

MALAWI SCHOOL CERTIFICATE OF EDUCATION

CHEMISTRY BLEND

A Candidate's Companion towards a good grade in MSCE Physical Science

SYLLABUS BASED AND
MSCE EXAM ORIENTED

FR. BENEDICT ZELE

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Feedback and order: 0999 230 936
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Chapter One:

GENERAL INTRODUCTION BASIC TERMS AND FACTS

- ✓ MATTER: Anything that has mass and occupies space. The three states of matter are solids, liquids and gases.
- ✓ An ATOM is the simplest particle of matter.
- ✓ A MOLECULE is the smallest particle of matter that can exist separately.

WHAT DOES THE KINETIC THEORY OF MATTER STATE?

Facts about the kinetic theory of matter must be learnt like the multiplication table

- All matter is made up of tiny particles. Particles refer to atoms or molecules or ions.
 - The particles are always in constant rapid/haphazard motion. In solids they vibrate about their fixed positions. In liquids they slide past each other. In gases they move randomly.
 - There are forces of attraction called Intermolecular Forces (IMF) also known as Van Der Waals forces. These forces act outside and between molecules.
 - Temperature increases the kinetic energy of the particles and weakens the IMF.
 - Lighter particles diffuse faster than heavier particles.
- ❖ *You are building your house on sand if you continue studying physical science before learning verbatim the above facts and terms!!!!*

INTERMOLECULAR FORCES

They do a “cement bonding” job. These forces act between and outside the molecules of a substance. Solids have the strongest IMF. Liquids have IMF of intermediate strength. Gases have the weakest IMF. As heat energy is supplied to the substance the IMF are steadily weakened.

Chapter two

THE STRUCTURE OF AN ATOM

o. INTRODUCTION

As milk is to a newly born baby, so is an understanding of an atom to a Chemistry student.

1. THE STRUCTURE OF AN ATOM

An atom has a centre called NUCLEUS. The nucleus contains two types of particles namely PROTONS and NEUTRONS. Protons and neutrons are collectively called NUCLEONS. Around the nucleus electrons rotate at very high speeds. The paths electrons follow are called shells, orbitals or energy levels. They are called energy levels because they are associated with a certain form of energy called kinetic energy. Protons carry a positive charge of +1 and have a mass in atomic mass unit (amu) of 1. Neutrons have no charge but have a mass of 1 amu. Electrons carry a negative charge. The mass of electrons is so negligible (insignificant) that it is regarded as zero. Consider adding a bucket of water to Lake Malawi. What effect will it have on the level of the water in the lake?

COMPLETE THE TABLE BELOW

PARTICLE NAME	MASS IN AMU	CHARGE
PROTON		
ELECTRON		
NEUTRON		

Table 2.1 shows the mass and charge of particles in an atom

An atom is electrically neutral. MEANING: it has an equal number of protons and electrons. The positive charges and negative charges balance and cancel each other out. The electrons are at a distance from the nucleus. The radius of the atom is more than 10000 times larger than the nucleus. The atom is mostly empty space.

EXERCISE 1a

1. *Name the part of the atom which is heaviest?*
✓ **Nucleus**
2. *GIVE A REASON:*
✓ It contains particles (protons and neutrons) which have mass.
3. *What is the overall charge of the nucleus?*
✓ **Positive**
4. *WHY?*
✓ Because of the presence of the positively charged protons.
5. *Why does the nucleus attract the energy levels?*
✓ Because the nucleus and the energy levels are oppositely charged.

2. ATOMIC NUMBER

Atomic number of an element is given by the number of protons in the nucleus of an atom.

SYMBOL for Atomic Number: Z

- Atomic number gives an atom its identity.
- It is very unique for a particular element.
- It determines the electron configuration which decides chemical properties of an element. (Electron configuration is discussed under the periodic table). The element carbon is carbon because it has 6 protons. To speak of carbon with anything other than 6 protons is like speaking of a triangle whose sides are not 3 – a flying elephant, mafupa ang’ombe mu nsomba.

3. ATOMIC MASS/ ATOMIC MASS NUMBER/ MASS NUMBER/ NUCLEON NUMBER

It is the total number of protons and neutrons in the nucleus of an atom.

SYMBOL: A

Number of Neutrons = Mass Number (A) – Atomic number (Z)

4. NUCLEAR NOTATION

It is a conventional way of representing an element.

Fig 2.1. The nuclear notation of an element.

Another frequently used notation is to use the name of the element followed by the atomic mass number e.g. U – 235, U – 236.

The nuclear notation for proton, neutron and electron are as follows: $p_1^1 \quad n_0^1$
 e_1^0

5. ISOTOPES AND ISOBARS

Definition of Isotopes: Isotopes are atoms of the same element that have the same atomic number but different mass numbers because of differences in the number of neutrons in their nuclei.

Why do isotopes have similar chemical properties?

- ❖ They have the same electron configuration.

Examples of isotopes

- ✓ Isotopes of carbon are C-12 (with 6 protons, 6 electrons and 6 neutrons), C-13 (with 6 protons, 6 electrons and 7neutrons) and C-14 (with 6 protons 6 electrons and 8 neutrons).

- ✓ Isotopes of hydrogen are H-1 called a proton (with 1 proton, 1 electron and 0 neutrons), H-2 called deuterium (with 1 proton, 1 electron and 1 neutron) and H-3 called tritium (with 1 proton, 1 electron and 2 neutrons)

WORKING OUT THE AVERAGE MASS OF AN ATOM

The average mass of an atom is the average mass of the isotopes of that atom.

WORKING:

(Mass of isotope 1 X its percentage) + (mass of isotope 2 X its percentage) divided by 100

Copper has two main isotopes Cu-63 and Cu-65. In nature they occur in the proportions 69% and 31% respectively. Work out the average mass of the isotopes of copper.

$$\frac{63 \times 69 + 65 \times 31}{100}$$

$$\frac{4347 + 2015}{100}$$

$$\frac{6362}{100} = \underline{\underline{63.62 \text{ amu. Ans.}}}$$

EXERCISE 2b

1. The two main isotopes of carbon are C-12 and C-13. These are present in nature in the proportions 98.89% and 1.11% respectively. Work out the average mass of the isotopes of carbon.
2. iii. The two isotopes of Bromine are Br-79 and Br-81. In nature they occur in equal proportions. Work out the average mass of the isotopes of bromine
3. Chlorine has two isotopes, Cl -35 and Cl-37. For every 4 chlorine atoms three are Cl-35 and one is Cl -37. Work out the average mass of the isotopes of chlorine.

DEFINITION OF ISOBARS:

Isobars are nuclei of the same mass number but different atomic numbers.

Why do isobars have different chemical properties?

- ❖ They have different electron configurations.

Examples of isobars

- ✓ Nitrogen-14 and carbon-14
- ✓ Xenon-131 and Iodine-131

6. RELATIVE ATOMIC MASS (RAM)

It is the average mass of the isotopes of an element compared to $\frac{1}{12}$ of the mass of Carbon-12 atom.

$$\text{FORMULA: RAM} = \frac{\text{Average mass of isotopes of an element (amu)}}{\frac{1}{12} \text{of the mass of 1 atom of C-12 (amu)}}$$

- Average mass of isotopes and RAM are exactly the same in value. However RAM has no SI units. It is a ratio. The amu above and below cancel each other out.
- RAM can have values which are not whole numbers e.g. RAM of chlorine is 35.5. It is found from the average of the different atoms making up the element in nature.
- To make things simple the RAM value of an atom of an element is given to the nearest whole number.

7. RELATIVE FORMULA MASS (RFM) / RELATIVE MOLECULAR MASS (RMM)

It is the mass of one formula unit compared to $\frac{1}{12}$ of the mass of Carbon-12 atom.

$$\text{FORMULA: RFM} = \frac{\text{Mass of one formula unit (amu)}}{\frac{1}{12} \text{of the mass of 1 atom of C-12 (amu)}}$$

EXERCISE 2c

- Work out the molecular mass of ethanoic acid $\text{C}_2\text{H}_5\text{COOH}$
- Work out the RFM of Calcium sulphate CaSO_4 .

8. COMPOSITION CALCULATIONS

$$\text{Percentage mass of an element} = \frac{\text{Mass of an element in compound} \times 100}{\text{RFM of compound}}$$

EXERCISE 2d

- Work out the percentage oxygen in $\text{Cu}(\text{NO}_3)_2$.
- Work out the percentage of hydrogen in ethanol.

9. REACTING MASSES

Consider Magnesium reacting with oxygen to produce magnesium oxide

I. Work out a balanced chemical equation = $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$

OSAIWALA! Always use balanced chemical equations in calculations dealing with reacting masses.

II. Calculate the number of Mg used up, O₂ required and MgO produced.

$$2\text{Mg} = 2 \times 24 = 48\text{g}, \text{O}_2 = 2 \times 16 = 32\text{g}, 2\text{MgO} = (2 \times 24) + (2 \times 16) = 48 + 32 = 80\text{g}$$

III. Suppose there are 12g of Mg used up work out the mass of MgO produced

$$\text{USE PROPORTION! } 48/12 = 80/x, 48x = 12 \times 80, x = \frac{12 \times 80}{48} = 20\text{g}$$

Chapter three

THE PERIODIC TABLE OF ELEMENTS

O. INTRODUCTION

- a. **The Periodic Table:** is a chart in which elements are arranged according to their atomic numbers.
- b. **Periodicity:** is the arrangement of elements in a table according to their properties.
- c. **Valence Electrons:** is the number of electrons in the outermost shell.

SIGNIFICANCE OF VALENCE ELECTRONS:

- ✓ Determining the number of chemical bonds an atom can make.
- ✓ Determining the group of an element.

- d. **Electron Configuration:** is the number and arrangement of electrons in an atom.

SIGNIFICANCE OF ELECTRON CONFIGURATION

- ✓ Determining the chemical properties of an element.

HOW DO ELECTRONS FILL UP ENERGY LEVELS?

Two electrons fill the first energy level. Eight electrons fill the second energy level. Eighteen electrons fill the third energy level. Thirty two electrons fill the fourth energy level.

KOMATU Izi osazigwirtsitsa (Overlook this as of now) because for MSCE chemistry understanding we say two electrons fill the first energy level and eight electrons fill rest of the energy levels.

- e. **GROUPS or FAMILIES** are vertical columns of elements in the periodic table.

What decides the group of an element?

- Number of valence electron

- f. **SERIES or PERIODS** are horizontal rows of elements in the periodic table.

What decides the period of an element?

- The number of shells.

1. ATOMIC RADIUS

Atomic Radius is the distance from the centre of the nucleus of an atom to the outermost shell.

FACTORS AFFECTING ATOMIC RADIUS

- a. Number of shells. An increase in the number of shells means that outermost shell electrons are placed farther away from the nucleus to increase atomic radius.
- b. Nuclear Charge/ Atomic number/Proton number
An increase in the number of protons increases the nuclear charge. This has a pulling effect on the outermost shell electrons and reduces the atomic radius.

WHAT HAPPENS TO ATOMIC RADIUS AS WE GO ACROSS A SERIES?

It decreases.

Explanation: Going across a series the number of protons increases resulting into an increase in nuclear charge. This has a pulling effect on the outermost shell electrons and reduces the atomic radius.

EXERCISE 3a

Explain why atomic radius increases as we go down the group?

2. ELECTRONEGATIVITY

Electronegativity is the attracting ability of an atom to its electrons

CAUSE OF THIS ATTRACTION: The positively charged nucleus attracts the negatively charged energy levels because opposite charges attract.

FACTORS AFFECTING ELECTRONEGATIVITY

a. Atomic radius: An increase in the atomic radius means that outermost shell electrons are placed farther away from the nucleus which holds onto the electrons. The effect is that electrons are pulled less strongly to the nucleus. This decreases electronegativity.

b. Nuclear Charge/ Atomic number/Proton number

An increase in the number of protons increases the nuclear charge. This has a pulling effect on the outermost shell electrons. This strengthens the attracting ability of an atom to its electrons and so increases electronegativity.

EXERCISE 3b

Explain what happens to electronegativity of elements as we go across a series on the periodic table

3. GROUP ONE ELEMENTS

MEMBERS: Lithium, Sodium, Potassium, Rubidium, Caesium and Francium.

COMMON NAME: Alkali Metals

WHY THEY ARE XALLED ALKALI METALS: They react with water to produce alkaline solutions. The other product is hydrogen gas.

a. PHYSICAL PROPERTIES OF ALKALI METALS

1. They are soft and so are easily cut by a knife.
2. They are good conductors of heat and electricity.
3. They have low density
4. They are ductile – easily drawn into wires.
5. They are malleable – can be hammered into any shape.
6. They are shiny when freshly cut.
7. They are solids at room temperature.

What determines the physical properties of alkali metals?

- It is the metallic bonding.

Why do alkali metals have low density? Why are they soft?

- The metallic bonding is weak.

Why are alkali metals ductile, malleable, good conductors of heat and electricity?

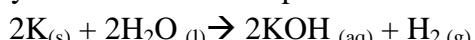
- The metallic bonding has free electrons.

Why do alkali metals become softer, more ductile, less dense on going down group one?

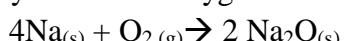
- The metallic bonding weakens more.

b. CHEMICAL PROPERTIES OF ALKALI METALS

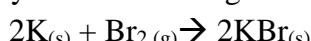
- i. They react with water to produce alkaline solution and hydrogen gas.



- ii. They react with oxygen to form a white solid oxide.



