higher rate of a chemical reactions.

Light

Some reactions need light. Photosynthesis is Avery good example. ***Why and how?*** Light acts as an external source of activation energy. The increases in light intensively, increases frequency collision between particles of which increases the rate of chemical reaction.

Catalyst

This can alter the rate of a reaction without being chemically changed itself. A catalyst increases the rate by providing an alternative reaction path with lower **activation energy.**

Activation energy is the energy barriers which reactants must overcome, when their particles collide, to react successfully and form products.

11.12 Reversible reaction.

The reaction can go in either directions**.** This is presented by arrows pointing in both directions. I.e.

$$A+B \rightleftarrows C+D$$

Where **A** and **B** are reactants and **C** and **D** are products.

Reversible reaction is based on the fact that products can react to form the reactants agian. Therefore, C and D can react to form A and B in reversible reactions.

The good example, of reversible reaction is esterification which involves;

Ethanoic acid + ethanol $\xrightleftharpoons{H^+}$ ethanoate + water

$CH_3COOH_{(aq)} + C_2H_5OH \xrightleftharpoons{H^+} CH_3COOC_2H_{5(aq)} + H_2O_{(l)}$

The double arrows, means the reaction is reversible. Therefore, ethyl ethanoate plus water can react to give ethanoic acid and ethanol again. As such;

$CH_3CCOC_2H_{5(aq)}\ H2O\ (l) \xrightleftharpoons{H} CH_3COOH_{(aq)} + C_2H_5OH_{(g)}$

NB*: Reactants and products are formed at the same rate.*

11.13 Irreversible reaction

The reaction can go in one direction only. This means it is one way. Therefore, it is presented by single arrows pointing in one direction

$$A + B \longrightarrow C + D$$

A and **B** are reactants, and C and D are products just like one we saw in reversible reaction to form **A** and **B** group again.

All combustion reactions are good examples of irreversible reactions. Once the products are formed, the recantants cannot be reformed. For example;

Magnesium + oxygen $\longrightarrow$ Magnesium oxide

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

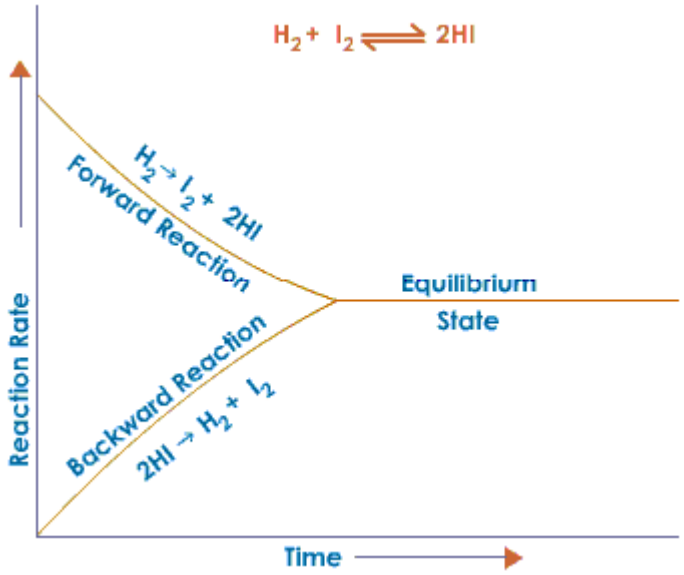
The magnesium oxide cannot form magnesium and oxygen basic.

Self-Assessment 11.2

1. Describe two factors that affect reaction rate.
2. Differentiate reversible and irreversible teaction.
3. State two examples of:
   Reversile reaction.
   Irreversile reaction

11.14 Chemical equilibrium.

In a chemical reaction, **chemical equilibrium** is the state in which both reactants and products are present in **concentrations** which have no further tendency to change with time, so that there is no observable change in the properties of the system. Usually, this state results when the **forward reaction** proceeds at the same rate as the **reverse reaction**. The ***reaction rates*** of the forward and backward reactions are generally not zero, but equal. Thus, there are no net changes in the concentrations of the reactant(s) and product(s). Such a state is known as **dynamic .**The chemical equilibrium is also called the **state of balance.**



The figure The graph of chemical equilibrium

Effects of different factors on the position of equlibrium.

Factor	**Effects.**
Increase for reactants (decrease for products).	• This favours forward reaction. • Chemical equilibrium shifts from left to right. • More products are formed.
Decreases for reactants (Increase for products).	• Favours backward reaction. • Equilibrium shift from right to left. • More reactants are formed.

Increases in the number of moles of functional of products. (Pressure increases)	• Favours reverse reaction. • Equilibrium is shifted to the left. • More reactants are formed in reverse reaction.
Decreases in the number of moles on formation of products. (Pressure increases)	• This favours forward reaction. • Equilibrium is shifted to right. • More products here are formed in forward reaction.
Increase in the number of moles on formation of products.(Pressure decreases)	• This favour forward reaction. • Equilibrium is shifted to the right. • More products are formed.
Decrease in the number of moles on formation of products. (Pressure decreases)	• This favours reverse reaction. • Shift equilibrium from right to left. • More reactions are produced.
Temperature increase in exothermic reactions.	• Favour reverse reaction. • Shift from right to left. • More reactants are formed.
Temperature decrease in exothermic reactions.	• Favour forward reaction. • Equilibrium shift from left to right, • More products are formed.
Temperature decrease in endothermic reactions.	• Favour reverse reactions. • Equilibrium shift from right to left. • More reactants are formed.

Self-Assessment 11.3

1. Define chemical equilibrium.
2. Draw the sample graph for chemical equilibrium.
3. Describe the following factors on chemical equilibrium.
    i. Increase in the number of moles on formation of products. (Pressure decreases)
    ii. Increase for reactants (decrease for products).
    iii. Temperature decrease in endothermic reactions.

11.15 Summary

The reaction rate can be defined as a rate of change of an amount or concentration of a particular reactants or product per unit. Several factors; surface area, light, temperature, concentration and catalyst affect the rate of reaction by altering the frequency of collision between particles. In chemical reaction, the equilibrium is attached when the rate of forward reaction is equal to the rate of backward reaction. The chemical equilibrium is also called the state of balance. This changed when the factors that affect the rate of reactions have been altered.

11.16 Revision Exercise.

1. (a) Define the rates of reaction
    (b) Explain how the following factors affect the rate of reaction
    i. Temperature
    ii. Concertation of reactants
    (b) Differentiate reversible and irreversible reactions
    (c) How does activation energy affect chemical reactions?
2. (a) Define “chemical equilibrium”

(a) Explain how does the following factors affect chemical equilibrium;

i. Increases in temperature
ii. Decreases in pressure
iii. Increases in concertation of reactants
iv. Decrease in temperature in endothermic reaction

3. (a) Mention three factors that facilitates the chemical reaction to take place

UNIT 12

Acids and bases II

Learning objectives.

By the end of this unit, you should be able to:

a) Define acid and base.
b) Identify conjugate acid- base pairs from given equation.
c) Explain the difference between strength and concentration of an acid or base.
d) Explain ways of regulating pH in different environment.
e) Classify oxides as acidic, basic or amphoteric.
f) Describe different ways of preparing salts.
g) Design an experiment to prepare and purify a salt.
h) State the application of precipitation.

12.10 Introduction.

In JC chemistry, you looked at acids and bases. According to Lowry bronsted theory, we stated that an acid is a proton donor while a base is proton acceptor. An acid donates hydrogen ions $\mathbf{H^+}$ to the base, and bases accept the hydrogen ion H+ from the acid in the chemical reaction.

The hydrogen $\mathbf{H^+}$ is called proton. Hydrogen ions **($\mathbf{H^+}$)** are responsible for *acidic properties* while the $\mathbf{OH^-}$ are responsible for the *basic properties*. Let us consider the chemical equation below;

$$HCl_{(aq)} + NH_{3(aq)} \rightleftharpoons \mathbf{NH_4^+} + Cl^-$$

In this case, hydrochloric acid (HCl) dissociate to form ions ($H^+ + Cl^-$). Therefore, it loses hydrogen ions, $\mathbf{H^+}$ (**proton**) while ammonia (NH^3) gets the hydrogen ion (NH_3+**$\mathbf{H^+}$**) to form ammonium **($\mathbf{NH_4}$).** The study of acid and base is called **proton transfer.**

12.11 Properties of Acids and Bases.

Properties of acids.

The following are the properties of acids.

- They conduct electricity
- Facilitate rusting
- Have pH less than 7
- React with metals higher than hydrogen in the electrochemical series
- Change the blue litmus paper to red
- React with base in neutralisation reaction to produce salt and water.
- Have sour taste

Properties of bases.

Properties of base are;

- Have bitter taste
- Have pH greater than 7
- Change red litmus paper to blue
- They conduct electricity
- They react with acid to form salt and water.

Self-Assessment 12.1

Define Acid.

State three properties of:

Acids.

Base.

12.12 Conjugate acid- base pairs.

Conjugate acid is the base that have accepted a proton in form of hydrogen ion (H^+). It is also defined as the substance formed when an acid loses hydrogen ion.

For example:

$$\mathbf{HCl_{(aq)} + NH_{3(aq)} \longrightarrow NH_{4(q)} + Cl^-}$$

The ammonia (NH_3) has accepted the proton ($H^{+)}$ to form ammonium (NH_4). Therefore, ammonium (NH_4) is called ***conjugate acid.***

Conjugate base is the acid that has donated a proton in form of hydrogen ions (H+). let's take the same equation of hydrochloric acid and ammonia to find out what a conjugate base is.

$$\mathbf{HCl_{(aq)} + NH_{3(aq)} \rightleftharpoons NH_{4(q)} + Cl^-}$$

In this case, hydrochloric acid (HCl) has lost a proton in form of hydrogen ions (H+) and becomes **Cl-.** Therefore, **Cl**-is a ***conjugate base.***

In general, **HA +B** $\rightleftharpoons$ **HB^++A-**, where **HA** is an acid while **B** is base. In the reaction, **HA** react with **B**, the hydrogen ions will be donated and becomes **HB^+** which is **conjugate acid.** This leaves **A**- as **conjugate base.**

To get the conjugate of an acid subtract a hydrogen (proton) from the acid. For instance,

CH_3COOH-**H** $\longrightarrow$ **CH_3COO^-**

Acid conjugate

Conjugate of acid means ***conjugate base.***

To get the conjugate of base, add a hydrogen **(proton)** to the base. For instance,

$H_2O+ \ \mathbf{H^+} \longrightarrow H_3O+$

Base proton

The conjugate of base means ***conjugate acid.***

Example 1

In the following equation give the conjugate acid base pair.

$CH_3COOH_{(aq)}\ H_2O \rightleftharpoons CH_3COO^-_{(q)} + H_3O+_{(aq)}$

Solution.

Conjugate acid base pair

i. CH_3COOH/CH_3COO^-
ii. H_3O / H_2O

Example 2

Write down the conjugate acid base pair in the following reaction.

$NH_{3(aq)} + HCl_{(aq)} \rightleftharpoons NH_4^+{}_{(aq)}\ Cl^-_{(aq)}$

Solution

Conjugate acid base pair.

i. HCL/Cl-
ii. NH4+/NH3

Formation of hydronium ion (H_3O^+)

This is formed when two water molecules react where one water molecule acts as an acid while the other one acts as a base.

- $H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+ + OH^-$

  Acid base hydronium hydroxide

- H_3O+ is called a hydronium ion.

Conjugate acid base Paris

i. H_2O/OH^-
ii. H_3O/H_2O

Water is an amphoteric substance because it can act as an acid or a base depending on

Self-Assessment 12.2

Define.

Conjugate base.

Conjugate acid.

Write down the conjugate acid-base pairs of the following.

$CH3COOH(aq) + H2O(l) \longrightarrow CH3COO^- + H3O(l)$

Describe how hydronium is formed?

Hydrolysis.

This is reaction with water. Water can also react with ions.

Example

$HC_3^-(aq) + H_2O(l) \quad H_3O^+ + CO_2^-$

$HSO^-_{4(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+ + SO^{3-}$

Further examples;

Water acting as an acid

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

Water acting as a base

$HCl_{(aq)} + H_2O \rightleftharpoons H_3O^+ + Cl^-$

12.13 Strength of acids and bases.

pH scale.

The measure of acidity and basicity of a solution. The P_H ranges from 0 to 14.