- iii. They react with halogens to produce white solid halides.



Why are alkali metals the most reactive metals on the periodic table?

- ❖ They have one valence electron which is loosely held by the nucleus.
During chemical reactions it is easily lost.

What happens to reactivity of alkali metals as we go down group one

- ❖ Reactivity increases.

EXPLANATION: Going down group one the number of shells increase. This increases atomic radius. The Valence electron experiences less forces of attraction by the nucleus. Elements easily lose the electron.

EXERCISE 3c

A newly discovered element X is suspected to be a group one metal on the periodic table. Describe an experiment you will carry out to verify this supposition. (*Maneb, 2004*)

4. GROUP TWO ELEMENTS (ALKALINE EARTH METALS)

MEMBERS: Magnesium, Calcium

They exist in nature in the earth's crust as carbonate, silicate and sulphates.

They react with water vapour to produce bases and hydrogen gas.

**Alkalies are bases which are soluble in water.

5. GROUP SEVEN ELEMENTS

MEMBERS: Fluorine, Chlorine, Bromine and Iodine.

COMMON NAME: Halogens.

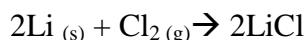
a. PHYSICAL PROPERTIES OF GROUP SEVEN ELEMENTS (Halogens)

1. They produce a pungent smell.
2. They have varying colours e.g. fluorine is very pale yellow gas, chlorine is a yellowish green gas, bromine is a reddish brown liquid, iodine is a blue black shiny solid.
3. They do not conduct electricity because they are molecular substances.
4. They are molecular substances being covalently bonded.
5. They are diatomic.
6. At room temperature they exist in all the three states. Fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid.
7. They are poisonous.
8. They have relatively low boiling and melting points.

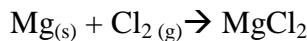
b. CHEMICAL PROPERTIES OF HALOGENS

- i. They react with metals to form halides.

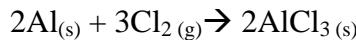
- a. *Reaction with alkali metals*



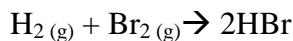
- b. *Reaction with group two metals*



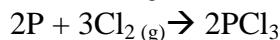
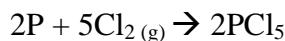
- c. *Reaction with aluminium*



- ii. They react with hydrogen to form hydrogen halides. Hydrogen halides dissolve in water to form acidic solutions.



- iii. They react with phosphorous to produce covalent phosphorous pentahalides and phosphorous trihalides.



Why are halogens the most reactive members of the nonmetal group?

- ❖ Halogens have seven valence electrons. When they react they easily gain one electron which they require to attain an electron configuration of a noble gas.

What happens to reactivity as we go down group seven?

- ❖ It decreases

EXPLANATION: Going down group seven the number of shells increase. Atomic radius increases and electronegativity decreases. Halogens gain electrons in chemical reactions and the incoming electron is less attracted to the nucleus resulting into decrease in reactivity.

What happens to boiling point of elements as we go down group seven?

It increases.

EXPLANATION: Going down group seven molecular sizes increase. Making Intermolecular forces to increase too. Greater heat is therefore required to separate these strong intermolecular forces.

c. USES OF HALOGENS

i. Fluorine

- ✓ Used in toothpaste in a form of fluoride to reduce tooth decay.

ii. Chlorine

- a. Used as a disinfectant.
- b. Used as a bleaching agent.
- c. Making pesticides.
- d. Making herbicides (weed killer chemicals).
- e. Making typex correction fluid.
- f. Making solvents for dry cleaning.
- g. Making plastics e.g. polyvinyl chloride.

iii. Iodine

- Making drugs.
- Making printing ink.
- Making animal feed.
- Making dyes.
- Making iodised salt.

iv. Bromine

- ✓ Used as a photographic film in a compound of silver called silver bromide.
- ✓ Used in making fuel additives.
- ✓ Making pesticides.
- ✓ Making of dyes.

d. HOW TO PREPARE HALOGENS

Halogens occur naturally as halides. To prepare halogens concentrated sulphuric acid acts on alkali metal halides in the presence of manganese dioxide.



6. GROUP EIGHT ELEMENTS

MEMBERS: Helium, Neon, Argon, Krypton And Xenon.

COMMON NAME: Noble/Inert/Stable Gases

a. PHYSICAL PROPERTIES

1. They are monoatomic.
2. They have low melting and boiling points.
3. They are gases at room temperature. They are present in air in very small amounts.
4. They are colourless.
5. They are odourless.
6. They have a low density.

b. CHEMICAL PROPERTIES

- ❖ The Noble Gases have no chemical properties because their outermost valence shells are filled with electrons.

c. USES OF THE NOBLE GASES

i. Argon

- ✓ To fill ordinary and long life light bulbs to prevent the tungsten filament from reacting with oxygen in the air and forming an oxide.
- ✓ To provide an inert atmosphere in arc welding and in the production of titanium metal.
- ✓ In Geiger-Muller tubes which are used for the detection of radioactivity.

ii. Neon

- In advertising signs, because it glows red when electricity is passed through it.
- In the helium-neon gas laser used in eye surgery

iii. Helium

1. To provide an inert atmosphere for welding
2. As a coolant in nuclear reactor
3. With 20% oxygen as a breathing gas used by deep sea divers.
4. To inflate the tyre of large aircrafts
5. In the helium-neon laser
6. To fill ships and weather balloons

EXERCISE 3d

- a)* Why are atoms said to be electrically neutral?
- b)* Why is Calcium placed in group two on the periodic table?
- c)* Why is Helium included in group eight?
- d)* Why are Germanium and Silicon regarded as metalloids?

Chapter four

TYPES OF CHEMICAL BONDING

(Covalent, ionic and metallic bonding)

O. INTRODUCTION

A chemical reaction is the rearrangement of atoms to form new substances.

WHY CHEMICAL REACTIONS TAKE PLACE?

It is in order for an atom to attain the electron configuration of a noble gas (Octet rule or duplet rule) and so

- i. Find stability
- ii. Reach a lower energy level
- iii. Experience greater forces of attraction.

Imagine sleeping all alone in the hostel. You feel unsafe and require your friend's company.

HOW DO ATOMS FIND STABILITY?

1. By gaining electrons.
2. By losing electrons.
3. By sharing electrons.
4. In Radioactivity atoms find stability by random spontaneous disintegration of unstable heavy nuclei (Discussed in **Nuclear Physics**).

VALENCY: is the number of chemical bonds an atom can make. In Ionic compounds it is given by the number of electrons lost or gained. In molecular substances it is given by the number of electrons offered for sharing.

1. COVALENT BONDING

This is the sharing of electrons between non-metals in order to attain the electron configuration of a noble gas.

A. WHAT FORMS THE COVALENT BOND?

The attraction between the shared electrons with the nuclei of the atoms involved.

B. EXAMPLES TO ILLUSTRATE COVALENT BONDING

I. Consider carbon (C) and hydrogen (H) reacting to form methane

C has 4 electrons in the outermost shell. It requires 4 electrons to attain an electron configuration of a noble gas. Hydrogen has 1 electron in the outermost shell. It requires 1 electron to reach an electron configuration of a noble gas (See, Fig. 4.1a.).

When carbon and hydrogen react one carbon atom offers four electrons for sharing to four hydrogen atoms. Each of the four hydrogen atoms offers one electron for sharing with the carbon atom. The hydrogen atom with one shell and two electrons in the

outermost shell attains noble gas electron configuration. The carbon atom now with eight electrons in the outermost shell also attains an electron configuration of a noble gas. The attraction between the shared electrons and the nuclei of carbon and hydrogen forms the covalent bond

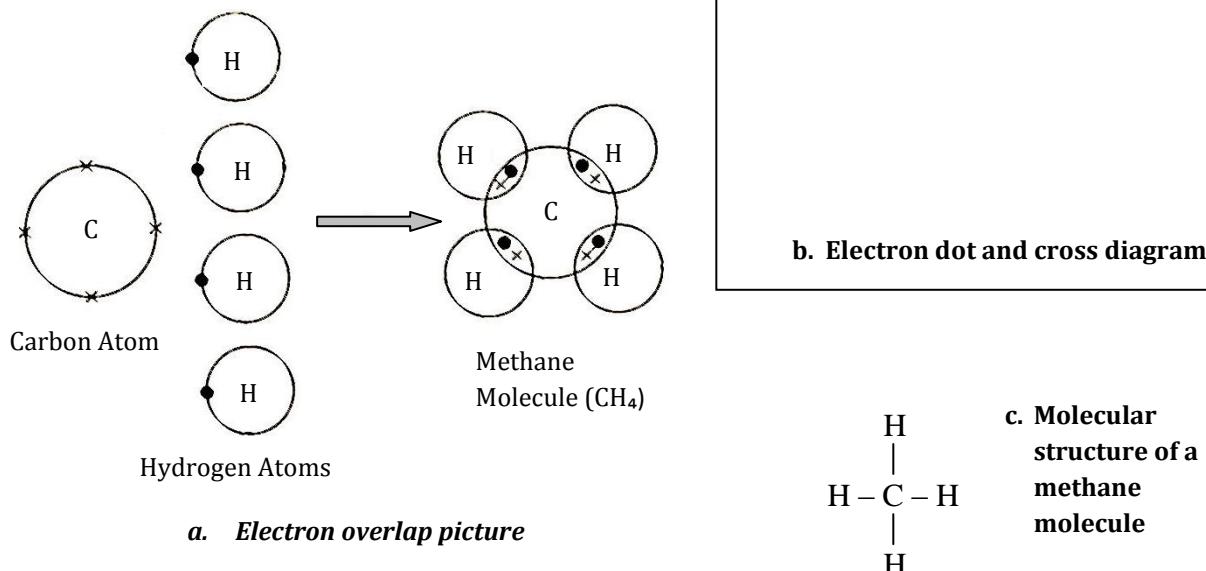


Fig. 4.1 (a) shows the electron overlap picture, (b) the electron dot and cross diagram and (c) molecular structure of a methane molecule and its bonds.

II. Consider two nitrogen atoms reacting to form nitrogen gas

Nitrogen has five electrons in the outermost shell and requires three electrons to attain noble gas electron configuration. When two nitrogen atoms react, each offers three electrons for sharing and both of them attain octet rule. The attraction between the shared pairs of electrons and the nuclei of the two atoms forms the covalent bond.

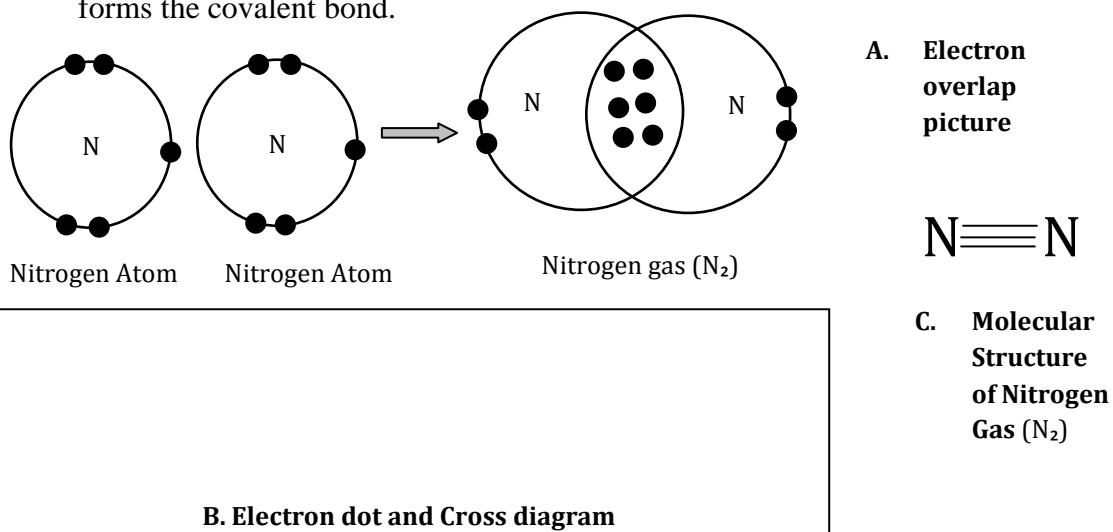


Fig 4.2 (A) the electron overlap picture, (B) electron dot and cross diagram and (C) the molecular structure of Nitrogen gas showing a triple bond.

EXERCISE 4a

1. Discuss how hydrogen and oxygen combine to form water.
2. Discuss the bonding between hydrogen and chlorine to form hydrochloric acid.

C. PROPERTIES OF COVALENT SUBSTANCES (Molecular Compounds)

1. They are made of molecules or atoms and not ions
2. They have weak IMF
3. They are not very soluble in water
4. They have low melting and boiling points
5. They can easily be vapourised since they are volatile
6. They have a particular smell e.g. camphor
7. They are bad conductors of heat and electricity.

D. TYPES OF COVALENT BONDS

- a. **Single Bonds:** When atoms form the bond by sharing a single pair of electrons e.g. chlorine gas Cl-Cl.
- b. **Double Bonds:** When atoms form the bond by sharing two pairs of electrons e.g. CO_2 O=C=O.
- c. **Triple Bonds:** These form when atoms share three pairs of electrons e.g. N_2 N≡N.
- d. **Dative/Coordinate Bonds:** This is formed in a covalent bond where the shared pair of electrons is provided by one of the bonded atoms. The atom which provides both electrons for the covalent bond is called the DONOR. The atom which accepts both electrons is called the ACCEPTOR. E.g. Ammonium ion - nitrogen atom in ammonia offers both electrons for sharing with hydrogen ion to form ammonia.