P_H is the concentration of hydrogen ions in a solution while the **concentration** *of an acid or base is the mass of the acid or base that is dissolved in a definite volume of a solution*. Therefore, it indicates an aqueous solution. Therefore, it can be also defined as the negative logarithm of the activity of the hydrogen ions in aqueous solution.

pH=-log (H+)

PH scale

Increasing acidity ⟷ ***neutral*** ***increasing basicity***

PH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
colour	Dank red	Red	Orange red	Light orange	Yellow	Light green	green	Dark green	Green blue	Green blue	blue	Dark blue	purle	purle	purle

Ph from O-6 means that the solution is **acidity**, P_H 7 means that the solution is **neutral** and P_H from 8-14 means the solution is **basic or alkaline**.

The smaller PH value the more acidic the solution is and the greater the PH values the more basic or alkaline the solution is.

The universal indicator is preferred when measuring the PH value of a solution because it gives the degree of acidity or basicity of a solution. This gives a wide range of colours depending on acidity or basicity of a solution.

The strength of an acid or base tells you how easily it dissociates (ionises) to produce **hydrogen ions** or **hydroxide ions** respectively.

Strong acids.

This is an acid that ionizes completely in water to form hydrogen ions (H+). Example of strong acids are.

- Hydrochloric acid (HCl)
  $HCl_{(aq)} + H_2O_{(l)} \rightarrow$ H+ Cl^-
- Sulphuric acid (H_2SO_4)

  $H_2SO_{4(aq)} + H_2O_{(l)} \longrightarrow H^+ + SO^2_4$
- Nitric acid (HNO3)

$HNO_{3(aq)} + H_2O_{(l)} \rightarrow$ H+ NO_3^-

Weak acids.

The acid which partially ionises when it dissolves in water. It produces few hydrogen ions (H+). For example,

- Acetic acid (CH3COOH)

- Methanoic acid (HCOOH)

Note: ***all carboxylic acids are weak acids.***

$CH_3COOH_{(q)} + H_2O_{(l)} \rightleftharpoons CH_3COO^-(aq) + H_3O_{(aq)}$

Strong bases.

This is a base that ionise completely when dissolved in water to produce hydroxide ion **(OH-)**

Examples of strong bases are;

- Sodium hydroxide (NaOH)
- Potassium hydroxide (KOH)

The equation can be shown as

$NaOH_{(aq)}\ H_2O_{(l)} \rightarrow Na^+ + OH^-$

Weak base.

The base which partially ionize when dissolved in water. This produces few hydroxide (OH-) ions

Example are;

- Ammonia (NH_3)

e.g. $NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH^+_{(aq)} + OH^-$

Ways of determining the strength of acid or base.

- Using universal indicator
- Using PH meter

Experiment 12.1

Aim: To identify a solution as strong acid, strong base, weak acid or base *using universal indicator*

Materials.

- 4 test tubes in a rank
- ammonia solution (NH_3), Dilute sulphuric acid (H_2SO_4)
- universal indicator, Sodium hydroxide (NaOH) solution, Measuring cylinder

Procedures

(a) Put $2cm^3$ of sulphuric acid in test tube
(b) Add 2 drops of universal indicator
(c) Note colour change of the solution
(d) Use universal indicator colour chat to determine the PH of the solution
(e) Repeat step a-d using sodium hydroxide, acetic acid, ammonia solution respective

Table results.

Solution	Dilute Sulphuric.	Sodium hydroxide.	Acetic acid.	Ammonia sodium.
Colour chat	Dark red	Purple	orange	Dark blue
PH	1	14	3	11

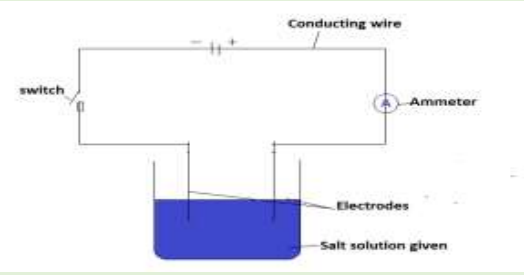
Based on the results, sulphuric acid is a strong acid while acetic acid is weak acid. Sodium hydroxide (NaOH) is strong base while ammonia solution is a weak base when compared with colour chart colour.

Experiment 12.2

Aim: Using conductivity apparatus, identify solution as strong acid, strong base, weak acid or weak base.

Material.

- beaker ,2 cells in holder
- sodium hydroxide (NaOH), dilute sulphuric acid (H_2SO_4)
- carbon electrode, acetic acid
- connecting wires, ammonia solution, Ammeter



Procedures

(a) set up experiment (figure)
(b) pour 75cm^3 of sulphuric acid in the beaker
(c) close the switch
(d) note the ammeter reading/ state of the bulb
(e) record in the table of results
(f) repeat the procedures with other solution, sodium hydroxide, acetic acid and ammonia solution

Solution	Ammeter reading	State of the bulb
Sulphuric acid	0.03	Bright
Sodium hydroxide	0.02	Bright
Acetic acid	0.01	Dim

Discussion.

From the results, sulphuric acid is a strong acid while acetic acid is a weak acid. Sodium hydroxide is a strong base while ammonia is a weak base. Sulphuric acid and sodium hydroxide are strong acid and base respectively because they produce a lot of ions when dissolve in water.Acetic acid and ammonia are weak acid and base because they produce few ions when dissolved in water.

Self-Assessment 12.3

1. Define the terms.
- pH.
- Concentration of acid.

2. Name two examples of:
- Strong acids.
- Weak acids.
- Strong base.
- Weak base.

3. Mention three ways of determining the strength of acids.

4. With the aid of well labelled diagram, Design an experiment to prove that Sulphuric acid is strong acid than methanoic acid.

12.14 Ways of regulating pH in different environments.

These are three applications of acid and base neutralization.

Stomach acid

The ant acid like milk for magnesia is taken to neutralise the excess stomach. Therefore taking milk of magnesia helps in reliving the pain

Insect's bits.

This reassesses methanoic which cause the pain. Therefore, dilute alkali can be applied on the affected area to neutralize it.

Soil

The lime is added to soil to increase its PH. When it has increase to the optimum PH level above, add the fertilizer of ammonium sulphate to reduce the PH.

2.15 Clasification of oxides.

The oxides are compounds formed when oxygen reacts with another element. The element can be *metals* e.g. sodium, magnesium and calcium and non- metal elements are carbon, sulphur and phosphorus. The oxides are classified into main three types:

(a) Acid oxides.

These are oxides of non-metals. Examples are

- Carbon dioxide (CO_2)
- Sulphuric dioxide (SO_2)
- Phosphorus pentoxide (P_2O_5)

These form carbonic acid, sulphurous and phosphoric acids respectively.

(b) Basic oxides.

These are oxides of metals. Examples are magnesium, iron, sodium, zinc and calcium.

When these oxides dissolve in water, they form alkaline solution. Good example of alkaline solution is sodium hydroxide (NaOH), magnesium hydroxide (NaOH) and calcium hydroxide $Ca(OH)_2$.

(c) Amphoteric oxides

Some oxides have both basic and acidic properties. They react with both base and acid in neutralization reaction to form **water** and **salt.** Examples are Aluminium oxide (Al_2O_3), zinc oxide (ZnO) and lead oxide (pbO).

Self-Assessment 12.4

1. State three of regulating pH in soil.
2. Define Oxides.
3. Name three examples of:
   Acid oxides.
   Basic Oxides.
   Amphoteric Oxides.

12.16 Formation of salts.

Let us consider the reaction between hydro chloric acid and sodium hydroxide (NaOH) to produce sodium chloride and water. i.e.

Hydrochloric acid + sodium hydrogen $\longrightarrow$ Sodium chloride+ water

$HCl_{(aq)} + NaOH(aq) \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

In this equation, the salt (NaCl) and water (H_2O) are the products produced. This kind of reaction is called neutralisation. Therefore, sodium chloride (NaCl) was produced as part of the neutralisation reaction. Compounds formed in this way are known as **normal salts**.

A normal salt is a compound that has been replaced by metal ions or by the ammonium ion (NH4+). These salt can be classified as those which are **soluble in water** or **insoluble in water**.

Examples of soluble salts in cold water are:

- All nitrates
- All common sodium, potassium and ammonium salts
- All chloride expect lead, silver and mercury
- All sulphates except lead, barium and calcium.

Table: useful salts.

Salts	Use
Ammonium chloride Ammonium sulphates	In torch batteries In fertilizer
Calcium carbonate Calcium chloride	Making cement 81 glass Extraction of sodium Making plaster boards Plaster for injured limbs
Iron ii sulphates Magnesium sulphates	In iron tablets In medicines
Potassium nitrate	In fertilizer Gun powder manufacture
Silver bromine	In photography
Sodium stereate	In soaps

Methods of preparing soluble salts.

There are four general reactions of preparing soluble salts.

Acid-metals reaction.

This method uses less reactive metals like magnesium, aluminium, zinc, iron and tin.

Consider this **experiment** below;

(a) Put 10 cm^3 of dilute nitric acid in test tube
(b) Add excess magnesium ribbon
(c) Observe what happens.

Observations/results.

Hydrogen gas is produced. In this reaction, the hydrogen ion from the nitric acid gain electrons from the metal atoms as the reaction proceeds.

Hydrogen ions + (electrons) $\longrightarrow$hydrogen gas

From metal

$2H^+ + 2e^- \longrightarrow H_{2(g)}$

The whole experiment can be summarised using equation below;

Magnesium + nitric acid $\longrightarrow$ magnesium nitrate + hydrogen gas

$Mg_{(g)} + 2HNO_{3(aq)} \quad Mg(NO_3)_{2(aq)} + H_{2(g)}$

Excess magnesium is removed by **filtration**

The magnesium nitrate solution is evaporated slowly to form a **saturated of the salt**.

Acid-carbonates reaction.

This involves reaction between any metal carbonate and any acid, the salt produced is similar to that carried out for an acid and a metal. For example, copper (ii) sulphate would be added in excess to dilute nitric acid.

- Effervescence would be observed due to the production of carbon dioxide (CO_2).

The equation below implies the whole explanation;

Copper (ii) carbon + nitric acid $\longrightarrow$ copper (ii) nitrate + carbon dioxide + water

$CaCO_{3(l)} + 2NHO_{3(aq)} \longrightarrow Cu(NO_3)_{2(aq)} + CO_{2(q)} + H_2O_{(l)}$

Metal carbonate contains carbonate ions, CO^{-2}_3. In this reaction the carbonate ions react with the hydrogen ions in the acid.

Carbon ions + hydrogen ions $\longrightarrow$ carbon dioxide + water

$CO^{2-}_3(aq) + 2H^+_{(aq)} \longrightarrow CO_2(g) + H_2O(l)$

Acid-alkali (soluble base) reaction.

This involves the reaction of acids and very reactive metals, sodium (Na) and potassium (K). it would be certainly be too dangerous to add the metal directly to the acid. Therefore, use an alkali which contains the particular reactive metal whose salt we wish to prepare. The soluble base is called **alkali**. However, there are several bases which are insoluble.

- In general, metal oxides and hydroxides are bases.
- Because in this neutralisation where an acid react base are in solution. The special technique **titration** is required
- Acid is slowly and added to a measured volume of alkali using a burette, until the indicator, usually phenolphthalein, changes colour
- An indicator is used to show when the alkali has been neutralised completely by the acid. This is called the end point.
- The solution which is produced can then be evaporated slowly **to obtain salt.**

Hydro chloric acid+ sodium hydroxide $\longrightarrow$ sodium chloride + water

$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

- This reaction can best be described by the ionic equation.

$H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$

Acid-insoluble base reaction.

This uses unreactive metals like lead or copper. In this case, it is not possible to use a direct reaction of the metals with an acid, so the acid is neutralised using particular metal oxides.The method is generally the same as that for the metal carbonate and an acid, though some warming of the reactants may be necessary. An example of a such a reaction is the neutralisation of sulphuric acid by copper ii oxide to produce **copper ii sulphate**.

This can be simplified by equation below:

Sulphuric acid + copper oxide $\longrightarrow$ copper ii sulphate + water

$H_2SO_{4(aq)}) \; CuO_{(s)} \longrightarrow CuSO_{4(aq)} + H_2O_{(l)}$

Metal oxides contain the oxide ions, O^{2-}. The ionic equation for this therefore,

- $\mathbf{2H^+_{(aq)} + O^{2-} \longrightarrow H_2O_{(l)}}$

Methods of preparing insoluble salts.

An insoluble salt, e.g. barium sulphate, can be made by **precipitation.** In this case, solutions of the two chosen soluble salts are mixed. The precipitates can ***be filtered off, washed with distilled water and dried***. The reaction that has occurred is:

Barium chloride + sodium $\longrightarrow$ barium sulphate + sodium chloride

$\mathbf{BaCl_{2(aq)} + Na_2SO_{4(aq)} \longrightarrow BaSO_{4(s)} + 2NaCl_{(aq)}}$

The ionic equation for the reaction is

$\mathbf{Ba^{+2}_{(aq)} + SO_4^{-2}{}_{(aq)} \longrightarrow BaSO_{4(s)}}$

This method is somethings known as **double decomposition** and may be summarised as follow:

Soluble + soluble salt $\longrightarrow$ insoluble salt + soluble salt

(AX) + BY $\longrightarrow$ BX + Ay

It should be noted even salts like barium sulphate dissolve to very small extent. For example, 1 liter of water will dissolve 2.2×10^{-3} of barium sulphate at 25^0C. These substances are said to be ***sparingly soluble***.

In general, all methods of salt preparation both soluble and insoluble can be categorised into three. These are;

- ❖ Direct combination
- ❖ Neutralisation
- ❖ Precipitation

In direct combination of element, we have started

Acid + insoluble base reaction

Acid and metal carbonate reaction

Acid and metal reaction

In naturalisation we have seen,

Acid and alkali (soluble base)

Finally, *is precipitation* which involves the reaction of **soluble salts** and **insoluble salts**

Application of precipitation reaction.

- Treatment of domestic water
- Treatment of industrial effluents

Ionic equation.

Let us consider the reaction of hydrochloric acid and sodium hydroxide solution

Hydrochloric acid + sodium hydroxide $\longrightarrow$ sodium chloride + water

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

When both hydrochloric acid and sodium hydroxide dissolves in water, the ions separate completely.

We may therefore write;

$$H^+_{(aq)}\ Cl^-\ aq + Na^+_{(aq)}\ OH^{-(aq)} \rightarrow \quad Na+_{(aq)} + Cl^-\ aq + H_2O_{(l)}$$

You may see that Na+ and Cl- are unchanged on both side of equation. These are called **spectator ions**. There are therefore eliminated from equation.

The equation now becomes;

$$H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$$

This type of equation is called **ionic equation**. It involves the use of ions to form product.

Self-Assessment 12.5.

1. Name two uses of the following salts.
   Potassium nitrate.
   Calcium chloride.
2. State two methods of preparing:
   Soluble salts.
   Insoluble salts.
3. State two application of precipitation.
4. Define spectator ion.

12.16 Summary.

The acid is a proton donor while a base is a proton acceptor in the reaction according to lowrly brownsted theory. Their strength is determined by universal indicator, PH meter and conductivity apparatus.

Under this chapter, we have also looked at salt which are produced by neutralisation and precipitation processes. Salts can be categorised into soluble and insoluble.

The process of precipitation involves the reaction of soluble salts and insoluble salts. This is very useful in treatment of domestic wastes and industrial effluents.

12.17 Revision

1 a Define the following terms

I. Acid
II. Base
III. Conjugate acid
IV. Conjugate base

(a) From the following equations write down acid base pair.

i. $HCl_{(aq)} + NH_{(q)} \quad NH_{4+(aq)} + Cl^-$

ii. $CH_3COOH_{(aq)} + C_2H_5OH_{(l)} \quad CH_3COO_2H + H_2O$

(b) Explain how hydronium ion is formed?

2 a I Define “hydrolysis”

(ii) Explain why water is called amphoteric substances

b. what is PH

c. what is the difference between **the strength** and **concentration** of acids or base.

3. a Give two examples of the followings:

i. strong acids
ii. weak acids
iii. strong base
iv. weak base

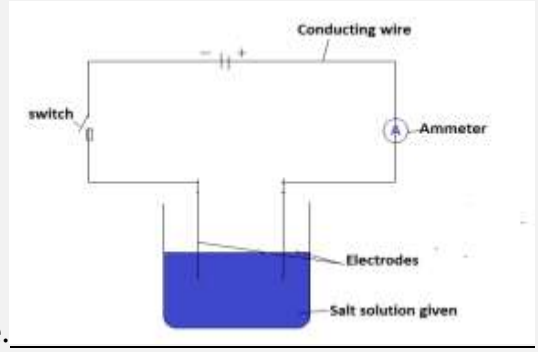
4. a. State three ways of expressing the strength of the acid/ bases.
   b. With the aid of a well labelled diagram, explain how can you determine hydroxide solution and ammonia solution.
   c. Explain how can you regulate PH in the following
   i. Stomach acid
   ii. Insects bits
   iii. Soil
5. a. Define "oxides"
   b. Mention 3 classification of oxides and two examples on each
   c. State two main methods of preparing salts
   d. I Define double decomposition precipitation
   ii. State two application of precipitation reaction.
6. a Define "ionic equation"
   b. write down ionic equation from the reaction.

$NaOH_{(aq)} + HCl_{(aq)}$ Nacl $+ H_2O$

   c. define "spectator ions"
7. Describe an experiment on how you can prepare insoluble salts using precipitation.
8. I Define "neutralisation process"
   ii. Explain how neutralisation can be used to prepare salts.

12.8 Practical question.

You are provided with beaker, 2 cells in holder, 100 ml of Sodium hydroxide solution, 100ml of dilute Sulphuric acid solution, 100ml of acetic acid, 100l of ammonia solution, carbon electrode, connecting wires and ammeter. Use the in procedures to determine the strength of acid and base.



Procedures

(a) set up experiment (figure)
(b) pour 75cm^3 of sulphuric acid in the beaker
(c) close the switch
(d) note the ammeter reading/ state of the bulb
(e) record in the table of results
(f) repeat the procedures with other solution, sodium hydroxide, acetic acid and ammonia solution

Table of results.

Solution	Ammeter reading	State of the bulb
Sulphuric acid		
Sodium hydroxide		
Acetic acid		
Ammonia solution		

1. Classify the acids and bases as:
   Strong acids.
   Weak acids.
   Strong bases.
   Weak bases.
2. Give two sources of errors.
3. Explain how errors can be reduced in an experiment?

UNIT 13

Oxidation and Reduction reactions.

Learning objectives

By the end of this unit, you should be able to;

a) Define the terms **"oxidation"** and **"reduction"** reactions.
b) Assign the oxidation number to an atom, ion and compound.
c) Identify the reducing and oxidizing agents.
d) Write half and overall redox equations.
e) Deduce the order of the reactivity of the metals.
f) Predict spontaneity of a reaction.
g) Describe the corrosion.
h) State the condition necessary for corrosion.
i) Describe the ways of preventing rusting.

13.10 Introduction.

Under organic chemistry, you have already applied some aspects of oxidation and reduction reaction. In this unit, you are going to study oxidation and reduction reactions that are applied in electrochemical cells, corrosion and rusting. You will also use their theories to arrange metals in order of their increasing reactivity in electrochemical series.

13.11 Oxidation number.

Oxidation number refers to the real charge of an atom in monoatomic ion or hypothetical charge in poly atomic ion.

Rules of assigning oxidation numbers.

1. Elements in their elemental form (uncombined state) have an oxidation number of 0**.** This is because all elements are made up neutral atoms that always have a final change of zero.
2. The oxidation of number of a polyatomic or monoatomic ion is equal to the charge of the ion
3. Nonmetals tend to have negative oxidation numbers, although some are positive in certain compounds or ions for example,
   - **Oxygen (O)** has an oxidation number of **-2,** except in the ***Peroxide ion*** in which it has an oxidation number **of -1**
   - **Hydrogen (H) is -1** when bonded to a metal**, H is +1** when bonded to a nonmetal.
   - **Chlorine** has oxidation number of **-1**

- These elements can be used as reference Points in assigning oxidation number to other elements.

4 The sum of the oxidation numbers in a neutral compound is zero (0)

5 For all neutral molecules, the sum of the separate Charges inside the molecule must add up to zero.

Example 1

Work out the oxidation number (charge) of the following of S in SO_2

Solution.

To work out the oxidation number of S in SO_2.

Charge of one oxygen atom is -2

Let **X** be oxidation number of **S**

X + 2O=0 *(oxidation number of neutral compound is 0)*

X+2(-2)=0

X-4=0

x=4

The oxidation number of S in SO2 is +4

Example 2.

Work out the oxidation number of Mn in MnO_4

To find oxidation number of Mn in MnO_4

- The charge for one oxygen aton is -2
- Let **X** be oxidation number for **Mn**.
- X+4(O)=-1
- X+4 (-2)=-1
- X- 8=-1
- X=-1+8
- X=+7
- Oxidation number of **Mn** in MnO_4 is +7.

Example 3

Work out the oxidation number of S in SO_4^{-2}

To find the oxidation number of S in SO_4^{2-}

The charge for one oxygen atom=-2

Let **X** be oxidation number of S

X+4(-2)=-2 *(oxidation number equals charge of an ion rule No.2)*

X-8=-2

X=-2+8

X=6

The oxidation of S in SO_4-2 is +6.

Example 4.

Find the charge of Mg in $MgCl_2$

To find oxidation number of Mg in $MgCl_2$

The charge of chloride atom =-1

.Let X be oxidation number of Mg.

X + 2(Cl)=0 *(oxidation number of neutral compound is 0)*

X+2(-1)=0

X-2=0

X=+2

Oxidation number of Mg in $MgCl_2$ is +2.

Example 5

Work out the oxidation number of Na in NaCl.

To find oxidation number of Na in NaCl

One chlorine atom has an oxidation number of -1

Let X be oxidation number of Na

X +1(Cl)=0 *(oxidation number of neutral compound is 0)*

.X+1(-1)=0

X-1=0

X=+1

The oxidation number of Na in NaCl is +1

Example 6

Calculate the oxidation number in N in NO_3

To find oxidation number of N in NO_3^{-1}

One oxygen atom has charge of -2

Let X be the oxidation number of N.

X+3(O)=-1*(oxidation number equal its charge)*

X =3(-2)=-1

x-6=-1+6

x=5

The oxidation number of N in NO_3^{-1} is +5.

Example 7.

Work out the oxidation number of S in H_2SO_4

To find oxidation number of S in H_2SO_4

- One Oxygen atom has a charge of -2
- Let X be oxidation number of S in H_2SO_4
- X +2(H) +4(O)=0*(oxidation number of neutral compound is 0)*
- X+2(1)+4(-2)=0
- X+2-8=0
- X-6=0
- X=6
- The oxidation number of S is 6.

Self-Assessment 13.1

1.Define the term oxidation number.

2.Given that the oxidation number of Oxygen atom is -2. Calculate the oxidation number of Mg in MgO.

13.12 Electron transfer reaction.

Oxidation.

- The loss of electrons
- The addition of oxygen atoms
- The increase in the charge or oxidation number
- The loss of hydrogen atoms

Reduction.

- The gaining of electrons
- The loss of oxygen
- The decrease in the charge or oxidation number.
- The gaining of hydrogen atoms

Oxidising agent is the substance that receives electrons in a redox reaction. This facilitates oxidation. While **Reducing agent** is the substance that has lost or donated electrons in a redox reaction.

Redox reaction.

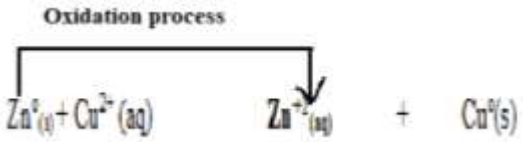
This is the reaction in which oxidation occurs together with reduction. The term **redo**x is an abbreviation of *reduction and oxidation*. The Oxidation and Reduction take part in the same chemical Reaction. If one substance is **oxidized** in the reaction, then the other substance must be **reduced** in the same reaction.

We can use oxidation number to determine if a substance has been oxidized or reduced. For example, let us consider the equation below:

$$Zn^{o}_{(s)} + Cu^{2+}(aq) \longrightarrow Zn^{+2}_{(aq)} + Cu^{o}(s)$$

In the reactants, $\mathbf{Zn^{o}_{(s)}}$means that it is neutral element while Cu^{2+} means that it is unstable, hence, it requires 2 elements to become stable(neutral) with the charge of zero ($Cu^{o}_{(s)}$).