Fig. 4.3 The dative bond

E. TYPES OF MOLECULAR SUBSTANCES

I. Polar Covalent Compounds

These are formed between two atoms which have combined where one is more electronegative than the other resulting into formation of partial charges. Partial charges form because the electron density favours the more electronegative element and electrons spend more time in its vicinity. An example is hydrogen chloride gas (HCl). Chlorine is more electronegative than hydrogen. Electrons spend more time in the vicinity of chlorine. The chlorine end of $\text{HCl}_{(g)}$ is partially negative (delta negative δ^-). The hydrogen end is delta positive (delta positive δ^+).

II. Non Polar Covalent Compounds

These are formed between two atoms which have combined and have the same electronegativity. The electron density is shared equally. E.g. O_2

III. Simple Covalent Compounds contain few atoms and have weak IMF

IV. Giant Molecular Substances contain hundreds of thousands of atoms. They have strong IMF e.g. Polymers like plastics, diamond and graphite.

2. IONIC BONDING/ELECTROVALENT BONDING

It is the electrical attraction between oppositely charged atoms which are produced when electrons are transferred from one atom to another during a chemical reaction.

A. WHAT FORMS THE IONIC BOND?

- ❖ The attraction between oppositely charged ions.

B. EXAMPLES TO ILLUSTRATE IONIC BONDING

I. Sodium reacting with chlorine to form sodium chloride

Sodium has one valence electron. Chlorine has seven valence electrons. In a chemical reaction sodium donates its electron to chlorine. There is electron transfer in which sodium loses an electron and chlorine gains an electron.

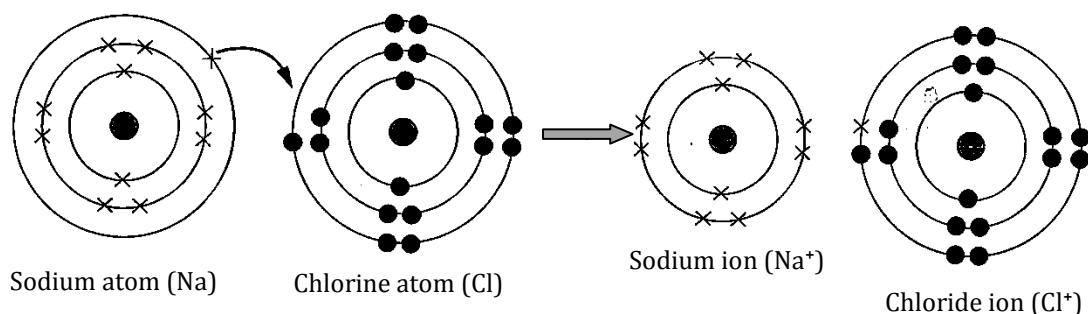


Fig 4.4 Reaction between Sodium and Chlorine to form Sodium Chloride

The sodium atom becomes a sodium cation with an electron configuration of Neon. The chlorine atom becomes a chloride ion with an electron configuration of Argon.

II. Consider the reaction between Lithium and oxygen to form lithium oxide.

Lithium with one valence electron has a valency of one. Oxygen with six valence electrons has a valency of two. In a chemical reaction two lithium atoms donate an electron each to one oxygen atom.

Here is an electron dot and cross diagram to explain the reaction. Take NOTE! We use valence electrons ONLY!



Fig. 4.5 electron dot and cross diagram to explain the reaction.

NOTES

- Ionic bonding happens between metals and non metals
- Metals lose electrons and become cations.
- A CATION is a positively charged ion which is attracted to the cathode. Non metals gain electrons and become anions.
- Ionic Lattice is a giant structure of many ions arranged and packed in a regular manner.
- LATTICE is the arrangement of ions in a well packed manner and in a regular pattern.

EXERCISE 4b

- i. Work out the formula of Aluminium oxide
- ii. Explain the reaction between calcium and chlorine to form calcium chloride. What is the charge of chlorine in the compound calcium chloride?
- iii. Why are negatively charged ions called anions?
- iv. What is an anode?

C. PROPERTIES OF IONIC COMPOUNDS

- I. They are made up of ions not atoms.
- II. The positive and negative ions are held together by strong electrostatic forces of attraction
- III. They have high melting and boiling points
- IV. They conduct electricity either when molten or in solution.
- V. They are soluble in water
- VI. They are hard brittle solids at room temperature.

EXERCISE 4b

- i. Why do ionic compounds conduct electricity when molten while molecular substances do not?
- ii. Explain why solid ionic compounds do not conduct electricity?

NOTES

- Valency is the combining power of an ion.
- The size of the charge on an ion is a measure of its valency
- The sum of the charges of the ions in the ionic compound must be equal to zero.
E.g. in magnesium chloride, $Mg^{+2}Cl_2^- = +2 + (-1 \times 2) = +2 - 2 = 0$
- Ionic compounds are neutral. Charges of the cation and anion cancel each other.
- A correct formula of an ionic compound can be verified by adding up the separate charges of the ions to see if they sum up to zero.

3. METALLIC BONDING

It is the electrical attraction between positive fixed ions and negative delocalised electrons in a metal. The metallic bond forms as gaseous metal cools to form a solid.

A. WHAT HAPPENS?

In a gas state the electrons of a metal are localised. Meaning: they are only to be found inside electron shells. Thus they are not free to move. As metal cools to form a solid metal atoms take up fixed positions and the valence electrons are delocalised. Meaning: they are set free from the energy levels. On being delocalised they become more stable, reach a lower energy level and experience stronger forces of attraction. The metallic bond results from electrostatic forces of attraction between the fixed metal ions and the negative delocalised electrons. We say the fixed metal ions are like stationary islands embedded in a sea of electrons.

Fig. 4.6. The metallic bond

B. PHYSICAL PROPERTIES OF METALS

1. They are good conductors of heat and electricity because the metallic bonding has free electrons.
2. They have high melting and boiling points because the metallic bond is very strong.
3. They have high density.
4. They are shiny e.g. silver, gold.
5. They are sonorous. Therefore they are used in making bells.
6. They have high tensile strength because the metallic bond is flexible.
7. They are ductile.
8. They are malleable.

C. SPECIAL METALS

- ✓ Mercury is a liquid at room temperature.
- ✓ Iron is magnetic.
- ✓ Alkali metals are soft and have no sonorous property.

Chapter five

NITROGEN, SULPHUR AND CARBON

0. INTRODUCTION

This chapter discusses important elements nitrogen, sulphur and carbon in chemistry.

1. NITROGEN

- Makes up 78.08% of the atmosphere
- It is obtained from air by fractional distillation of liquid air.

A. GASES THAT MAKE UP CLEAN AIR AND THEIR PERCENTAGES

Nitrogen 78.08%, Oxygen 20.95%, Argon 0.934%, carbon dioxide 0.035%, water vapour -, Neon 0.00181%, Helium 0.000524%, Krypton 0.00014%, Xenon 0.0000087%.

B. PROPERTIES OF NITROGEN

1. Has no smell
2. Has no colour
3. It is covalently bonded
4. It is very unreactive – we inhale and exhale N₂ without any danger.

NB: At very high temperatures nitrogen reacts with oxygen and hydrogen gas. Nitrogen oxides are harmful and cause pollution of the environment.

C. USES OF NITROGEN GAS

- I. Used to make ammonia gas through the Haber process. Ammonia is used in the making of fertilisers.
 - II. To freeze liquids in damaged pipes before opening them
 - III. Used to make nitric acid (HNO₃). Nitric acid is used for making explosives, dyes and nylon (artificial fibres)
 - IV. Used in food preservation many foods are packed in an atmosphere of nitrogen.
 - V. Filling emptied oil tankers for safety reasons.
- VI. In plant growth, nitrogen
- is a major component of proteins
 - increases leaf area and vegetative growth in crops
 - is responsible for the deep green colour, chlorophyll.
 - regulates the availability and utilisation of phosphorous and potassium
 - makes plants and their fruits succulent e.g. melons
 - increases grain yield by increasing grain size in legumes and cereals.

D. NITROGEN CYCLE

This is a repeated process through which nitrogen gets into the soil from the air, plants and animals and out of the soil.

NITRATES: These are nitrogen containing compounds

I. HOW NITROGEN GETS INTO THE SOIL

a. Through the work of nitrogen fixing bacteria

Leguminous plants have bacteria in the nodules on their roots. These bacteria take nitrogen from the atmosphere and convert it into nitrates. Nitrates are soluble compounds and plants absorb and convert them to make proteins

b. Through acid rain

Nitrogen dioxide in the air dissolves in the rain making acidic solution which forms nitrates in the soil.

NB: Lightening is a necessary condition for formation of nitrogen dioxide gas.

c. Organic matter – Decaying of dead plants and animals

Dead plants and animals produce nitrates as they decay. This is done by putrefying bacteria.

d. Artificial fertilisers e.g. potassium nitrate, ammonium nitrate, urea CAN.

II. HOW NITROGEN GETS OUT OF THE SOIL

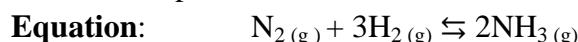
a. By the work of denitrifying bacteria a process called volatilisation.

b. Plants use nitrates from the soil and convert them into proteins.

c. Through immobilisation microorganisms use nitrogen from the soil to build up their bodies.

E. THE INDUSTRIAL PROCESS OF MAKING AMMONIA

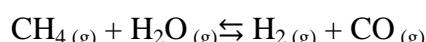
It is an example of a reversible reaction



Raw Materials: Nitrogen gas and hydrogen gas

Nitrogen gas is obtained from the fractional distillation of liquid air while hydrogen gas is obtained from steam reforming process.

STEAM REFORMING PROCESS:



STAGES:

- i. Obtaining raw materials.
- ii. Pressurising N₂ and H₂ to 350 atm at a temperature of 450°C in the presence of iron catalyst.
- iii. Cooling and condensing the ammonia gas.
- iv. Liquefying the ammonia formed.

NOTES

Reversible reactions go in both directions. As soon as the product ammonia forms, it decomposes to form N₂ and H₂.

2. SULPHUR

A. PHYSICAL PROPERTIES OF SULPHUR

- i. It is yellow in colour.
- ii. It is brittle.
- iii. It does not conduct electricity.
- iv. It is insoluble in water.
- v. It has a low melting point-melting point 113°C.

B. SOURCES OF SULPHUR

- i. Volcanic eruptions.
- ii. Fossil fuels e.g. coal, natural gas and petroleum.
- iii. Metal ores such as zinc blend (ZnS).
- iv. Sulphur beds about 200m below the ground. Sulphur from sulphur beds is obtained through the Frasch process.
- v. Eggs.

C. USES OF SULPHUR

- i. Manufacturing of matches
- ii. Making fireworks
- iii. Making vulcanised rubber
- iv. Making sulphur concrete. Sulphur concrete cannot be attacked by acid.
- v. Making paints and dyes
- vi. Making of drugs. Some of the drugs are used for sterilisation and others are used as skin ointments.
- vii. Making of pesticides, fungicides and insecticides
- viii. *Making sulphuric acid.*
- ix. **In plant growth**, sulphur
 - Increases oil content of oil crops
 - Constitutes the amino acids cystine, cysteine and thiamine.
 - Activates some proteolytic enzymes, e.g. papain
 - Is useful in nodule formation of legume roots for nitrogen fixation.
 - Is needed in protein synthesis and improves the biological value of proteins.

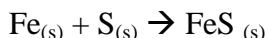
D. USES OF SULPHURIC ACID

- i. Making soapless detergents.
- ii. Used in car batteries.
- iii. Used as a dehydrating agent.
- iv. Making fertiliser.
- v. Manufacturing of plastics.
- vi. Used in the gold refining process.

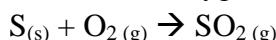
E. CHEMICAL PROPERTIES OF SULPHUR

I. Sulphur reacts with metals to form sulphides.

Iron + sulphur → iron sulphide



II. Sulphur reacts with oxygen to form sulphur dioxide



F. ALLOTROPES OF SULPHUR

Allotropy (POLYMORPHISM)

This is when an element can exist in more than one physical form in the same state. Allotropes are different physical forms of the same element in the same state. The two allotropes of sulphur are rhombic sulphur and monoclinic sulphur.

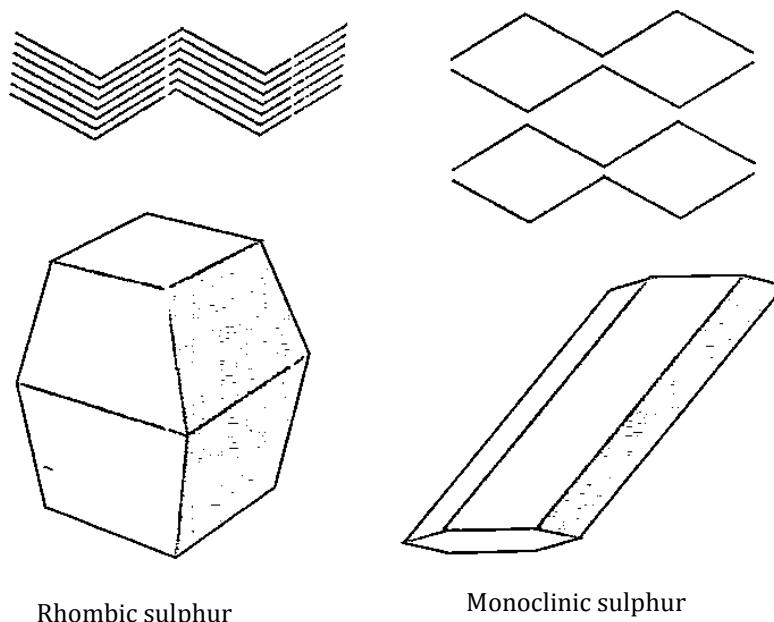


Fig 5.1. The two Allotropes of Sulphur.