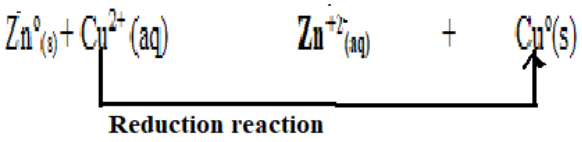
Zinc charges from 0-2 i.e from $\mathbf{Zn^{0}} \longrightarrow \mathbf{Zn^{2}+}$.Its oxidation number has increased. By changing from Zn^{0} to Zn^{2+}, zinc has lost electrons. This is called **oxidation**. In short



When Zn^{0} (zinc) metal Ionized, it changes its state to aqueous form. Thus, zinc ion can be written as $\mathbf{Zn^{2+}(aq)}$ in the products. $\mathbf{Zn^{2+}(aq)}$ means that zinc metal has lost two electrons. This oxidation can be represented by **half equation of oxidation**. i.e.

$\mathbf{Zn_{(s)} \longrightarrow Zn^{+2}_{(aq)}}$

Copper changes from 2 to 0 i.e. Cu^{+2} to Cu^{0}. The oxidation number has decreased. By changing from Cu^{2+} to Cu^{0}, copper ion (Cu^{2+}) has gained electron lost by zinc. This is called **Reduction process**.



When copper ion $\mathbf{(CU^{2+})}$ has gained electrons (+2e) lost by zinc, it becomes a neutral atom Cu^{0}(s). Therefore, it changes its state to solid form. This can be presented by **half equation of reduction.**

$Cu^{+2}{}_{(aq)} + 2e^{-} \longrightarrow Cu^{0}{}_{(s)}$

In the equation, Zinc metal (Zn^{0}s) is being oxidized. Therefore, it is called ***Reducing agent***.

While copper ion $Cu^{2+}{}_{(aq)}$ has been reduced by gaining electrons lost by Zinc hence it is called **Oxidizing agent.** Reducing agent get their oxidation number increased while oxidizing agent get their oxidation number decreased:

Cu^{2+} to Cu^{0}(s).

This can be summarized in the equation:

Oxidized reduced

$Zn^{o}_{(s)} + Cu^{2+}(aq) \longrightarrow Zn^{+2}_{(aq)} + Cu^{o}(s)$

Reducing agent Oxidizing agent.

Half equation of oxidation

$Zn_{(s)} \longrightarrow Zn^{+2}_{(aq)}$

Half equation of Reduction

$Cu^{+2}_{(aq)} + 2e^{-} \longrightarrow Cu^{0}_{(s)}$

Example 8

Study the equation below and answer the given questions:

$Mg^{0}(s)+2Ag^{+}(ag)……..Mg^{2+}(aq)+2Ag^{0}(s)$.

Which substance has been oxidized?

Magnesium $Mg^{0}(s)$. *why?* It has lost electrons.

Which substance has been reduced?

Silver ion($2Ag^{+}(ag)$). *Why?* Has gained electrons.

Which substance is reducing agent?

Magnesium($Mg^{0}(s)$) *Why?* It has oxidised

Which substance is oxidising agent?

Silver ion ($2Ag^{+}(ag)$) Why? It has been reduced.

Is this redox Reaction? Explain why?

Yes. The oxidation and Reduction is taking place simultaneously.

Write half equation for oxidation and Reduction.

- Half equation for oxidation
- **$Mg^{0}(s) \longrightarrow Mg^{2+}(ag)+2e^{-}$**
- (ii) Half equation for Reduction,
- **$2Ag^{+}(ag)+2e^{-} \longrightarrow 2Ag^{0}(s)$,**

Example 9

$CuO_{(s)}+C_{(s)}$……$CO_2(g)+Cu(s)$,

which substance is oxidized?

Carbon. ***Why?*** Has gained oxygen atom.

Which substance is reduced?

Copper. oxidized (CuO) **Why?** Has lost oxygen(O) atom.

Self-Assessment 13.2

1.Define the terms:

Oxidation

Reduction.

2.Study the equation below and answer the given questions:

$Ca^0(s)+2Ag^+(ag)$……..$Ca^{2+}(aq)+2Ag^0(s)$.

i. Which substance has been oxidized?

ii. Which substance has been reduced?

iii. Which substance is reducing agent?

iv. Which substance is oxidising agent?

v. Is this redox Reaction? Explain why?

vi. Write half equation for oxidation and Reduction.

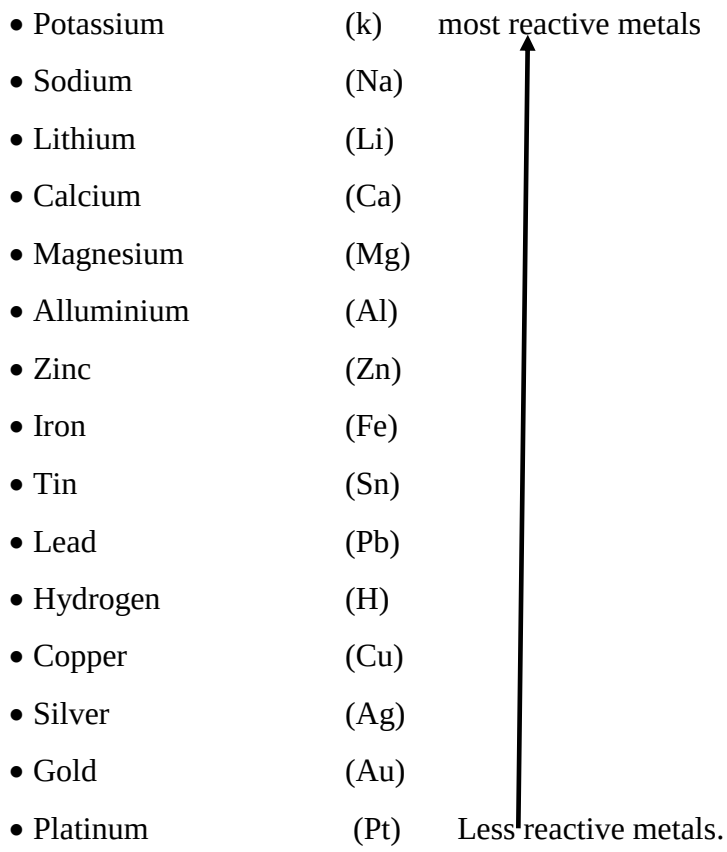
3.Describe the difference between reducing agent and oxidizing agent.

13.13 Displacement reaction.

Apart from halogens, displacement reaction also involves reaction of metals with other metals in aqueous solution. The metal reacts by donating or losing electrons in their outer energy level. The electrons lost are gained by metal ions in solution. Therefore, the metal donating the electrons is oxidized while the metal in solution gaining the electrons is reduced. Displacement reaction can be considered as **redox reaction**. The metal losing electrons is the ***reducing agent*** and the substance gaining the electrons is the ***oxidizing agent.*** Therefore, **displacement reactions of metals**, are reactions in which electrons are transferred from more reactive metal to less reactive metals.

The displacement series/reactivity series.

This shows metal in order of their increasing reactivity. The displacement or Reactivity series is also called **Electro Chemical Series**.



The figure 13.13.The Reactivity Series of metals.

The metal which are higher in the series will donate electrons to metals which are lower in the series. Therefore, a chemical reaction will take place.

For example: $\mathbf{Zn^0(s)+Cu^{2+}(ag) \longrightarrow Zn^{2+}(ag)+Cu^0(s)}$

Here the reaction has taken place because Zincn(Zn) is more reactive than copper (Cu), hence it donates electrons to copper(Cu^{2+}) ion.

Metals which are lower in the reactivity series will not donate electrons to metals higher in the reactivity series hence no chemical reaction. For instance;

$Mg^{+2}{}_{(aq)} + Cu^0{}_{(s)} \longrightarrow$ No reaction

Copper substance is less reactive than magnesium hence it cannot donate electrons to magnesium.

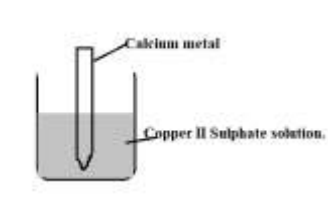
The combination of metal which are far apart in the reactivity series would give higher voltage in the voltage series (Electrochemical series). This is because, the metals far apart give large differences of a half cell voltages and therefore larger electromotive force (EMF) values.

Potassium is most electropositive element because it easily gives away its outermost shell electrons as compared to other metals in the series. **Why?** Potassium (K) has largest atomic radius. Therefore, its single electron in the outermost shell experience less force of attraction from the nucleus. As such, it is easily lost unlike in all other metals.

In **spontaneity of a reaction,** a displacement series can be used to predict whether reaction will take place or not.

Example 10

Study the figure below and answer the given questions,



(a) Explain what would happen after 10 minutes?

***Calcium(Ca) would displace Copper from the solution and produces Calcium sulphate and Copper metal.** *WHY?* Calcium is the most reactive metal than Copper.

(b) Write the balanced equation of the reaction:

$Ca^0(s) + Cu^{2+}SO_4$-2(ag)……$Ca^{2+}SO_4(aq)+Cu^0(s)$,

This equation can simply written as

$Ca^0(s) + Cu^{2+}(aq)$ $Ca^{2+}(ag) + Cu^0(s)$

(c)Which substance has been oxidized?

Calcium. WHY? **Has lost electrons**

(d) which substance has been reduced?

Copper ions.WHY? **Has gained electrons**

(e)*Write down half equation for oxidation and Reduction*

(i)Half equation for oxidation

$Ca^0(s) \longrightarrow Ca^{2+}(ag)\ \ 2e^-$

(ii)Half equation for Reduction,

$Cu^{2+} + 2e^- \longrightarrow Cu^0(s)$

Experiment 13.1

Aim **:** To carry out investigation on displacement series using different solutions and metals.

Materials:

- Four Beakers, Sand Paper
- Distilled Water, Copper Sulphate Solution
- Zinc Sulphate Solution, Iron Sulphate Solution
- Magnesium Sulphate Solution, Measuring Cylinder.
- Magnesium, copper, zinc, iron metal.

Procedures

a. Pour $5cm^3$ of Copper Sulphate Solution into each of the four beakers

b. Clean the Copper, Zinc, Iron and Magnesium metal using Sand Paper.

c. Put a piece of each Metal into each of the four beakers containing Copper Sulphate solution

d. Observe the contents of the beaker for 2 to 3 minutes.

e. Record the Result in the table below by indicating **"Reaction**" or "**No reaction**".

f. Rise the beakers with distilled water

g. Repeat the steps **a to** f using solutions of Zinc Sulphate, Magnesium Sulphate and iron sulphate respectively.

Table of results.

Metal Solution	Copper (Cu)	Zinc (Zn)	Iron (Fe)	Magnesium (Mg)
Copper Sulphate		Reaction	Reaction	Reaction
Zinc Sulphate	No Reaction		No reaction	Reaction
Iron Sulphate	No Reaction	Reaction		Reaction
Magnesium Sulphate	No Reaction	No Reaction	No reaction	

1(a) Use the results to arrange the metals in order of increasing reactivity.

- Magnesium (Mg
- Zinc (Zn)
- Iron (Fe)
- Copper (Cu)

(b)Explain the reason to your answer in 1a.

- *The reducing power of metals decreases down the reactivity series hence they above metals displace metals below them.*

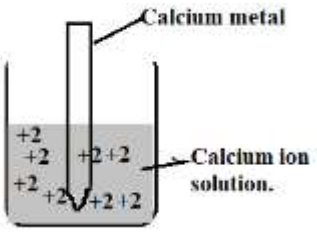
(c) From your observation, which combination of metals could give a higher voltage and state the reason for your answer stated.

Magnesium (Mg) and copper (Cu) would give a higher voltage because metals which are far apart in the voltage series (Electrochemical) would give large differences of a half cell voltage and therefore larger electromotive force (EMFs) values.

Self-Assessment 13.3

13.14 The electrochemical cell.

In this, Reduction-oxidation reactions are used to generate an electric current. In electrochemical cell, the reactive metal is dipped into an electrolyte containing the same ion. **An electrolyte** *is the substance that conduct electricity when in the molten state or solution*. In this electricity is carried through **ions**. The process of splitting up (decomposing) substances by passing an electric current is called **electrolysis**. However, you will study **electrolysis** as a special topic after this unit.



The figure 13.14. Explain partially what happens in electrochemical cell before.

Calcium metal (Ca) has been dipped in an electrolyte containing Ca^{2+} ions of the same metal. In this case, Calcium metal will lose +2 electrons. This can be summarized by **half equation of oxidation**

$$\mathbf{Ca^0(s) \longrightarrow Ca^{2+} + 2e^-}$$

The electrons lost go into the solution as positive ions. When electrons accumulate on the Calcium

Metal surface, it becomes negatively charged. The Calcium ions (Cu^{2+}) in the solution increase until some start to gain electron and form atoms. This overall can be summarized by half equation of Reduction.

$Ca^{2+}(ag) + 2e^{-} \longrightarrow Ca^{0}(s)$

These two reactions of Oxidation and Reduction take place at the same time at the

Equal rate. Therefore, it is called **Redox Reaction**.

The ability of a metal to lose electrons depend on their position in the electrochemical series. The metal that loses electrons easily, will acquire a much negative charges while the metal which loses electrons slowly will have smaller(few) negative charge on its surface.

When the metal i.e. Calcium metal becomes negatively charged, it attracts a layer of positive ions, of which forms two layers.

This result in **potential difference** between metal (**Calcium**) and its ions produced. This is called

electrode potential of the metal. Therefore, potential difference depends on the greater the ability of the metal to form ions. Hence, the greater the ability of the metal to ionize, the greater the potential deference (**electrode potential**). It is impossible to measure the electrode potential of one electrode but it is possible with two electrodes. This is easily done by connecting the two electrodes with **metal wire** and connecting the two solutions by **Salt bridge.**

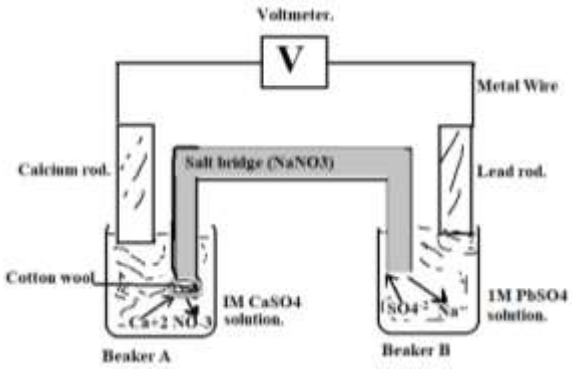
This should be in a way that ions can slow between them.

In this case above where the metal rod of Calcium was dipped in a salt solution of Calcium Sulphate, it is Called a **half - cell.**

Two half – cells connected through the two electrodes is what we call **electrochemical** or **Voltaic cell.**

Let us use the simple diagram below to see how redox reaction helps to produce current in

Electrochemical cell.



In the beaker A, the Calcium rod dissolves by losing two electrons to form Calcium ions (Ca^{2+}(aq).This is called **Oxidation Process.** It is summarised by half equation below:

$Ca_{(s} \longrightarrow Ca^{2+}(ag) + 2e^-$

The electrons ($2e^-$) flow through the external wire to the beaker labeled B. The lead ion ($Pb^{2}_{+(aq)}$) will react with these electrons to form lead metal deposits. This is called **Reduction Process.** Therefore, it can be summarised by half equation

$Pb^{2+} + 2e^- \longrightarrow Pb(s).$

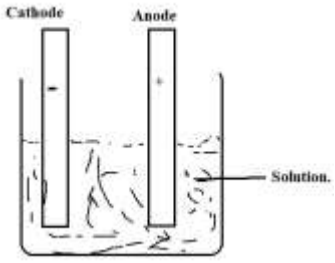
The overall equation:

$Ca_{(s)} + Pb^{2+}$(aq)……Ca^{2+} (aq) + $Pb_{(s)}$

Normally, **Oxidation** always occurs at the **Anode** (beaker A) while **Reduction** Process occured in **Cathode** (beaker B) of the Electrochemical cell. **The anode,** is the positive electrode which attracts **anions** (**negatively charged ions**).

The Cathode, is the negative electrode which attracts **Cations** (**positively charged ions**).

The **Anode** and **Cathode** can be presented in the figure below:



The electrons are transported by the ions (anions) in the solution up to the electrode and through the Wire to the other electrode. **In the salt bridge**, different concentration in Na^{+} and $NO3^{-}$, and Ca^{2+} and $SO4^{-2}$ allows Ca^{2+} and $SO4^{-2}$ to diffuse in while Na^{+} and $NO3^{-}$ are diffusing out.

The potential difference of the cell is measured by **Voltmeter in Volts**. This gives the Voltage of the cell called **electromotive force** (emf).

Mathematically.

Electromotive force of cell = Electromotive force of anode – Electromotive force of Cathode.

emf of the cell=emf of Anode(+) – emf of Cathode.

The Electromotive force **(potential difference)** is also affected by:

- Temperature
- Concentration
- The state of the metal.

The Electromotive force can be simply symbolized by **E**.As times goes ,Calcium metal becomes thinner as it ionizes while lead electrode becomes thinker and thinker as it gain electrons. When lead sulphate($PbSO_4$) is used up by the second half-cell, solution ofNat ions from the salt bridge flow to balance the increased negative charge at the Cathode. Therefore, the concentration of Ca^{2+} ions at the anode increases while of Pb^{2+} ion at the Cathode decreases.

Self-Assessment 13.4

1. With the Aid of well labelled diagram, describe how electrochemical cell produces the current.
2. Define electromotive force.
3. Describe how can you find the electromotive force.
4. Name three factors that affects electromotive force.

13.15 Corrosion/rusting.

Corrosion is the reaction process which occurs when some metals are exposed to air(oxygen) and water. In this process, the metal is dissolved (ionized) slowly and gradually wear out. Most metals on corrosion form **oxides**. The good example of corrosion is (**Rusting).** Therefore, **rusting is** *the process whereby Ion react with water in the presence of oxygen.*

Conditions necessary for Rusting.

- Presence of iron.
- Presence of Oxygen.
- Presence of Water (moisture).

Experiment 13.2

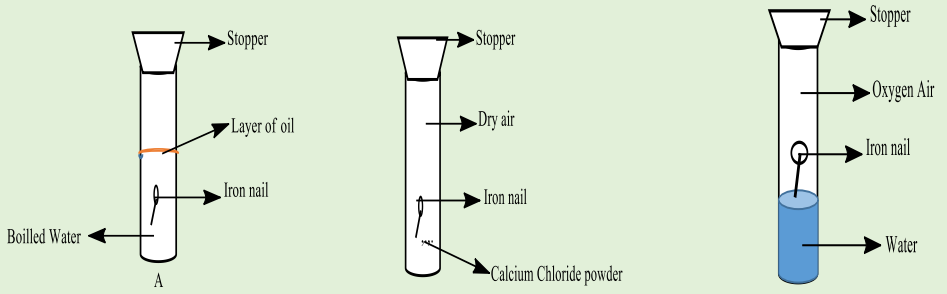
AIM: *To investigate Conditions necessary for rusting of Ion*

Materials

- test tubes
- Ion nails
- Ionic stoppers
- Tap water
- Boiled water
- Anhydrous calcium chloride.

Procedures

a. Set up experiment as shown below:



The figure 13.15a. Iron nail in three different medium.

(b).Leave the experiment for one week.

(c).Observe what happens to each nail.

Results/observation.

In the test tube A, rust will not take place because the air was removed by boiling the water.Test tube B, rusting will not take place because the *calcium chloride* absorbs water (It is dehydrating agent). In test tube C, the nail will rust because of the **Air (Oxygen)** and **Water**.

How is rust formed?

Iron metal (Fe) must dissolve first in water and ionize (equation 1):

$$Fe^{0}_{(s)} \longrightarrow Fe^{+2}_{(aq)} + 2e^{-} \qquad 1$$

The electrons released are used by oxygen gas and water (equation 2)

$$2H_2O_{(l)} + O_{2(g)} + 4e^{-} \longrightarrow 4OH^{-}_{(aq)} \qquad 2$$

The hydroxyl ions react with Fe^{2+} ions and more water and air to give rust which is hydrated Iron III oxide (equation 3).

$$2Fe(OH)_{2(aq)} + 2H_2O_{(l)} + O_{2(g)} \longrightarrow Fe_2O_3.\, n\, H_2O_{(s)} \quad 3$$

n stands for number of water molecules which can vary.Corrosion increases when the Ph. value is very low i.e. when the conditions are acidic.

Ways of preventing corrosion/rusting.

Painting.

This prevents water from reaching inside which is necessary conditions for rusting

Applying oil or Grease on substance.

This prevents both water and oxygen gas from reaching Iron or any metal inside.

Electroplating the substance.

This prevent oxygen and water from reaching inside the substance.hence, no rust take place

Galvanization.

This means covering Iron with Zinc .Zinc does not rust easly.

How does Galvanization prevent Rusting?

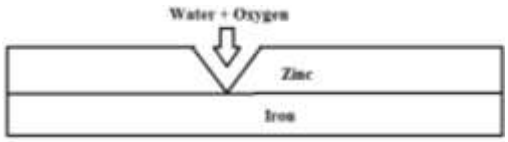


Figure 13.15b. Galvanization process.

When zinc is punctured, water and oxygen occupies the gap.This forces Iron(Fe) to ionize and become $Fe^{2+}(ag)$ ion ready to form rust. Immediately, zinc will donate electron to Fe^{2+} to avoid unnecessary rusting.

This is shown by half equation below:

$$Zn_{(s)} + Fe^{+2}{}_{(aq)} \longrightarrow Zn^{+2}{}_{(aq)} + Fe^{0}{}_{(s)}$$

Half equation for Oxidation

$$Zn_{(s)} \longrightarrow Zn^{+2}{}_{(aq)} + 2e^{-}$$

Half equation for Reduction

$$Fe^{+2}_{(aq)} + 2e^{-} \longrightarrow Fe^{0}_{(aq)}$$

In this case, Iron will remain solid hence no rusting.

Sacrificial protection

This protects underground Iron pipes from rusting. Zinc or Magnesium blocks are attached to them. Since they are more reactive, they corrode instead of Iron. This is called ***Sacrificial protection*** because zinc and magnesium are Sacrified to protect Iron from rusting.

Self-Assessment 13.5

1. Define corrosion.
2. Name three necessary conditions for rusting.
3. With the aid of well labeled diagram, describe the experiment to demonstrate necessary conditions for rusting.
4. With references of chemical equation, explain how rust is formed?
5. Describe three ways of preventing rusting.
6. With the aid of labeled diagram, explain how galvaniszation process prevent rusting.

13.6 Summary

Oxidation number number is a real charge in monoatomic and hypothetical charged in poly atomic ion. Oxidation involves loss of electron, gaining of oxygen atom, loss of hydrogen atom and increase in the charge of atoms, whereas Reduction is the opposite of oxidation.

When oxidation and Reduction is taking place at the same time, the term called **Redox Reaction** can be used. **Redox** means *Reduction* and *Oxidation.* These kinds of reaction are commonly used in electrochemical cell displacement series and rusting of metals. Rusting/corrosion occurs when Iron or any other metal is exposed to oxygen, and water. However, the condition can be prevented by painting the material, greasing or oiling, galvanizing and electroplating.

13.7 Revision Exercise.

1(a) Define the following terms in terms of electron:

(i)Oxidation

(ii)Reduction

(b) Explain the difference between oxidation and reaction in term of hydrogen atoms

(c) Define "redox reaction"

2 Study the equation below and answer the the given questions

$Zn(s) + Ca^{+2}(aq) \rightarrow Zn^{+2}_{(aq)} + Ca^{0}_{(s)}$

(a) What does 0 in calcium means

Which substance(s) has been

i. Oxidised
ii. Reduced
iii. Oxidising agent
iv. Reducing agent

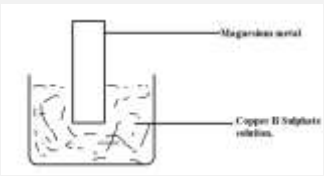
b,(i)Write down Half equation from oxidation and Reduction.

(ii)Is this redox reaction? Explain why?

c, Find the oxidation number of sulphur(s) in sulphulic acid(H_2SO4) if oxidation number of oxygen(O) is -2 and Hydrogen is (I).

(b) define oxidation number in poly atomic atoms.

3 Study the figure below and answer the question



a.(i)Explain what would happen after 5 minutes

(ii)Write down balanced equation for reactions.

b.Will this reaction take place fast? Explain why?

4a (i)Define “Corrosion”

(ii)State three necessary for rusting .

(iii)Explain how rust is formed?

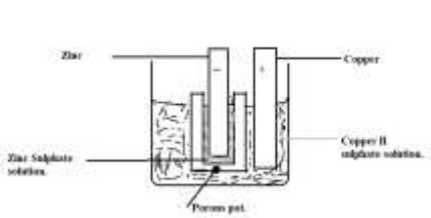
b with the aid of well labelled diagram explain how galvanization helps to prevent rusting.