FORMATION OF THE ALLOTROPES OF SULPHUR

Explaining why Rhombic Sulphur is more stable than Monoclinic Sulphur

Sulphur is dissolved in methylbenzene and excessive heat is supplied. Insoluble particles are filtered out. The filtrate is allowed to cool. Crystals that are formed above 96°C are needle like and they are called monoclinic or B sulphur.

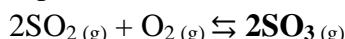
Crystals that are formed below 96°C form rhombic sulphur. In rhombic sulphur the molecules are more closely packed. The IMF are stronger than in monoclinic sulphur. Therefore rhombic sulphur is more stable than monoclinic sulphur. The Melting point of sulphur is 113°C and the boiling point is 444°C.

G. CONTACT PROCESS

This is the industrial process through which sulphuric acid is formed.

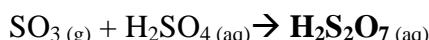
STAGES:

- i. Burning sulphur in oxygen gas, O₂ to produce sulphur dioxide, SO₂.
- ii. Pure sulphur dioxide gas is reacted with pure oxygen gas in the presence of Vanadium oxide catalyst at a temperature of 450°C and a pressure of 1 atmosphere (1atm).

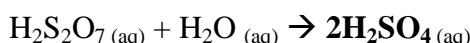


These gases are purified through a dust precipitator and passed through a drier.

- iii. Sulphur trioxide is reacted with concentrated sulphuric acid to produce OLEUM



- iv. The Oleum is reacted to the correct amount of water to produce sulphuric acid.



H. PROPERTIES OF SULPHURIC ACID

1. pH lower than 7.
2. Conducts electricity.
3. Changes colour of indicators.
4. Reacts with bases to produce salts and water. This is neutralisation reaction. Salts of sulphuric acid are called SULPHATES.

I. USES OF SULPHATES

- i. Ammonium sulphate ((NH₄)₂SO₄) is used for making fertiliser.
- ii. Magnesium sulphate (MgSO₄) is used as a laxative
- iii. Manufacturing of plaster of Paris to support fracture. CaSO₄ sets bones.
- iv. Diagnostic medical X-ray studies e.g. Barium sulphate (BaSO₄)
- v. Hydrated sodium sulphate (Na₂SO₄·10H₂O) is used in the manufacturing of paper and glass.

J. SULPHUR DIOXIDE

Type of Bonding: Covalent

Formation: Sulphur reacting with oxygen in the air.

PHYSICAL PROPERTIES OF SULPHUR DIOXIDE

- i. Colourless with choking smell
- ii. Has greater density than air
- iii. It is soluble in water.

CHEMICAL PROPERTIES OF SULPHUR DIOXIDE

- i. Reacts with hydrogen sulphide to produce sulphur and water
- ii. Changes orange colour of potassium dichromate to green colour
- iii. It reacts with water to produce sulphurous acid. In rain water it forms acid rain.

USES OF SULPHUROUS ACID

- i. Bleaching agent.
- ii. Fumigant.
- iii. It is used as a preservative in canned fruit, jam and fruit juices.
- iv. Used for bleaching straw and wool.

DANGERS OF SULPHUR DIOXIDE

- i. Irritates the lungs if inhaled to cause bronchitis and other diseases.
- ii. Dissolves in rain water to form acid rain which destroys forests

ACID RAIN is rain which is abnormally acidic produced as a result of toxic gases reacting with rain water

3. CARBON

Type of Bonding: Covalent

Allotropes: Diamond and graphite

a. GRAPHITE

Type of Bonding: Covalent

It is a giant covalent molecule. Three covalent bonds are made. One electron is left free

PROPERTIES

- i. Dark grey shiny solid
- ii. Soft with slippery feel
- iii. Conducts electricity. WHY? One electron is left free.
- iv. Not so dense.

USES OF GRAPHITE

- i. Making pencils
- ii. Making electrodes
- iii. Making lubricants
- iv. It used as a component in certain sport equipment such as tennis and squash rackets.

b. DIAMOND

Type of Bonding: Covalent

It is a giant covalent molecule. All four electrons are used up in bonding hence no free electrons.

PROPERTIES

- i. Does not conduct electricity.
- ii. Has high boiling and melting points. Its boiling point is 3550°C.
- iii. Transparent and sparkles in light.
- iv. Very hard substance.

USES OF DIAMOND

- i. It is used for making jewellery and ornaments.
- ii. It is used for making drilling bits in derricks
- iii. Making diamond studded saws
- iv. Making glass cutters.
- v. Used for making polishers.

Chapter six

MOLARITY, REACTING MASSES, MOLAR VOLUME, EMPIRICAL FORMULA AND STANDARD SOLUTIONS

a. INTRODUCTION

Weighing Relative Atomic Mass (RAM) of an element in grams contains 6.02×10^{23} atoms. Weighing Relative Formula Mass of a molecular substance contains 6.02×10^{23} molecules. Weighing RFM of an ionic compound contains 6.02×10^{23} ions.

- ✓ 23g of sodium (RAM of Na is 23) contain 6.02×10^{23} atoms.
- ✓ _____ g of Aluminium contain $2 \times 6.02 \times 10^{23}$ atoms
- ✓ 3g of carbon contain _____ atoms.
- ✓ _____ g of calcium contain 6.02×10^{23} atoms
- ✓ 18 g of water contain 6.02×10^{23} molecules
- ✓ _____ of sulphuric acid contain 6.02×10^{23} molecules.
- ✓ 44g of _____ contain 6.02×10^{23} molecules.

6.02×10^{23} is called Avogadro number. The quantity of particles is called the mole.

b. MOLES AND AVOGADRO

A mole of a substance contains 6.02×10^{23} particles of that substance and is found by weighing RAM or RFM of the substance in grams.

Avogadro Number is the number of carbon atoms contained in exactly 12 grams of the C-12 isotope and its value is 6.02×10^{23} .

Formula: No of Moles = $\frac{\text{Mass in grams}}{\text{RAM or RFM}}$

USING THE FORMULA

1. Make mass the subject of the formula?
2. Make RAM or RFM the subject of the formula?

EXERCISE 6a “*Question three has been worked out as an example!*”

- i. Work out the mass of gold (Au) which contains the same number of atoms as 3g of carbon.
- ii. Work out the number of moles in 10g of sulphur dioxide.
- iii. What mass of CuSO_4 contains 0.2 moles?

$$\text{No of moles} = \frac{\text{Mass}}{\text{RFM}}$$

$$\text{Mass} = \text{RFM} \times \text{No of moles}$$

$$\text{Mass} = 160 \times 0.2 = 32\text{g}$$

- iv. Work out the percentage of water in hydrated sodium sulphate.

c. MOLARITY

Definition: It is the concentration of a solution expressed in number of moles per unit volume.

SI Units:

Number of moles is measured in moles. Volume is measured in litres or dm³. 1 mole per 1 litre gives a 1 MOLAR SOLUTION or 1M solution.

WAYS OF EXPRESSING MOLARITY

- (i) Number of moles per unit volume
- (ii) Grams per unit volume
- (iii) As a percentage

FORMULAE

- a. Molarity = $\frac{\text{No of moles}}{\text{volume}}$
- b. No of moles = $\frac{\text{Mass}}{\text{RFM}}$
- c. By substitution Molarity = $\frac{\text{Mass}}{\text{RFM} \times \text{volume}}$
- d. Mass = Molarity x RFM x volume
This is rewritten as $\frac{\text{mass}}{\text{volume}} = \text{RFM} \times \text{Volume}$
But $\frac{\text{mass}}{\text{volume}}$ gives us concentration.
- e. So Concentration = Molarity x RFM
Concentration is mass of solute particles dissolved in a given volume of a solution.
1 litre = 1dm³ = 1000cm³ = 1000ml. 1 litre of water has a mass of 1kg.
- f. C₁V₁ = C₂V₂. Here C is concentration and V is volume.

EXERCISE 6b

- i. What is the mass of solute particles in 250cm³ 1 molar hydrochloric acid solution?
- ii. Given 20g of NaOH pellets in what volume of water can you dissolve it to have 2 molar solution?
- iii. What is the molarity of a solution of sodium chloride labelled 50% (HINT: 50% means 50g dissolved in 100ml of water).
- iv. Work out the molarity of NaCl solution prepared by dissolving 5.8g in 200cm³
- v. What is the molarity of 21% HNO₃?
- vi. 60cm³ of a solution whose concentration is 15g/cm³ was diluted with distilled water by raising its volume to 80cm³. Calculate the concentration of the new solution.
- vii. Calculate the mass of solute particles when 100ml salt solution is removed from a stock of 200g/l
- viii. Explain how you can make a concentration of 20g/l from sugar solution whose concentration is 50g/l.
- ix. A laboratory technician has 1 litre of 2M HCl (aq) solution. Describe how she could prepare 250cm³ from it.
- x. Given 2 litres of 0.2M NaOH solution describe how you can use evaporation method to prepare a 1 NaOH solution from it.

d. REACTING MASSES

Consider Magnesium reacting with oxygen to produce magnesium oxide

- a. Work out a balanced chemical equation = $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$
“OSAIWALA!” Always use balanced chemical equations in calculations dealing with reacting masses.

- b. Calculate the number of Mg used up, O₂ required and MgO produced.

$$2\text{Mg} = 2 \times 24 = 48\text{g}$$

$$\text{O}_2 = 2 \times 16 = 32\text{g}$$

$$2\text{MgO} = (2 \times 24) + (2 \times 16) = 48 + 32 = 80\text{g}$$

- c. Suppose there are 12g of Mg used up work out the mass of MgO produced
USE PROPORTION!

Let MgO be = z

$$\text{Then } 48/12 = 80/z.$$

$$48z = 12 \times 80$$

$$z = \frac{12 \times 80}{48} = 20\text{g}$$

$$z = 20\text{g}$$

therefore the **Mass of MgO = 20g**

Here is another consideration:

Consider Propane burning in air to produce carbon dioxide and water.

Work out a balanced chemical equation



The balanced chemical equation means

- 1 mole of propane gas requires 5 moles of oxygen gas to produce 3 moles of carbon dioxide gas and 4 moles of water molecules
- 44g of propane gas require 160g of oxygen gas to burn completely and produce 132g of carbon dioxide gas and 72g of water molecules.
- 44g of propane produce 72g of water molecules

Now use proportions to work out

- Moles of oxygen gas needed to burn 66g of propane.
- Mass of water molecules produced when 66g of propane is used.

EXERCISE 6c

- (i) Work out the mass of CO_2 formed when 72g of methane burn completely in air?
- (ii) 10g of calcium metal reacts with dilute sulphuric acid which is in excess. CaSO_4 and hydrogen gas are the products.
- Write down a balanced equation?
 - Work out the number of moles of calcium metal used up?
 - Work out the moles of sulphuric acid needed?
 - work out the volume of hydrogen gas evolved at STP?
- (iii) The formula of hydrated sodium carbonate is $\text{Na}_2\text{CO}_3 \cdot \text{H}_2\text{O}$ where X represents the number of molecules of water of crystallisation. To determine the value of X, a sample of hydrates sodium carbonate crystals was heated to constant mass. The results obtained were as follows:
- | | |
|--|-------|
| Mass of hydrated sodium carbonate before heating | 7.15g |
| Mass of sodium carbonate after heating | 2.65g |
- Calculate the decrease in mass?
 - How many moles of water were lost in the experiment?
 - How many moles of anhydrous sodium carbonate were formed at the end of the experiment?
 - Now work out the value of X in the formula.

e. MOLAR VOLUME

DEFINITION: It is the volume which 1 mole of a gas occupies at standard temperature and pressure and it is 22.4dm^3 .

By standard temperature we mean 0°C . Standard pressure is 1 atm.

There is also molar volume of a gas at room temperature and pressure and it is 24dm^3 .

Room/Normal temperature and pressure is 25°C and 1 atm.

EXERCISE 6d

Consider ethane burning completely in air.

1. Work out a balanced chemical equation
2. Work out the volume of CO_2 evolved at STP.
3. Work out the volume of CO_2 evolved when 70g of ethane are used.

f. EMPIRICAL FORMULA

Definition: It is the simplest formula that shows the lowest ratio of the atoms that make up a compound.

- Glucose molecule $\text{C}_6\text{H}_{12}\text{O}_6$ has the empirical formula CH_2O
- Ethanoic acid CH_3COOH can be rewritten $\text{C}_2\text{H}_4\text{O}_2$. Its empirical formula is CH_2O .
- Butene is C_4H_8 . Its empirical formula is CH_2 .

Working out empirical formula given grams of elements making up a compound

- ✓ Work out the number of moles.
- ✓ Work out the lowest mole ratio.

EXERCISE 6e

- i. In an experiment 2.6g of chromium were in excess of chlorine gas 7.93g of chromium chloride were formed. Work out the correct formula of chromium chloride and work out a balanced chemical equation for the reaction.
- ii. Calculate the empirical formula of an organic compound containing 48.0g of carbon, 12.0g of hydrogen and 32.0g of oxygen.

g. STANDARD SOLUTIONS

A. DEFINITION: A standard solution is a solution whose concentration is known

B. PREPARATION OF STANDARD SOLUTIONS

CASE STUDY:

PREPARATION OF 500cm³ OF 0.1 MOLAR SOLUTION FROM AN ANYHYDROUS SODIUM CARBONATE (Na₂CO₃).

a. APPARATUS/MATERIALS

- ✓ Triple beam balance for measuring exact masses of Na₂CO₃ crystals.
- ✓ Volumetric apparatus: 500cm³ volumetric flask, burette, pipette, pipette filler, clamp, clamp stand.

b. MATHEMATICAL CALCULATIONS

Working out the mass of Na₂CO₃ which must be dissolved in 500cm³ to give us a 0.1 Molar solution.

$$\text{Mass} = \text{Molarity} \times \text{RFM} \times \text{Volume}$$

$$\text{Molarity} = 0.1\text{M}$$

$$\text{RFM} = 46 + 12 + 48 = 106. \text{ (Can you figure out how this has been worked out)}$$

$$\text{Volume} = 500/1000 = 0.5\text{l}$$

$$\text{Mass} = 0.1\text{M} \times 106 \times 0.5 = 5.3\text{g}.$$

This means to prepare a 0.1Molar solution of Na₂CO₃ we have to dissolve 5.3g in 500cm³.

c. HOW TO WEIGH THE EXACT 5.3g of Na₂CO₃.

1. Heat the Na₂CO₃ gently in a dish to remove any absorbed water. Leave the dish in a dessicator. A dessicator is a closed glass container which contains a drying agent e.g. silica gel.
2. Weigh a clean dry bottle on a triple beam balance and record its mass as M₁. Using a spatula transfer the cool dried sodium carbonate into the weighing bottle until its mass is exactly 5.3g exactly greater than M₁.

d. DISSOLVING THE 5.3g SODIUM CARBONATE CRYSTALS IN 500cm³ WATER.

1. Transfer the 5.3g sodium carbonate crystals into a 100cm³ beaker.
2. Add warm distilled water enough to cover the solid sodium carbonate.
3. Stir the contents with a stirring rod for the solid sodium carbonate crystals to dissolve completely.
4. Clean the stirring rod with distilled water and wash it into the 100cm³ beaker.
5. Use a funnel to transfer the solution and the washings from the beaker into a clean dry 500cm³ volumetric flask.
6. Use a special glass stopper to close the volumetric flask at the top.
7. **Shake well** the contents for complete mixing.

h. TITRATION

DEFINITIONS:

- ✓ **Titration** is the gradual addition of one liquid to another.
- ✓ **End Point/Neutral Point:** It is the balanced point whereby an acid is completely neutralised by a base. The neutral point is *shown by colour change using an indicator.*
- ✓ **Titre:** This is the liquid that is gradually added to a standard solution

ACID-BASE INDICATORS WITH CORRESPONDING COLOUR CHANGES

INDICATOR	COLOUR IN ACID	COLOUR IN BASE
Blue Litmus	Red	Blue
Red Litmus	Red	Blue
Methyl orange	Pink	Yellow
Methyl red	Red	Yellow
Phenolphthalein	Colourless	Pink
Universal Indicator	Red, orange, yellow	Blue, violet, purple
Bromothymol Blue	Yellow	Blue

Table 6.1 shows effect of acids and bases to indicator

WORKING OUT THE MOLARITY OF A SOLUTION GIVEN A STANDARD SOLUTION

Formula:
$$\frac{M_1 V_1}{N_1} = \frac{M_2 V_2}{N_2}$$

DETERMINING THE MOLARITY OF A SOLUTION USING TITRATION GIVEN A STANDARD SOLUTION

CASE STUDY:

Give 2 Molar solution of sodium hydroxide how can you determine the molarity of HCl.

APPARATUS

Titration Apparatus.

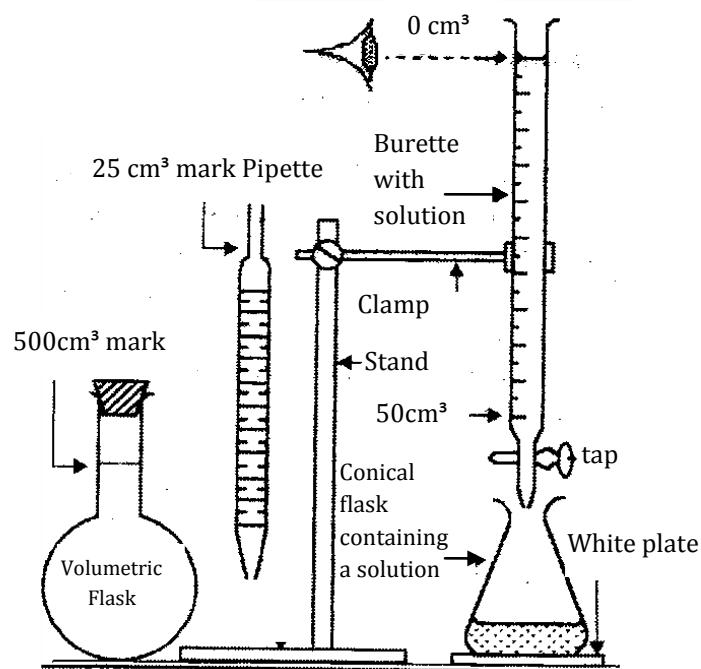
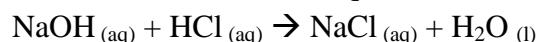


Fig. 6.1. The Titration Apparatus

PROCEDURE

- ✓ Measure 20cm³ NaOH solution and pour it in a conical flask.
- ✓ Add two drops of phenolphthalein indicator to 20cm³ NaOH solution. The colour changes to pink.
- ✓ Put HCl into the burette up to the zero mark
- ✓ Open the tap to add HCl solution slowly to the NaOH solution. Gently swirl around the flask to help the contents to mix well.
- ✓ Observe the colour change. When the last drop of acid from the burette neutralises the base in the flask the colour changes from pink to colourless. This is called END POINT.
- ✓ Work out a balanced chemical equation.



- ✓ Use the formula $\frac{M_1 V_1}{N_1} = \frac{M_2 V_2}{N_2}$ to work out the molarity of the acid.

SOURCES OF ERROR IN TITRATION

1. Misjudging the colour of the indicator near the end point. Different people have different sensitivity to colours.
2. Misreading the volume
3. Using contaminated solutions
4. Using wrong amount of indicator.
5. Not filling the burette properly
6. Titrating at wrong temperature
7. Vigorous swirling making the solution to splash from the titration conical flask leading to losing the solution.
8. Leaking burette.

EXERCISE 6f

- i. Calculate the volume of 0.1M sodium hydroxide that is needed to neutralise 20cm^3 of 0.1 M hydrochloric acid.
- ii. 20cm^3 of potassium hydroxide solution completely neutralise 40cm^3 of 0.1 M sulphuric acid. Work out the molarity of the potassium hydroxide solution.
- iii. Given 0.25 Molar solution of sodium carbonate. Describe how you can find out the molarity of $\text{HCl}_{(\text{aq})}$ by titration.

Chapter seven

ELECTROLYSIS

O. INTRODUCTION

DEFINITION: *Electrolysis is the process where electric current in a form of ions flows through a liquid and decomposes (breaks down) the liquid.*

ELECTROLYTE is a substance that conducts electricity by means of ions either in solution or when molten (fused).

Why are liquid metals non electrolytes?

They conduct electricity by means of electrons and not ions.

CONDUCTIVITY is the ability of a substance to allow the flow of electric current.

ELECTRIC CURRENT is made by the uniform flow of electrons in a wire. In a series circuit, electric current is the same at all points. Before an electric current can flow in a circuit, the circuit must be complete. If an ammeter shows a current reading, then the circuit is said to be complete.

1. CONDUCTIVITY CIRCUIT or The ELECTROLYSIS APPARATUS

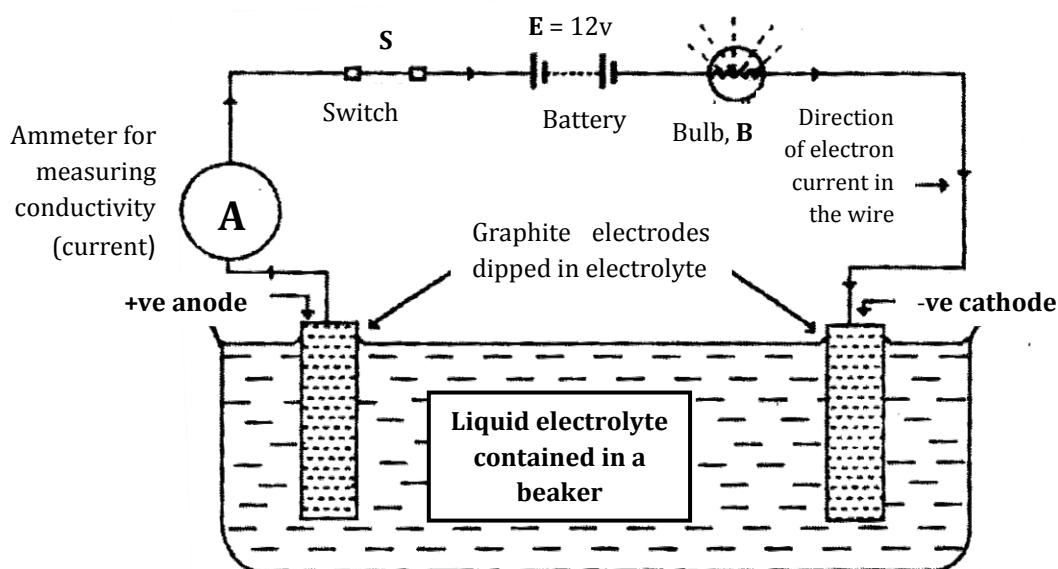


Fig 7.1 Conductivity Circuit

PARTS AND FUNCTIONS OF A CONDUCTIVITY CIRCUIT

1. **The Cell/Battery:** Gives a push-pull effect on electrons. Electrons are repelled on the negative terminal and attracted towards the positive terminal. The battery pumps out electrons through the bulb to light it up and through the ammeter to show the reading.
2. **Ammeter**
 - i. Measures the current of a complete circuit
 - ii. Indicates the conducting ability of the electrolyte

Advantages of an ammeter over a bulb

- iii. Gives accurate readings
 - iv. Measures small amounts of electric current
3. **Bulb:** Shows the presence of electric current and also its rough size.
4. **Anode:** It is a positive electrode. It is made positive by its connection to the positive side of the battery.
5. **Cathode:** It is the negative electrode. It attracts positively charged ions called Cations.
6. **Electrodes**

FUNCTION OF ELECTRODES

- (i) Providing a place where the products of electrolysis can be formed.
- (ii) Providing a place in the electrolyte liquid where the electrons can be given off or taken in, from the external circuit.

USES OF THE CONDUCTIVITY CIRCUIT

- a. To show that electric current flows in all parts of the circuit.
- b. To find the kind of electrolyte present in the beaker from the reading on the ammeter. Ammeter reading shows the conductivity of the liquid in the beaker.
- c. To act as an apparatus for electrolysis.

2. IONIC THEORY

It uses ions and atoms to explain what happens when electric current is passed through a solution.

IONISATION is the breaking free of ions from their compounds.

MSCE Chemistry does not clearly differentiate ionisation and disassociation. Strictly speaking **DISASSOCIATION** is the breaking free of ions from ionic compounds. The ions are already present. Ionisation is when polar covalent compounds break their bonds to make ions present.

3. ELECTROLYSIS OF LEAD BROMIDE (PbBr_2)

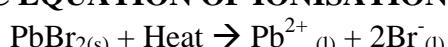
Lead bromide solid powder does not conduct electricity.

WHY? The ions are held together by strong ionic bonds and are not free to move.

Fused lead bromide conducts electricity.

WHY? Melting lead bromide causes the ions to break free as it turns to liquid. Heat energy breaks the strong ionic bonds and the ions are free to move.

IONIC EQUATION OF IONISATION OF LEAD BROMIDE



IONS PRESENT IN THE BEAKER

Positive lead ions (cations), Pb^{2+} and negative bromide ions (anions), 2Br^-

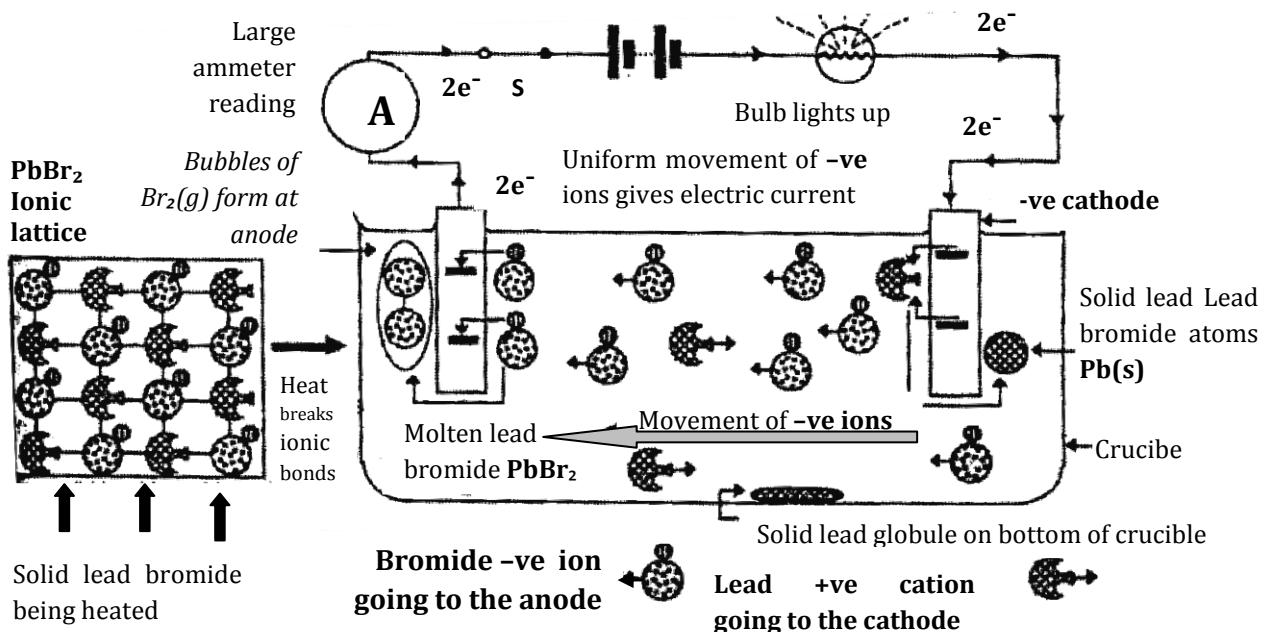


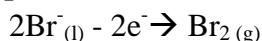
Fig 7.2. Conductivity circuit with fused lead bromide as electrolyte

OBSERVATIONS: When the switch is closed

i. BROMINE GAS FORMS AT THE ANODE.