5 a Define “Electromotive force”

b Draw the Electrochemical cell using calcium and lead metals and other two fast solution.

c Mention three ways of preventing rusting.

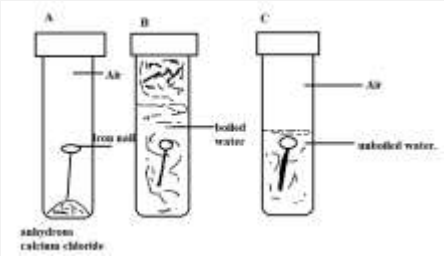
6. One of the first practical electrochemical cell was the Daniell cell inverted by John Daniel in 1836. A diagram of this type of cell is shown below.



It is capable of generating about 1.1 volts and was used to operate small electrical items such as doorbells.

a) The electron reaction taking place at a copper anode is
   a. Cu+2(aq) + 2e- Cu(s).
b) Write the electrode equation for the process taking place at the cathode.
c) Which way would the electrons flow in the wire connected to the voltmeter?
d) Why would the copper II sulphates crystalise at the bottom of the outer container?
e) What is the function of the porous pot?
f) There are problems associated with the Daniell cells which has led it to being replaced by other types of cells. State two reasons why Daniell cell are no longer in use today?

7. Three test tubes were set up as shown to investigate the rusting of iron. After one week the rusting was observed in tube C.



i. Why is anhydrous calcium chloride placed in tube A?
ii. Why is the water in tube B boiled?
iii. What would you expect to happen to the nail in tube B if the stopper was removed and the tube left for a further week?
iv. From the results of the above experiments, state the conditions necessary for the rusting of Iron.

13.8 Practical Question.

You are provided with the following materials, four beakers, sand paper, distilled water, copper sulphate solution, zinc sulphate solution, Iron ……sulphate solution and a measuring cylinder, copper metal, Iron, zinc and magnesium metals.

Procedures

(a) Put 5cm^3 copper sulphate solution into each of the four beakers.
(b) Clean the copper, zinc, Iron, magnesium metals using sand paper.
(c) Dip the metal into each of the four beakers.
(d) Observes the contents of the beaker after 5 minutes.
(e) Rise the beaker with distilled water.
(f) Repeat the steps a to f using the other given solution.

Table of result

Solution(s) **Metal(s)**	**Copper sulphate**	**Zinc suphate**	**Iron sulphate**	**Magnesium sulphate**
Copper				
Zink				
Iron				
Magnetfun				

(i).Fill the result by writing **Reaction** or no reaction.

(ii).Arrange the metal in order of their reactivity.

(iii).From your observation, which reaction was very fast and why?

(iv). State two sources of errors in the experiment.

(v). Describe two methods on how you can reduce errors mention.

UNIT 14.

Electrolysis

Learning objectives.

By the end of this unit, you should be able to:

a) Define "electrolysis"
b) State the role of water in electrolysis
c) Describe the purification of copper.
d) Describe the process of electroplating and its uses.
e) State the uses of electrolysis.

14.10 Introduction.

In chemistry and manufacturing company, **electrolysis** is a technique that uses a direct electric current (DC) to drive an otherwise non-spontaneous chemical reaction. Electrolysis is commercially important as a stage in the separation of elements from naturally occurring sources such as ores using an **electrolytic cell**. The voltage that is needed for electrolysis to occur is called the **decomposition potential.** This theory of electrolysis is very important in chemical industry for it is used in batteries.

In this unit, you are going into details of electrolysis and its application to purification of copper and electroplating process.

14.11 Electrolysis.

By definition, **electrolysis** *is the process of splitting up (decomposing) substances by passing an electric current through them*. The substance which is decomposed is called **the electrolyte.** Therefore, **the electrolyte** is the *moiten state substance or solution that conducts electricity.* The electricity is carried through the electrolyte by **ions**. This leaves the electrolyte through **electrodes. Electrodes** are ususlly made of *unreactive metal e. g platinum* and *non-metal carbon* (**inert electrodes).** Normally we have two electrodes, **Cathode** and **Anode**. The **Cathode** is the negative electrode which attract **Cations** (***positively charged ions***) while **Anode** is the positive electrode which attracts anions (***negatively charged ions***).

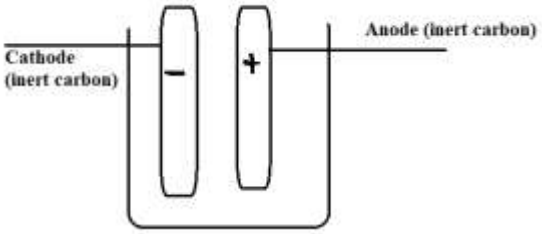


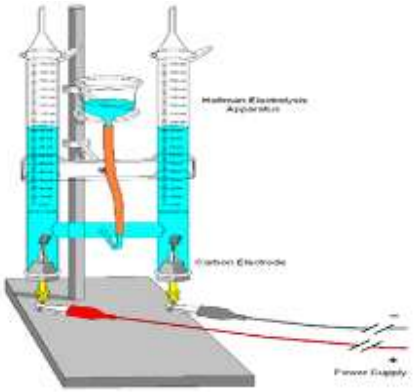
Figure 14.11b. The electrodes on electrolysis.

Electrolysis of aqueous solution.

The other industrial processes involve the Electrolysis of aqueous solutions. Therefore, let us consider the electrolysis of water.

Electrolysis of water (H_2O).

Pure water is a very poor conductor of ecletricity because there are few ions in it.The instrument that is used in electrolysis of water is called **Hofmann Voltameter**.



The figure 14.11b. Hofmann voltameter.

Sodium hydroxide (NaOH) solution or Sulphuric acid (H_2SO_4) which is diluted must be added to water to increase its conductivity. When the power supply is turned on, current start flowing through this solution. The gases (**oxygen and Hydrogen)** can be produced at the both electrode**,** Anode (+) and Cathode(-).These gases are collected in the side arm of the apparatus.After 20 minutes,roughly twice as much as gas is produced at the Cathode as at Anode. The gas collected at the cathode burns with a squeaky pop. This means that it is hydrogen gas (H_2). For the Hydrogen

gas to be collected, the positively charged hydrogen ions must have moved to the Cathode.

This is shown by chemical equation below:

$$\mathbf{4H^+(aq) + 4e^- \longrightarrow 2H_{2(g)}}$$

When water loses H^+ (Hydrogen ions), the remaining portion is Hydrogen ions (OH^-). This is attaracted to the Anode.The gas collected here, relights a glowing splint showing that it is **oxygen**.

The chemical equation below shows how oxygen gas is produced:

$$4OH^-_{(aq)} \longrightarrow 2H_2O_{(l)} + O_{2(g)} + 4e^-$$

This process/experiment proved that the formula of water is **(H_2O).**

The role of water in electrolysis.

This dissolves most ionic compounds because it is polar, therefore it attracts the positive and Ions.

When the water is partially Ionised, the ions are free to move at random. This makes water to be weak electrolyte. In electrolysis, the **H^+** are attracted to the cathode electrode to be reduced and **OH^-** are attracted to the anode electrode to oxidised. At each each elcecrode, only one type of ions gets discharged hence it is called **Preferential discharge.**

Factors that affecting preferential discharge.

- Position of water in reactivity series.
- Concentration of ions in solution.
- Nature of electrode.

Self-Assessment 14.1

1. Define the term:
   Electrolysis.
   Preferential discharge.
2. Describe the role of water in electrolysis.
3. Name three factors that affecting preferential discharge.

14.12 Purification of copper using electrolysis.

Copper is a very good conductor of electricity, it is used for electrical wiring and cables. Any amount of impurities cut down or reduce its conductivity. Therefore, the newly extracted copper has to be purified by electrolysis to make sure that it is 99.99% pure.

The impure copper is used as the anode and the pure copper is use as cathode. The electrolyte is a solution of copper sulphate acidified with a solution of sulphulic acid to improve its conductivity.

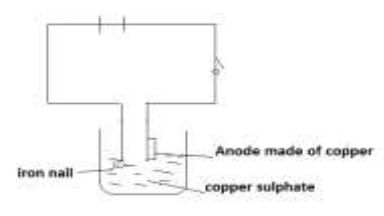
Experiment 14.1

AIM**:** Purification of copper using electrolysis.

Materials

- Two cells in order
- Beaker
- Copper…… Sulphate solution
- Sulphuric Acid solution
- Connecting wires.

The figure



Procedures.

(a)Set up experiment as shown in the diagram

(b)Close the switch

(c)Observe what happens on the anode and cathode after 10 minutes.

Table of results.

Electrode(s)	**Observation(s)**
Anode	Loses mass (bulbles of gas)
Cathode	Gain mass (bubles of gas

Discussion.

When the current is switched on, the copper moves from the impure to the pure Cathode.During the process, the impure anode loses mass because the copper atoms lose electrons and becomes ions **(**CU^{2+}(aq)**.**

This can be summarized by half equation below:

$Cu_{(aq)} \longrightarrow Cu^{+2}_{(aq)} + 2e^-$

The process is called **Oxidation**.It takes place at the Anode.The electrons released at the anode travel around the external circuit to the Cathode.The electrons are passed on the copper ions from the copper II sulphate solution and copper is deposited onto the Cathode.This causes increase in the mass of cathode. The reaction takes place at the cathode is called **Reduction reaction**.This can be summarized by half equation below:

Cu^{+2}(aq) +2e-$\longrightarrow$Cu(s)

In this process any impurities fall to the bottom of the cell and collected in the anode in form of the **slime.**

Conclusion.

Copper metal can be purified using electrolysis.

NB: The slime is rich in precious metals like gold and silver.

Self-Assessment 14.2

1. Define Purification of copper.
2. With the aid of well labeled diagram, explain how can you purify copper metal using electrolysis.

14.13 Electroplating.

The process involving electrolysis which Coats ametal with another metal using electric current. This also deposits metals from solution in form of thin layer on the other suface.

Reasons for electroplating.

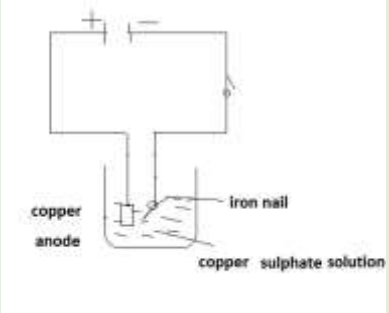
- To prevent rusting.
- To make materials look shiny and attractive.

Experiment 14.2

AIM: Electroplating of Iron nail with copper using redox reduction.

Materials:

- Iron nail
- beaker
- Two cells in holder
- Copper electrode(Anode)
- Copper II suilphate solution.
- Connecting wires
- switch



Procedures

(a) Set up the apparatus as shown above.

(b) close the switch.

(c)Observe what happens

Table Observation/Results

Iron nail	Observation (s)
Mass	
Colour	
Mass of anode	

Discussion.

When the switch is closed, electric current will start flowing and the copper on the anode will dissolve and ionise. The anode reaction is as follows

$$Cu_{(aq)} \longrightarrow Cu^{+2}_{(aq)} + 2e^-$$

This reaction is called **Oxidation.**

The copper ions will be attracted to the cathode. The electron lost by copper will travel through the external cilcut to the cathode where they will be gained by copper **($CU^{2+}aq$)**.

The cathode reaction is called **Reduction.** This can be summarized by half equation below:

$$Cu^{+2}(aq) + 2e^- \longrightarrow Cu(s)$$

The reacton shows that copper will be doposited on the ion nail.

Conclusion.

Iron nail can be electroplanted with copper using Oxidation-Reduction reaction in electrolysis.

Applications of electroplating.

- Jewerly industry
- Manufacture of computer hard drives (Electrical components)
- Making engine parts
- Corrosion protection
- Automotive application e.g tyre rims.

14. 14 Uses of electrolysis.

- Electroplating of metals
- Purification of metals (refining).
- Extraction of metals.

Self-Assessment 14.3

1. Define electroplating.
2. State two reasons for electroplating.
3. Describe how you can electroplate iron nail using copper II sulphates solution.
4. Describe three application of electroplating.
5. What are three uses of electrolysis?

14.15 Summary.

The electrolysis is aprocess in which a chemical reaction is caused by the passage of an electric current. In this process non-metals are produnced at the anode whereas metals and hydrogen gas are procuced at the cathode. At the anode, chlorine, bromide and Iodine are produced in preference to oxgygen. While at the cathode, hydrogen is produced in preference to metals unless unreactive metals such as copper and nickel are present. The theory is used in purificationof copper and electroplanting process.