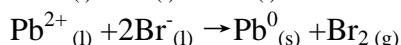
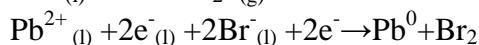
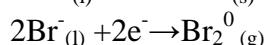
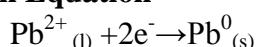
Bromide ions are attracted to the anode because they are oppositely charged to the anode and unlike charges attract. They give up their electrons to the anode to form bromine atoms. Two bromine atoms form one molecule of bromine gas.

Half equation at the anode



The two electrons given up by bromine ions go to the external circuit and reach the cathode.

Overall Equation



ii. AMMETER GIVES A READING

The electrons given up at the anode are attracted to the positive end of the cell and flow up the copper wire in the external circuit. As the electrons flow through the ammeter gives a reading.

iii. BULB LIGHTS UP

The cell pumps out the electrons through the bulb. As electrons flow through the bulb, it lights up.

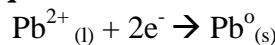
iv. LEAD DEPOSITS FORM AT THE CATHODE

Positively charged lead ions are attracted to the cathode

WHY? They are unlike charges.

Two electrons arriving at the cathode join with a lead ion to form lead metal.

Half equation at the cathode



Definition of Half Equation

A **half equation** is an equation showing the chemical reaction taking place at an electrode.

4. CLASSIFICATION OF ELECTROLYTES

i. STRONG ELECTROLYTES

- ✓ Pass a large current in the conductivity circuit.
- ✓ Have many free ions present

EXAMPLES:

Nitric acid in water, sulphuric acid in water, alkali hydroxides, NaCl in water, Fused salts, fused acids, fused bases, fused lead bromide, fused sodium chloride, fused calcium oxide, fused aluminium oxide.

ii. WEAK ELECTROLYTES

- ✓ Pass small current through the conductivity circuit
- ✓ Have few free ions.

EXAMPLES:

Distilled water, Hydrofluoric acid in water, organic acids, silver chloride in water and lead bromide in water.

5. NON ELECTROLYTES

- ✓ Pass zero current through the conductivity circuit.
- ✓ There are no ions present.

EXAMPLES:

hydrocarbons, liquid non metals, liquefied hydrogen chloride gas.

Note: that liquid metals conduct electricity but they are non electrolytes because they use electrons and not ions.

6. FACTORS AFFECTING CONDUCTIVITY OF ELECTROLYTES

IONIC BOATS PICTURE: Here we speak of ions in terms of ionic boats. We speak of the liquid electrolyte in terms of a river. The ionic boats have to cross the electrolyte river from one electrode to the other to bring about conductivity.

i. CONCENTRATION

A more concentrated electrolyte has more ionic boats. The number of charges crossing the electrolyte river per second is increased thereby increasing the conductivity.

ii. DISTANCE BETWEEN THE ELECTRODES

The magnitude of an electric current depends on the speed at which the ionic boats travel across the electrolyte river. SPEED = No of ions per time taken. Moving the electrodes closer decreases the time taken for the charges and increases the speed making conductivity higher than before. Decreasing the distance between electrodes does not make ions move physically faster.

iii. NUMBER OF CELLS

Increasing the number of cells increases the push-pull effect. The ions move physically faster. More ionic boats arrive at the electrode per second. This increases conductivity.

iv. DEPTH OF ELECTRODES

If the electrodes are lowered deeper into the electrolyte, ions have more space to unload their electrons. More ionic boats cross the electrolyte river per second. This increases conductivity.

v. SOLUBILITY IN WATER

Partially soluble compounds ionise partially. Water breaks up few ions from the ionic lattice and few free ions are present making conductivity low. Soluble electrolytes almost ionise completely in water leading to high conductivity.

SOLUBLE ELECTROLYTE: Nearly complete ionisation – High conductivity – strong electrolyte. SIGN: Single arrow →

LOW SOLUBILITY ELECTROLYTE: partial ionisation – low conductivity – weak electrolyte. SIGN: Double arrow ⇤

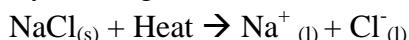
7. ELECTROLYSIS OF SODIUM CHLORIDE (NaCl)

Explain why solid sodium chloride does not conduct electricity?

How can one make solid sodium chloride conduct electricity?

A. IONIC EQUATIONS OF IONISATION

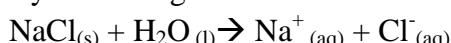
I. By Heating



Heat increases the kinetic energy of the NaCl particles and weakens the ionic bonds. Then the solid NaCl changes its physical state from solid to liquid, hence ionisation.

**Melting point of NaCl is 801°C

II. By dissolving NaCl in water



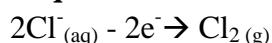
How does water break the strong ionic bonds in NaCl at room temperature?

Water is a polar covalent substance because the oxygen atom is more electronegative than hydrogen and so water has partial charges. Hydrogen is delta positive and oxygen is delta negative. Dissolving NaCl in water, sets up electrostatic forces of attraction. The delta positive hydrogen end attracts the chloride ion. The delta negative oxygen end attracts the positive sodium ion. These electrostatic forces of attraction break up the ionic bonds in NaCl to cause ionisation.

OBSERVATIONS

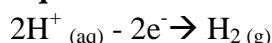
- a. AT THE ANODE: Chlorine gas is formed.

Half equation at the anode



- b. AT THE CATHODE: Hydrogen gas is formed

Half equation at the cathode



B. THE PREFERENTIAL DISCHARGE OF HYDROGEN IONS AT THE CATHODE

- In all aqueous solutions water ionises a little. The ionic equation of the ionisation of water is $\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq})$
- Hydrogen ions and hydroxide ions are also present in the solution in addition to sodium ions and chloride ions.
- Both sodium and hydrogen ions are present at the cathode. Hydrogen ions are preferentially discharged at the cathode because they need less energy to discharge than sodium ions.

FACTORS AFFECTING PREFERENTIAL DISCHARGE OF IONS

I. CONCENTRATION

The ion with highest concentration can be preferred. In the electrolysis of weak NaCl solution oxygen gas is discharged at the anode instead of chlorine gas. Electrolysis of dilute solutions is actually electrolysis of water.

II. THE ION WITH LOWEST ENERGY FOR DISCHARGE.

This can be likened to *kudzera njira yachidule* (taking a short path) to a destination.

III. NATURE OF ELECTRODES

In the electrolysis of CuSO₄, oxygen gas forms at the anode if the apparatus uses carbon electrodes. When the electrodes are replaced by pure copper metal rods, the anode dissolves.

Note: Preferential discharge only takes place in aqueous solutions. In fused melts there is no choice of ions.

RULES FOLLOWED IN PREFERENTIAL DISCHARGE

At the cathode

- Hydrogen ions accept electrons to form hydrogen gas when competing with more reactive metals. More reactive metals ‘like’ to exist as ions.
- Less reactive metals will preferentially be discharged ahead of hydrogen ions.

At the anode

- If halogen ions are not present, hydroxide ions give up electrons more readily than other non-metal ions.

8. ELECTROLYSIS OF DILUTE SULPHURIC ACID (H_2SO_4)

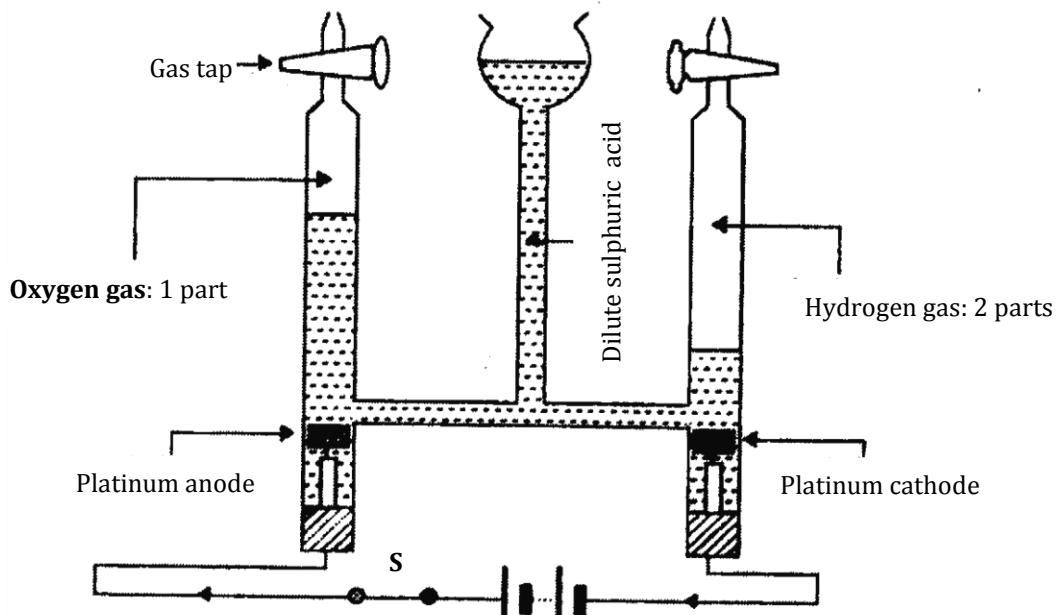
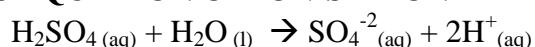
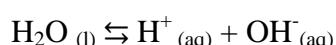


Fig 7.3. The Hoffman Voltameter

IONIC EQUATION OF IONISATION



Partial ionisation of water



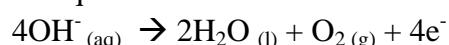
THE IONS PRESENT ARE

- Hydrogen ions, sulphate ions and hydroxide ions

OBSERVATIONS

I. OXYGEN GAS IS PRODUCED AT THE ANODE.

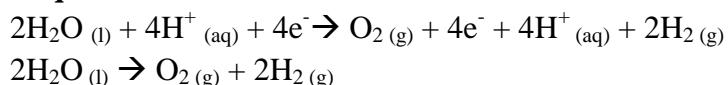
Explanation: There is preferential discharge of hydroxide OH^- ions according to the half equation



II. HYDROGEN GAS FORMS AT THE CATHODE.

The four electrons go to the external circuit and reach the cathode. Hydrogen ions inside the conductivity circuit are attracted to the anode. There they gain the electrons and form hydrogen gas according to following half chemical equation: $4\text{H}^+_{(\text{aq})} + 4\text{e}^- \rightarrow 2\text{H}_2_{(\text{g})}$

Overall Equation



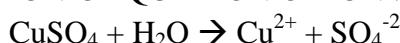
The overall reaction shows that it is electrolysis of water. Volumes of hydrogen gas and oxygen gas are collected in the ratio 2:1. As a rule, electrolysis of dilute solutions is electrolysis of water.

III. BULB LIGHTS UP.

IV. AMMETER GIVES A READING.

9. ELECTROLYSIS OF COPPER SULPHATE (CuSO_4)

IONIC EQUATION OF IONISATION



The ions present in the beaker are copper ions and sulphate ions.

I. USING CARBON ELECTRODES

OBSERVATIONS:

- OXYGEN GAS IS PRODUCED AT THE ANODE.

Explanation: There is preferential discharge of hydroxide, OH^- , ions according to the half equation $4\text{OH}^-_{(\text{aq})} - 4\text{e}^- \rightarrow 2\text{H}_2\text{O}_{(\text{l})} + \text{O}_2_{(\text{g})}$

- COPPER METAL FORMS AT THE CATHODE.

Half Equation at the cathode: $2\text{Cu}^{2+}_{(\text{aq})} + 4\text{e}^- \rightarrow 2\text{Cu}^0_{(\text{s})}$

- BULB LIGHTS UP BUT BECOMES DIMMER WITH TIME. Copper ions become fewer with time as the electrolyte is used up. The decrease in concentration decreases the number of ionic boats reaching the electrode per second. This decreases conductivity.

- AMMETER GIVES A READING WHICH DECREASES WITH TIME.

- THE SOLUTION BECOMES PALER BLUE. As the liquid electrolyte is used up the copper ions decrease in concentration.

- With time the cathode preferentially discharges hydrogen gas because the copper ions decrease in concentration.

EXERCISE 7a

What is responsible for the blue colour of copper sulphate solution?

II. USING COPPER ELECTRODES

OBSERVATIONS:

- a. THE ANODE WHICH IS A COPPER ELECTRODE DISSOLVES.
This is an example of preferential discharge. It is the lowest energy route for the reaction to follow.
Half Equation at the anode: $\text{Cu}^0_{(s)} - 2\text{e}^- \rightarrow \text{Cu}^{2+}_{(aq)}$
- b. COPPER DEPOSITS FORM AT THE CATHODE
Half Equation at the cathode: $\text{Cu}^{2+}_{(aq)} + 2\text{e}^- \rightarrow \text{Cu}^0_{(s)}$
- c. THE SOLUTION STAYS BLUE AS ELECTROLYSIS IS TAKING PLACE.
The copper ion concentration stays constant. The rate at which copper ions are being used up at the cathode is equal to the rate at which the copper ions are being replaced by the dissolving anode.
- d. AMMETER GIVES A READING WHICH STAYS THE SAME.
- e. BRIGHTNESS OF THE BULB REMAINS THE SAME

10. ELECTROLYSIS OF COVALENT SUBSTANCES

Covalent substances are non-electrolytes. They have no ions. However some covalent substances ionise when dissolved in polar solvents.

BUBBLING HYDROGEN CHLORIDE GAS THROUGH WATER.

REVISION QUESTIONS

How do hydrogen and chlorine combine to form hydrogen chloride gas?

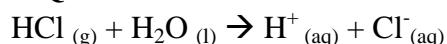
What makes hydrogen chloride gas a polar covalent molecule?

When $\text{HCl}_{(g)}$ is bubbled into water, electrical attractions are set up between the oppositely charged ends of the polar water molecule and $\text{HCl}_{(g)}$ polar molecule.

The delta positive hydrogen end of the water molecule faces inwards towards the delta negative chloride end of $\text{HCl}_{(g)}$. The delta negative oxygen end of the water molecule faces inwards towards the delta positive hydrogen end of the $\text{HCl}_{(g)}$ molecule.

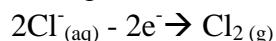
The $\text{HCl}_{(g)}$ molecule is pulled apart and the bond breaks. The shared electrons remain with chlorine making it a chloride anion. Hydrogen remains with no electrons and becomes a hydrogen cation.

IONIC EQUATION OF IONISATION

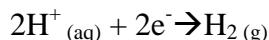


OBSERVATIONS

- a. Chlorine gas forms at the anode according to the half equation



b. Hydrogen gas forms at the cathode according to the half equation



- c. Ammeter gives a high reading and the bulb lights up. These indicate that $\text{HCl}_{(\text{aq})}$ is a strong electrolyte.
- d. There is temperature increase indicating that the reaction is exothermic. The making of new bonds liberates heat to the surrounding.
- e. The solution turns blue litmus paper red. The presence of hydrogen ions makes the solution acidic.
- f. The solution reacts with magnesium metal to give hydrogen gas. Acids dissolve metals which are higher than hydrogen on the electrochemical series.

**When hydrogen chloride is bubbled through a non-polar covalent liquid no chemical reaction takes place.

11. CONDUCTIVITY CHANGES IN A TITRATION

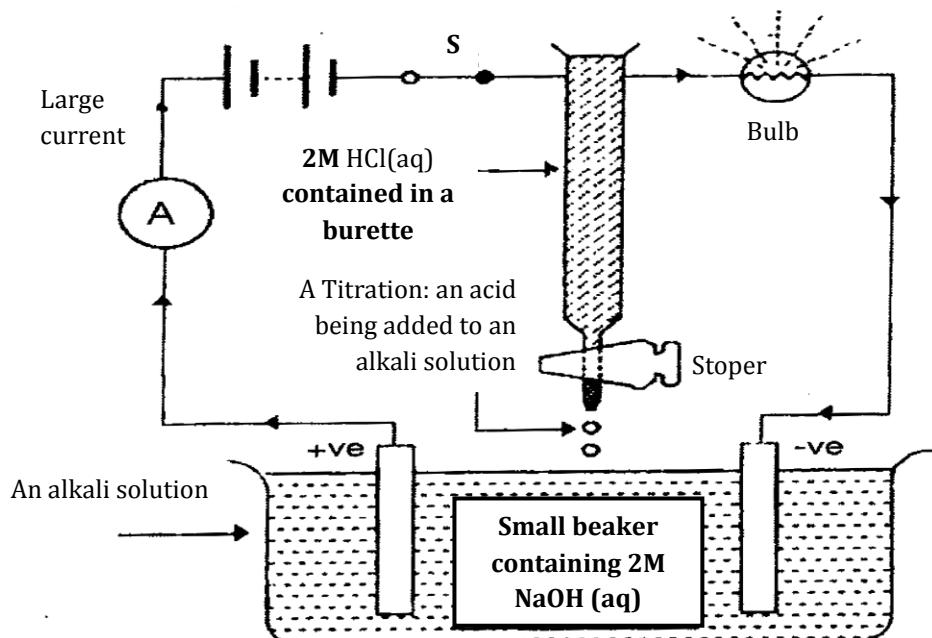


Fig 7.4. Experiment setup for conductivity changes in a Titration

- 20cm^3 of 2M $\text{NaOH}_{(\text{aq})}$ solution is measured out exactly using a pipette and transferred into a 100cm^3 beaker.
- Two drops of phenolphthalein are added to give a pink colour showing the presence of a base.
- The solution is placed in a conductivity circuit by dipping the electrodes into the solution and closing the switch.
- The ammeter reading is noted.

- HCl _(aq) of the same 2M concentration is then added slowly from the burette. After every addition of the HCl _(aq) the contents are swirled around to make sure of good mixing of acid and base.
- The ammeter reading is taken

OBSERVATIONS:

1. When 20cm³ of HCl _(aq) is added to the NaOH _(aq) there is a sudden colour change from pink to colourless. This is called the END POINT of the titration.
2. Ammeter reading decreases with the addition of HCl _(aq) up to the end point.
3. Further addition of HCl _(aq) increases ammeter reading.

TABLE OF RESULTS

ACID VOLUME (cm ³)	0	5	10	12	15	18	20	21	22	24	25
CONDUCTIVITY (mA)	880	840	800	780	760	730	710	740	800	860	900

Table 7.1 Increase in acid volume and change in conductivity

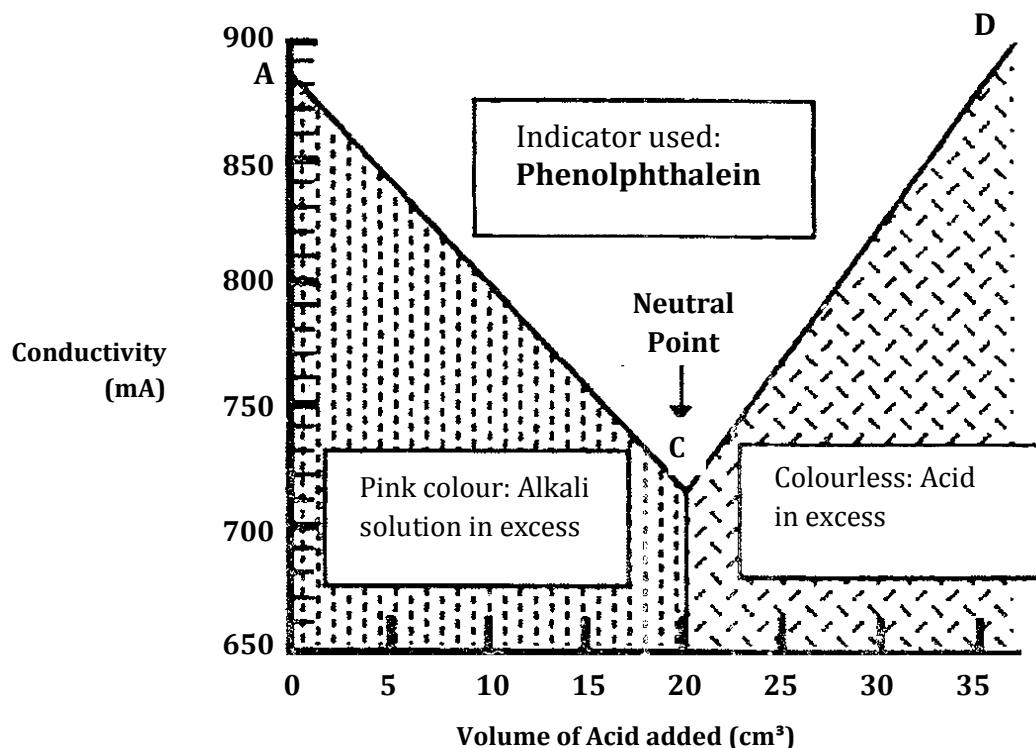
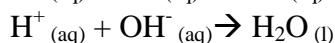
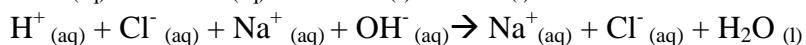
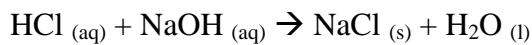


Fig 7.5. Graph of Conductivity(mA) against Volume HCl _(aq) added into a beaker from the burette in a titration (Results in Table 7.1. above).

- ✓ Point A shows the greatest current reading to start with. Ionisation of NaOH produces ions. The highest concentration of the ions is at A
- ✓ Addition of the HCl reduces the concentration of ions in the solution. The net ionic equation is that of the formation of water
ACID + BASE → SALT + WATER



Water is a weak electrolyte and ionises very little. Therefore the ion concentration is decreasing. $\text{Na}^+_{(\text{aq})}$ and $\text{Cl}^-_{(\text{aq})}$ ions are present in the beaker but do not take part in the reaction. They are spectator ions.

- ✓ Point C is the end point
- ✓ After point C the acid is in excess and the hydrogen ions are increasing in concentration.

12. ELECTROPLATING

Electrolysis is important in electroplating and metal ore purification processes

Definition: ELECTROPLATING is the covering of one metal by another by means of electrolysis.

CASE STUDY:

ELECTROPLATING AN IRON NAIL WITH COPPER

APPARATUS: Conductivity circuit with copper anode, iron nail cathode and copper chloride electrolyte.

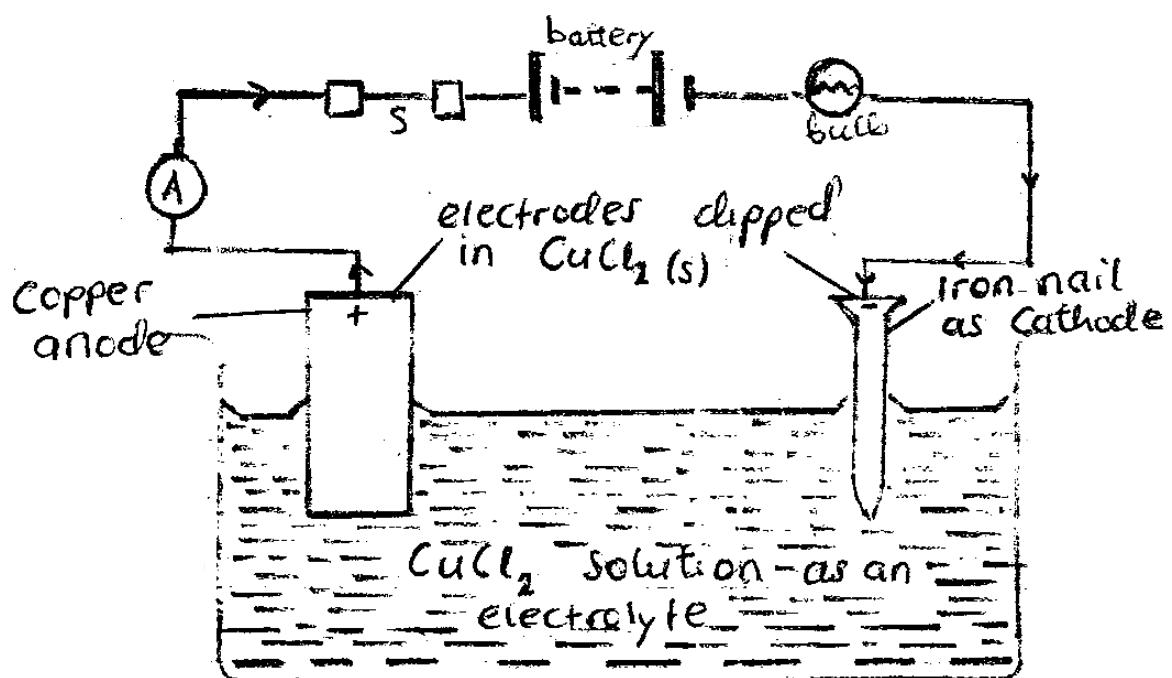
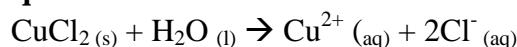


Fig 7.6. electroplating of an iron nail with copper metal

From my student, 2012 Pius XII Seminary: "Learn from this to draw with creativity"

Ionic Equation of ionisation



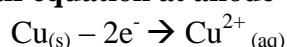
THE IONS PRESENT IN THE BEAKER ARE

- copper ions and chloride ions.

OBSERVATIONS:

- a. THE COPPER ANODE DISSOLVES. This is the lowest energy ionic route for the reaction to follow.

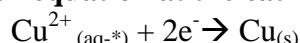
Half equation at anode



The two electrons go through the external circuit to the cathode. The copper cation goes to the cathode.