Electroplanting is the process of depositing metals from solution in the form of a thin layer on other surfaces such as metal or plastic. The theory of electroplanting is used to make metals look shiny and attractive. This also helps to prevent rusting.Electrolysis involves the use of both Oxidation at the anode and Reduction at the Cathode.

14.16 Revision Exercise.

QUESTIONS

1 a. Define the following terms:

(i) Electrolysis

(ii) Electrolyte

(iii) Cathode

(iv) Anode

(b) Explain the role of water in electrolysis

2(a) Define “Preferential discharge”

(b) Metion two factor that affect Preferential discharge

(c).State three uses of electrolysis.

3(a) Define Electroplating.

(b) State three importance of electroplating.

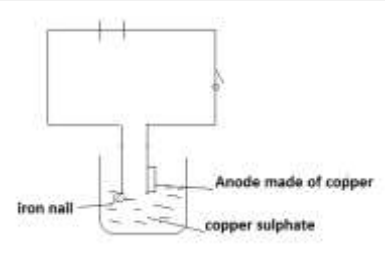
(c) Explain why electrolysis is used to prevent rusting based on your knowledge.

4 With the Aid of a well labeled diagram, explain how you can purify copper using electrolysis.

5 With the Aid of a well labeled diagram, explain how you can electroplate silver spoon using copper II Sulphate solutions.

14.7 Practical Question.

You are provided with the following materials, two cells, beaker, nails, copper II sulphate solution, copper electrodes(anode),connecting wires and switch.



(a)Connect the circuit (Figure)

(b) close the switch

(c)observe what happens

Table of results

Iron nail	**Observation**
Mass of mode	
Colour of pron nail	
Olmer observation	
In pron nail	

Write down half equation in:

(i)Iron nail

(ii)Anode

UNIT 15.

Water.

Learning objectives.

By the end of this topic, you should be able to:

a) Identify natural sources of water.
b) Decribe the water cycle.
c) Describe the physical properties of water.
d) Explain the importance of water.
e) Describe water hardness and its effects.
f) Describe methods used to remove water hardness.
g) Describe sources, effects and prevention of water pollution.
h) State impact of human activities on the atmosphere.

15.10 Introduction

Water is very important natural resource that is widely used in different process. In living organisms, water aid various biochemical process that helps to produce energy for the body activities. In the previous section, you have seen how water helps various kind of chemical reactions. In this unit, you are going to study water only as outlined by learning objectives.

15.11 Natural resources of water.

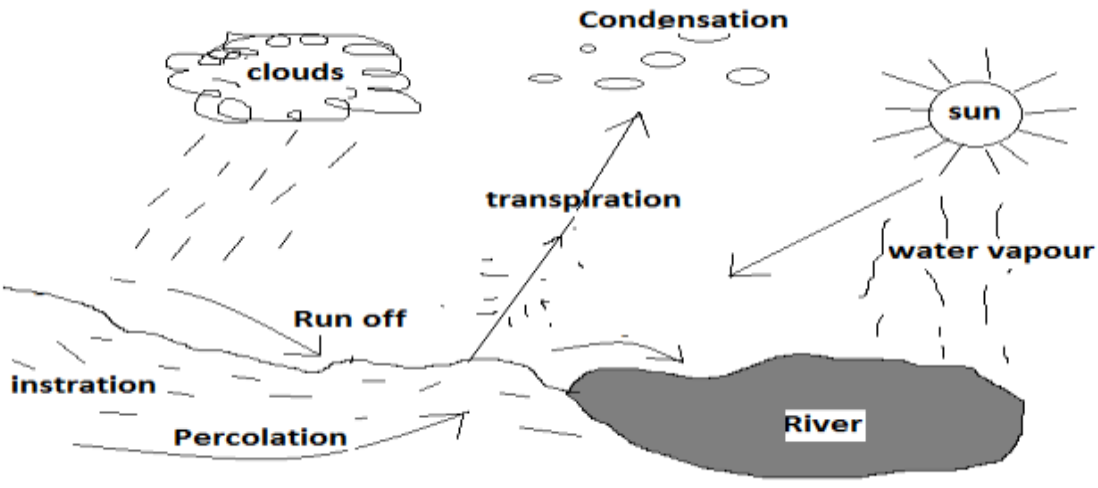
- Wells
- Rivers
- Seas
- Dams
- Lakes
- Ponds
- Rain
- Spring Water
- Borehore.

15.12 Water cycle.

This cycle shows how water circulates around the Earth. Water cycle is also called **hydrological cycle.** The driving force behind the water cycle is the **sun**. The sun provide heat energy.

Steps in water cycle (Hydrological cycle).

- Heat from the sun causes **Evapolation** from sea, lakes, rivers, and oceans in form of water vapour.
- This water vapour also comes from plants leaves through **Transpiration** and **Respiration.**
- When water vapours rises, it cools and condense to form tiny droplets of water through the process called **Condensation.**
- These tiny droplets of water vapour form **clouds** which are moved by air currents.
- When they reach certain size, they fall as rain.
- The water that falls as rain, runs into streams and rivers and then on into lakes, seas and oceans.
- The process in which heavy clouds release moisture back to the earth's surface in form of rain, hailstorm and snow is called **Precipiration.**
- Therefore, some water infiltrate into the soil through **infiltration Process**, where it moves through the soil and its layers by gravity and Cappilarity forces through **Percolation.**



The figure 15.12. Water cycle.

Importance of water cycle.

- Ensures constant supply of fresh water in lakes and rivers
- Encourages soil formation through leaching process.
- Provide water used by plants and animals.
- Maintains a balance between the water on the surface, underground and in the atmosphere.

15.13 Physical properties of water.

- Colourles
- Tastless
- clourless
- Has boiling point of 100^{o}C and melting point of 0^{o}C.
- has the density of 1g/cm^{3}.

15.14 Importance of water in everyday life.

- Plants use water for transportation of nutrients.
- Domestics uses e.g cooking, drinking and washing.
- in agriculture, water is used for irrigation of crops
- generation of electricity e.g Mkula falls
- In laboratories, water is used as solvent. It dissolves polac substances.
- In automobile, water is used for cooling engines.

Self-Assessment 15.1

1. Name three natural sources of water.
2. Define water cycle.
3. Draw the water cycle.
4. Differentiate transpiration and evaporation in water cycle.
5. Explain why water cycle is importance?
6. Name three physical properties of water.
7. What are the three importance of water?

15.15 Hardness in water.

Rain water dissolves carbon dioxide as it falls through the atmosphere. A small fraction of this dissolved carbon dioxide with water to form carbonic acid.

For example:

Water + Carbon dioxide = carbonic acid (double arrows)

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

Carbonic acid is a weak acid.

When carbonic acid (H_2CO_3) passes over the rocks containing limestone and dolomite, it dissolves slowly. The dissolved substance are called **calcium and magnesium hydrogen carbonates.** The presence of this dissolved substances in water causes **Water hardness.**

Therefore, **Water hardness** does not react with soap (soap is insoluble in it) hence it forms **grey-white suspension** which is called **scum**. By definition **water hardness** is *a kind of water that does not form lather easily with soap*. The good example is **tap water** this has salt elements dissolved in it. The opposite of water hardness is **soft water**. This form lather easily with soap. The good examples are **distilled water** and **rain water**.

Importance of soft water.

- Used as a solvent in laboratory.
- Used for cleaning test tubes.
- Used for washing cloth.

Types of water hardness.

There are two types of water hardness.

(a) Temporary hardness.

In **temporary hardness**, water contains dissolved calcium and magnesium hydrogen carbonates. This causes carbon dioxide that forms carbonic acid. Temporary hardness is easily removed by bo**iling.**

b) Permanent hardness.

In permanent hardness, water contains Calcium Sulphates and Magnesium Sulphates. This cannot be easily removed by **boiling.** However, **chemical treatment** and **Distillation** can be used to remove permanent hardness in water. The same methods can also remove *Temporary hardness*.

Advantages of hard water.

- Development of healty teeth and bones.
- Formation of shells in animals.
- For the healthy growth of plants in agricultural industries.
- Helps to form lead carbonates coat in the water.
- Good for brewing beer.

Disadvantages of hard water.

- It is not good for washing
- Lowers the quality of fibers in cloths when you hence used it for washing.
- Blocks pipes used by water boardy system.
- Causes kettles to fur.

Self-Assessment 15.2

1. Describe the causes of water hardness.
2. Describe two types of water hardness.
3. Name three importance of soft water.
4. Describe three:
    - Advantages of hard water.
    - Disadvantages of hard water.

15.16 Methods of removing water hardness.

Several ways are used to remove water hardness.

Addition of washing soda ($Na_2CO_3.10H_2O$) crystals.

Calcium and Magnesium in which actually causes the hardness are removed as a precipitates. Hence no longer hardness. This is summarized by equation below:

Calcium ion + Carbonate ion=Calcium Carbonate.

$$\mathbf{Cu^{+2}(aq) + CO^{2-}{}_3(aq) \longrightarrow CaCO_3(s)}$$

Ion exchange.

This uses resin. Resin is made up of insoluble sodium salt called **Sodium Permuitit**. Therefore, the method is also called **permuitit Process**. In this, the water is passed through a container filled with a suitable resin containing Sodium ion (Na+). The calcium or Magnesium ions causing the hardness are exchanged for the sodium ion in the resin.

The equation below summarises the given information above:

Calcium + Sodium-resin ⟶ Calcium-resin sodium

$$\mathbf{Ca(s) + Na\text{-}R(s) \longrightarrow Ca^{2+}\text{-}R(s) + 2Na^{+}(aq)}$$

When all the sodium ions have been removed from the resin, it can be regenerated by pouring a solution of a suitable sodium salt through it.

Distillation.

The water is distilled away from the dissolved substance. However ,it is not suitable to be used on a large scale.

Boiling.

This decomposes both calcium and Magnesium hydrogen carbonates, forming insoluble calcium carbonate. For instance, see the equation below:

Calcium hydrogen carbonate calcium carbonate + water + carbondioxide

$$Ca(HCO_3)_{2(aq)} \xrightarrow{heat} CaCO_{3(s)} + H_2O_{(l)} + CO_{2(g)}$$

However, the method is not suitable for removing permanent hardwater.

Self-Assessment 15.3.

1. Explain how the following methods are used to remove water hardness.
   Ion exchange.
   Addition of washing soda.

15.17 Water polution.

In JC chemistry, you learnt about the pollution of air and soil. The addition of harmful substances into water as a result of human activities causes **water pollution**. The substances that contaminate water are called ***Water pollutants***.

Examples of water pollutants.

- Toxic metals
- Plastic wastes
- Pesticides.
- Fertilizers.

Causes of water pollution.

- Poor agricultural practice.
- Oil spills from Ships
- Industrial wastes
- Detergents from washing in homes and hospitals.
- Hot water from industry.
- Waste and sewage from homes.

Effects of water pollution.

- Promotes growth of disease-causing micro-organisms like typhod, cholera and dysentery.
- Promotes entrophication (enriching the lake with nutrients and its effects) which causes the death of aquatic animals.

- Hot water reduces oxygen concentration in water bodies hence it causes the death of fish.

Ways of controlling water pollution.

- Treatment of sewage.
- Avoid use of excess toxic chemical.
- Do not urinate into water bodies
- Use farm-yard manure not excess fertilizer.
- Construct pit latrine should be constracted away from water bodies
- Industries shoube advised not to rease not water into waterbodies.

Purification and treatment of water.

This process involves both **Filtration** and **Chlorination**.These follows steps below:

(1) Impure water is first passed through screens to filter out floating debris.

(2) Aluminium Sulphate is added to coagulate small particles of clay-so that they form larger clumps, which settle more rapidly.

(3) Filtration through coarse sand traps larger, insoluble particle. The sand also contains specially grown microbes which remove some of the bacteria.

(4) The sedimentation tank has chemicals known as flocculants, for instance, Aluminium is added to it to make the small particles stick together and sink to the bottom of the tank.

(5) These particles are removed by further filtration, through fine sand.

(6) Finally, a little chroline gas is added to kill the remaining bacteria. This sterilizes the water.

The addition of chlorine gas makes the water more acidic, hence appropriate sodium hydroside solution are added.

15.18 Impact of human activities on the atmosphere.

- Air pollution

Self-Assessment 15.5

1. Name three water pollutants.
2. Describe two causes of water pollution.
3. State three effects of water pollution.
4. Explain two ways how water pollution can be controlled?
5. List two impacts of human activities on water pollution.

15.19 Summary.

According to what we have studied, water is very important in our everyday life. Therefore, treatment of water is very ensential to ensure good health.

15.20 Revision Exercise.

1(a) Mention three natural sources of water.

(b)Define water cycles.

(c)Explain how water cycle is very important (3 points).

2 Complete the figure below:

(a) Identify the steps

- A_____ B_____
- C_____D_____
- E_____F_____
- G_____

(b)Explain the difference between stage **E** and **F**

(c)State thee physical properties of water.

(d)Explain why water is important.(3 points)

3(a)Define “water hardness”

(b)Distingush permanent and Temporary water hardness.

(c)State 3 advantages of water hardness.

(d)State 3 ways of removing water hardness

4(a)Define water pollution.

(b)Mention 3 examples of water pollutants

(c)Explain how oil spills from ships cause water pollution on the lake.

(d)State three effects of water pollution

5(a)Mention three causes of water pollutions.

(b)State three ways of controlling water pollution.

6 State two main process in water purification and treatment.

7. Explain the process of water treatment and purification

With the aid of a well labelled diagram, describe the water cycle **(Hydrological cycle).**

UNIT 16

Greenhouse gases and ozone layer.

Learning Objectives.

By the end of this unit, you should be able to:

a) Describe greenhouse gases
b) Define "Ozone layer"
c) Describe standards on air quality.

16.10 Introduction.

The human activities and other natural disasters occurred cause the great accumulation of gases in the atmosphere. This has indeed damage the environment to the extent. The changes in environment has disturbed even rain system. Therefore, Scientists studies, focus on finding out *what causes change in environment*, and *how this can be regulated.* In this unit, you are going to focus on greenhouse gases and Ozone layer.

16.11 Greenhouse gases.

These are airs which accumulate in the atmosphere. Their abundance in the atmosphere leads to rise of global temperature. Example of these gases are;

- Carbondioxide
- Sulphar dioxide
- Nitrogen dioxide
- Nitrogen monoxide
- Methane

Sources of greenhouse gases.

Sources of greenhouse gases depends on the type of the gases.The following are greenhouse gases with their sources included.

Table 16.11 Greenhouse gases and their sources.

Greenhouse gases.	**Sources.**
Carbon dioxide.	• Combustion • Respiration • Fermentation of sugars • Decompositon of death plants and animals
Sulphar dioxide	• Burning of fossils fuels. • From roasted metals.
Nitrogen monoxide and dioxide.	• Lightening • Combustion of fossil fuel. • Internal combustion of engine
Methane.	• Wetlands • Oceans • Activities of termites.

Effects of greenhouse gases.

- Cause Acid rain
- Have adverse effects on human health
- Global warming

Global warming.

This is the continuous rise in average temperature on earth due to the effects of greenhouse gases on the earth.

Effects of Global warming.

Due to accumulation of greenhouse gases on the atmosphere, several effects are noticed. The well-known effect is **global warming.** This has the following effects:

- Melting of mountain glaciers and polar ice.
- Sub-merging of Islands
- Expansion of deserts
- Formation of hurricanes and typhoons.

However, **the effects of greenhouse gases** can be reduced by various ways to improve climatic conditions**.** ***How possible can it be?*** The following are ways of reducing the effects of this global warming.

- Use renewable source of energy.
- Afforestation (planting of trees on land)
- Catalytic Conveters should be sitted to Automobile engines
- Reafforestation. (plant trees to replace the cut down ones)

Self-Assessmment 16.1.

1. Describe greenhouse gases.
2. Name three examples of greenhouse gases and their sources.
3. State three effects of greenhouse gases.
4. Describe three effects of global warming.
5. Explain how global warming effects can be reduced.

16.12 Ozone layer.

The region of the earth's stratosphere that absorbs most of the sun's ultraviolent radiation. This mainly found in the lower parts of the stratosphere about 20km from the earth. However, the levels of Ozone layer vary seasonally and geographically. The ozone layer contains less than 10 parts per millions **(ppm)** of ozone, while the average ozone concentration in Earth's atmosphere as a whole is about 0.3 parts per million **(ppm)**. The important of this ozone layer is that, it absobs 93% to 99% of the sun's high frequency ultraviolent light. This helps to avoid the following bad;

- skin cancer
- damage to crops
- harm to aquatic life.

How does it get destroyed?

Aerosols, *coolants* and *foams* release chloroflouroCarbons(CFCs).In the atmosphere, chloroflouroCarbons release chlorine atom(Cl) which splits ozone layer (O_3) into oxygen atom(O) and oxygen molecules (O_2) or oxygen gas.The oxygen atom then combine with ozone molecule.This form another oxygen gas or oxygen molecules.

For example;

Ozone $\xrightarrow{\text{Chlorine}}$ oxygen molecule + Oxygen atom

$O_{3(g)} \xrightarrow{Cl} O_{2(g)} + O$

Therefore;

Ozone + oxygen atom $\longrightarrow$ oxygen molecule

$O_3(g) + O \longrightarrow 2O_2(g)$

The reaction of ozone layer (O_3) with oxygen atom (O) result into ***depletion of ozone layer***

Effects of ozone layer depletion.

This allows more harmful ultraviolent radiation to penetrate to the earth. Therefore,

it causes;

- skin cancer,
- damage to crops
- eye defects.

Ways of reducing ozone layer.

- Use of CFCs-free aensol sprays.

- Replacing stock refrigerants with less harmful hydro & Clorocarbons($HCRC_s$)Recycling empty household aerosol consumers.
- Adopt the use of non-refrigerants e.g Ammonia(Md_3), Propane(C_3H_8) and bulane($C4H_{10}$)

Self-Assessment 16.2

1. Define ozone layer.
2. Explain how ozone layer is depleted.
3. State three effects of ozone layer depletion.
4. Describe three methods how ozone layer can be reduced.

16.13 Air quality standards.

This is based on World Health Organisation for air pollutants. Therefore, the Malawi government through Malawi Bureau of Standards(MBS) published in the year 2005 guidelines on maximum allowable emisslons or both gaseons and particulate pollutant in a document called **Malawi Standards MS 737:ZW5.**

TABLE: Ambient Air Quality Standard Limits in Malawi

Pollutants	**Maximum concentration in ambient Air**	**Average period**
Suspended particulate matter	25 ug/m^3	1 day
Carbon dioxide	9ppm 35ppm	8 hours 1 hour
surphur dioxide	0.20ppm 0.08ppm 0.02ppm	1 hour 1 day 1 year
Nitrogen dioxide	0.12ppm 0.3ppm	1 hour 1 year
Ozone Lead , ug/m^3	0.12ppm 0.5ug/m^3	1 hour 1 year
Photochemical Oxidants	0.10ppm	1 hour

	0.08ppm	4 hours

KEY

- Ppm *(Parts Per Million)*
- *ug*/m^3 *(concentration per Unit Volume)*

16.14 Summary.

In this unit, you have learnt that greenhouse gases are carbon dioxide, sulphur dioxide, Nitrogen dioxide. These accumulate in the atmosphere. Therefore, their abundant in the atmosphere lead to rise in global temperature and depletion of ozone layer. Ozone layer absorps most of the suns ultraviolent light which protects plants and animals from harmful condition like skin cancer and damage of crops. As such, its depletion leads to the conditions mentioned. Air quality standards table has also been included in this book.

16.15 Revision Exercise.

1(a) State three examples of greenhouse gases.

(b)Mention two source of each of the greenhouse gase mentioned in 1a.

(c)List three effects of green house gases on the environment.

2(a)Define “global warming”.

(b)State three ways on how you can minimize the problems of global warming.

c(i)Define “ozone layer”

(ii)Explain how ozone layer is depleted?

(iii)State three effects of ozone layer depletion.

4 State four ways of reducing ozone layer destructions.

UNIT 17

Waste Management.

Learning Objectives.

By the end of this unit, you should be able to;

a) Classify different types of waste.

b) State major sources of waste.

c) Describe environment friendly ways of treating and disposing wastes

d) State ways of minimizing wastes.

e) Describe the social and economic importance of recycling metals and plastics

17.10 Introduction.

The term" **Waste**" means unwanted materials. Therefore, ***waste management*** involves *controlling, and monitoring of waste materials*. This is to ensure good environment around us.

In this unit, you will concentrate on clustes, sources and methods of disposing and treating wastes.

Table 17.10 Classification of waste and their examples.

Physical state.	• Liquid waste • Solid waste • Gas waste • Sludge wastes.
Level of dangers.	• Hazardous waste (toxic, flammable). • Non-hazardours wastes.
Biodegradability.	• Plants remain. • Animal remain. • Polythene. • Glass. • Plastics.

17.11 Sources of wastes.

- Domestic sources e.g *food leftover*.
- Industrial waste e.g *used oil and toxic gases,*
- Waste from furms e.g *Pesticides, herbicides.*
- Laboratories waste e.g *expired chemicals, plastics*.
- Agricultural sources e.g *Fertiliser and herbicides*.
- Medical sources e.g *expired drugs, linen*.

Effects of wastes on the environment.

- air pollution.
- soil pollution.
- water pollution
- Carry intectious diseases eg toxic gases
- They cause spread of cholera disease.
- Toxic gases increase respiratory infections.

Self-Assessment 17.1

1. Define the term waste.
2. State three classification of waste and their examples.
3. Name three sources of waste.
4. Describe three effects of waste on the environment

17.12 Methods of sound waste disposal.

Burning

Some wastes like papers, plastics, carton boxes, bread wrappers must be burt into ashes.

Burying

Broken bottles, pots, glass and cups must be buried deep down the soil

Animal feeds

Bones are dried and grounded into the powder of which are used as a dog's feeds.

Composite manure

Some wastes from pit can be used to improve soil fertility. When they have completely decomposed.

Recycling

Old papers, magazines, cartons, and packaging materials, should be recycled to the industry.

Re-use

Water for washing plate can be re-used for watering flowers and other vegetables.

Incineration

Medical wastes should be burnt by incinerators.

17.13 Ways of minimizing wastes.

This involves:

- Recover
- Recycle
- Re-use

Recycling.

This is based on reclamation of the wastes material and processing it. The recycled materials are usually obtained from which they get depleted over time like **soil.**For example, scrap metals from used copper and tin products are usually recycled so as to preserve their deposits for future generations.

Social and economic Important of recycling wastes.

- Helps in conserving non-renewable resources.
- Saves energy.
- Reduce air and water pollution

- Sources of employment e.g in town assemblies
- Waste material becom useful materials after being disposed.

Self-Assessment 17.2

1. Describe two methods of disposing waste in the industry.
2. Define recycling.
3. Name three ways of minimizing the wastes.
4. Mention three social and economic Important of recycling wastes.

17.14 Summary.

Wastes are unwanted materials which affects the environment. These are classified into physical state, level of dangers and biodegradability. However, wastes can be managed by recover, recycle and reuse. Recycling wastes reduces air and water pollution.

17.15 Revision Exercise

1(a) Define “wastes”

(b) State two classification of wastes and one example of each class.

(c)Mention two sources of the following wastes

(i)Domestic wastes.

(ii)Indusrial waste

(iii) Laboratery wastes

(iv) Medical waste

2(a) Mention three ways of disposing wastes

(b)State three ways of minimizing wastes.

(c)Explain why recycling of wastes is very important.

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Dedication of the document.

This book is dedicated to my mother, Rachael Kamanga Chuzu, who scourged in vain and left the world on 27th Augast, 2002 together with my sister (Grace Chuzu). May their soul rest in peace.

I should also dedicate this book to form 4 students at Top Hill Private Academy (2017-2018). Their excellent questions have assisted me a lot to come up with techniques that will surely help students to understand the concepts.

About the Author

Chuzu Elisha Chafika hails from Kasungu district traditional authority Kawamba in Chisamba village. He is a former student of Robert Blake Secondary School and Mwimba Primary School. Currently he is studying Bachelors degree in Veterinary Medicine at Lilongwe Universities Agriculture and Natural Resources (LUANAR: Bunda campus)

Chuzu Elisa was teaching Chemistry at Top Hill Private academy in 2017-2018 and at kakhome CDSS in 2018. He wrote this book in response to the excellent questions and positive comments from the students posed during the lesson.

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Many of the explanations, examples ……. And even the humor ….. Will have a Chemistry focus, simply because that is the perspective with which he is most familiar. But he very aware that this is not the only perspective, and look forward to learn from you, just as much as you learn from him. The basis of science is understanding and application. Hence, understanding is worth than memorizing. Once you understand share the knowledge by your friends to keep your head Excellency. Chemistry is fun.

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