- b. COPPER COATING AT IRON NAIL CATHODE.

Half equation at the cathode



The copper ions gain the two electrons at the cathode forming copper metal. The copper metal coats the iron nail.

- c. AMMETER GIVES A READING
- d. BULB LIGHTS UP.

CONDITIONS FOR ELECTROPLATING

- (i) The metal to be covered by another metal is made into the cathode
- (ii) The metal to cover the other metal made into the anode
- (iii) The electrolyte must be a soluble salt of the same metal that makes up the anode.

ADVANTAGES OF ELECTROPLATING

- (i) Prevents rusting/corrosion of metals.
- (ii) Improves the appearance of the metal.
- (iii) Gives extra thickness to worn out parts of machinery.

13. THE PURIFICATION OF COPPER

What is an ore?

An **ore** is a naturally occurring mineral from which a metal can be extracted

E.g. Malachite (CuCO_3) is a copper ore. It is green in colour and is sold for decoration. Electrolysis is used to obtain the copper metal from malachite.

APPARATUS FOR THE PURIFICATION OF COPPER

Conductivity circuit:

- Impure copper is made the anode.
- The cathode is made of pure copper.
- The liquid electrolyte is any soluble salt of copper e.g. CuSO_4 .

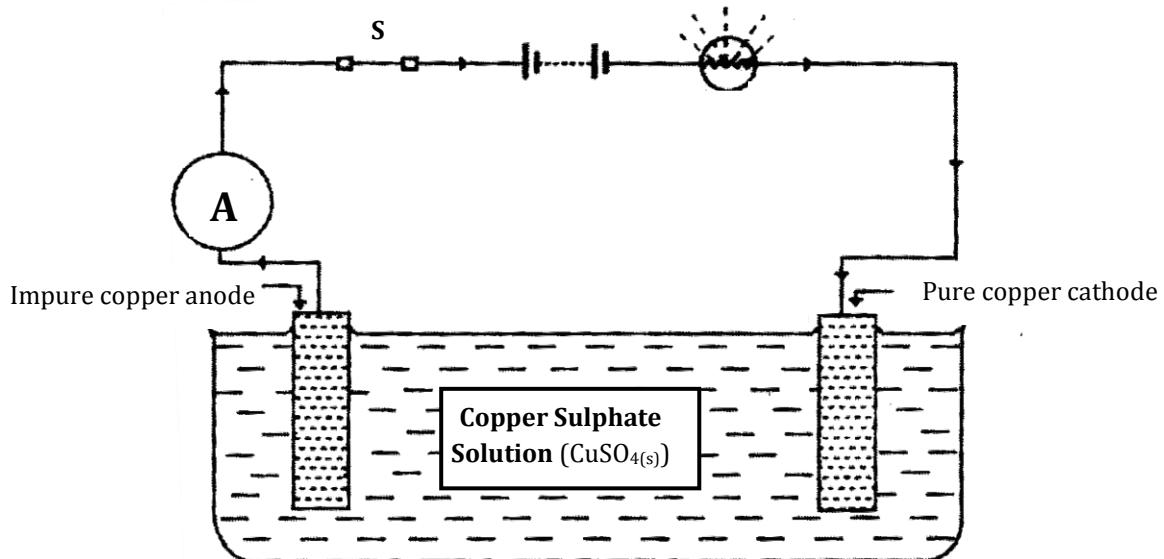
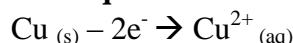


Fig 7.6. Apparatus for the purification of copper or electroplating of copper

OBSERVATIONS:

- THE COPPER ANODE DISSOLVES. This is the lowest energy ionic route for the reaction to follow.

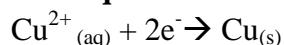
Half equation at anode



The two electrons go through the external circuit to the cathode. The copper cation goes to the cathode.

- COPPER COATING AT PURE COPPER CATHODE.

Half equation at the cathode



The copper ions gain the two electrons at the cathode forming copper metal. The copper metal coats the copper metal.

- DIRT IMPURITIES FALL OFF THE ANODE and form sludge on the bottom of the container. Formation of sludge is evidence that purification is taking place.

NOTES

- This is electrolysis of copper using copper electrodes
- Work out the overall equation for the reaction!
- It is simply copper atoms being transferred from the anode to the cathode.
 $\text{Cu}^0 \rightarrow \text{Cu}^0$
- We work out the overall equation by adding up the half equations.
- Solid calcium chloride and copper sulphate are dissolved in water separately.

Chapter eight

PRECIPITATION

O. INTRODUCTION

DEFINITIONS:

- A **precipitate** is a low solubility solid formed from the mixing of two strong electrolytes.
- **Spectator Ions** are ions present in a chemical reaction but take no part in the reaction.

1. OBSERVATION EXPERIMENT

PROCEDURE

- i. Dissolve solid lead nitrate, $\text{Pb}(\text{NO}_3)_2$ (s), in water. This gives a colourless solution.
- ii. Dissolve potassium iodide in water. This also gives a clear colourless solution.
 - a. Write down the ionic equations of ionisation for both lead nitrate and potassium iodide.
 - b. Name the ions present in the two beakers.
- iii. Pour the two solutions into a third beaker

OBSERVATION

- A thick yellow opaque muddy substance forms. This is called a precipitate.

How does the precipitate form?

Inside the beaker the lead ions, nitrate ions, potassium ions and iodide ions collide. The collisions between the oppositely charged ions might have caused the ions to stick together because of their attractions. This leads to an exchange of partners presented diagrammatically below

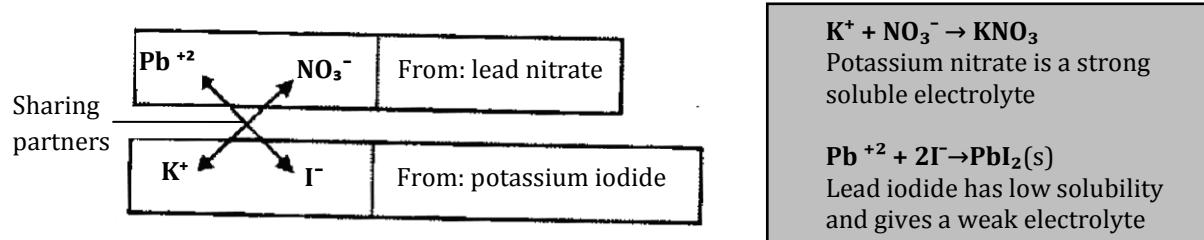
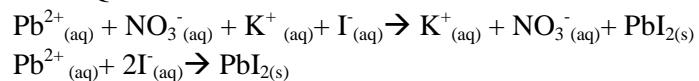


Fig 8.1. Diagram showing exchange of partners in formation of a precipitate and the chemical equations involved in the process

- Partners giving a soluble product will form a strong electrolyte and give no precipitate. Partners giving an insoluble product will stick together, form a weak electrolyte and give a precipitate.
- Read the solubility rules to name the precipitate and spectator ions formed.

NET IONIC EQUATION



2. SOLUBILITY RULES

- i. All group one and ammonium salts are soluble
- ii. All nitrates are soluble
- iii. All chlorides, iodides and bromides are soluble except silver and lead.
- iv. All sulphates are soluble except calcium and barium.
- v. Carbonates and hydroxides have low solubility except sodium, potassium and ammonium.

EXERCISE 8a

Solid barium chloride and sodium sulphate are dissolved in water separately. The solutions are then mixed in a third beaker. Predict whether or not a precipitate will form. Write down the net ionic equation.

Chapter nine

ELECTRON TRANSFER

O. DISPLACEMENT REACTIONS IN A BEAKER

OBSERVATION EXPERIMENTS

MATERIALS: Copper, zinc, iron and magnesium metals, copper sulphate, magnesium sulphate, zinc sulphate and iron sulphate, beakers distilled water and a measuring cylinder.

PROCEDURE:

- ✓ We poured about 2cm³ of copper sulphate into four identical beakers
- ✓ We placed each of the four metals into the four beakers containing copper sulphate solution
- ✓ We observed for about five minutes.
- ✓ We repeated the above procedures with solutions of zinc sulphate, magnesium sulphate and iron sulphate.

OBSERVATIONS – These are summarised in the table of results.

	Copper	Zinc	Iron	Magnesium
CuSO ₄				
ZnSO ₄				
FeSO ₄				
MgSO ₄				

Table 9.1 shows results of the experiments carried out

Record the results in the table by indicating ‘REACTION’ or ‘NO REACTION’

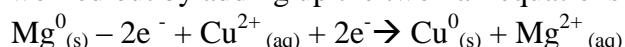
Arrange the metals in order of decreasing reactivity

Magnesium metal placed in CuSO_{4(aq)} gives ‘REACTION’ meaning

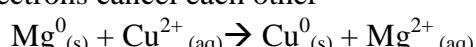
- ✓ The magnesium metal dissolves and forms Mg²⁺ ions according to the half equation Mg^{0(s)} – 2e⁻ → Mg²⁺.
- ✓ Copper metal coats magnesium rod. The copper ions receive the electrons and form copper metal according to the half equation Cu^{2+(aq)} + 2e⁻ → Cu^{0(s)}
- ✓ With time the solution becomes paler blue and finally the solution will be magnesium sulphate solution.

Overall Equation Ionic Displacement Equation

This is worked out by adding up the two half equations



The electrons cancel each other

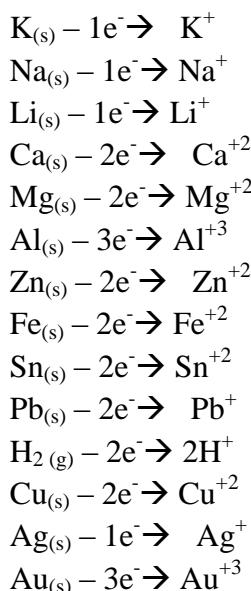


EXERCISE 9a

Work out the overall equation when calcium metal is placed in silver chloride solution

NOTES

- Any metal can donate electrons to another metal
- Any metal can gain electrons from another metal
- Continuing with displacement reactions using different metals and solutions leads us to a displacement table and the **Electrochemical Series (ECS)**.

1. THE ELECTROCHEMICAL SERIES (ECS)/ ACTIVITY SERIES**NOTES**

- ✓ Without the half equations the ECS is the displacement table of metals.
- ✓ Metals higher in the ECS will donate electrons to ions of metals lower down in the electrochemical series.
- ✓ To have a quick check on which pair of metal and ions will give a reaction we look for a top left- bottom right pattern.
- ✓ Metals near the top are group 1 and 2 members of the periodic table
- ✓ Hydrogen ions represent acidic solutions. Metals below hydrogen on the ECS will not react with acidic solutions.

2. DISPLACEMENT REACTIONS OF HALOGENS

More reactive halogens displace less reactive ones from their solutions.

EXERCISE 9b

- Arrange the halogens in order of decreasing reactivity
- Explain why chlorine is more reactive than bromine
- Predict what can happen if chlorine gas is bubbled into potassium iodide solution.

3. OXIDATION AND REDUCTION – REDOX REACTIONS

Oxidation means

- ✓ Gaining of oxygen atoms
- ✓ Loss of electrons
- ✓ Increase in charge
- ✓ Loss of hydrogen atoms

Reduction means

- ✓ Loss of oxygen atoms
- ✓ Gaining of electrons
- ✓ Decrease in charge
- ✓ Gaining of hydrogen atoms.

Redox Reactions

A REDOX REACTION is a chemical reaction where electron transfer occurs and oxidation and reduction take place simultaneously.

Examples of redox reactions

1. Displacement reactions in a beaker
2. Electrolysis
3. Electroplating
4. Purification of metal ores

4. REDUCING AGENT AND OXIDISING AGENT

- A **reducing agent** is a substance which facilitates reduction
- An **oxidising agent** is a substance which facilitates oxidation

Example: When zinc metal is placed in $\text{CuSO}_4 \text{ (aq)}$, the following things happen:

- zinc loses electrons and so is _____
- Cu^{2+} ions gain electrons and so are _____
- Zinc metal facilitates the _____ of Cu^{2+} ions and so is _____
- Cu^{2+} ions facilitate the _____ of zinc metal and so are _____

SUMMARISING REDOX REACTIONS

CHANGE	OXIDATION NUMBER	ELECTRONS ARE
Oxidation	Increased	Lost
Reduction	Decreased	Gained
Reducing agent	Increased	Lost
Oxidising agent	Decreased	Gained
Oxidised substance	Increased	Lost
Reduced substance	Decreased	Gained

Table 9.2. A Summary of Redox reactions

5. OXIDATION NUMBER

It is the charge on an ion.

EXERCISE 9c

- i. *Workout the molecular formula of aluminium nitrate*
- ii. *Calculate the oxidation number of:*
 - a. *Carbon in hydrogen carbonate ion, HCO_3^-*
 - b. *Sulphur in sulphur dioxide gas*

RULES WHEN WORKING OXIDATION NUMBERS

- ✓ All elements have an oxidation number of zero
- ✓ Oxygen has an oxidation number of -2.
- ✓ Hydrogen has an oxidation number of +1
- ✓ In neutral molecules the sum of separate charges inside the molecule must add up to zero.
- ✓ For a charged atom the sum of the separate charges inside the ion must be equal to the final overall charge on the ion itself.

6. ELECTROLYTIC CELLS

An electrolytic cell is formed when two different metals are placed in an electrolyte and joined by an outside wire.

EXAMPLES OF ELECTROLYTIC CELLS

I. An Orange Cell

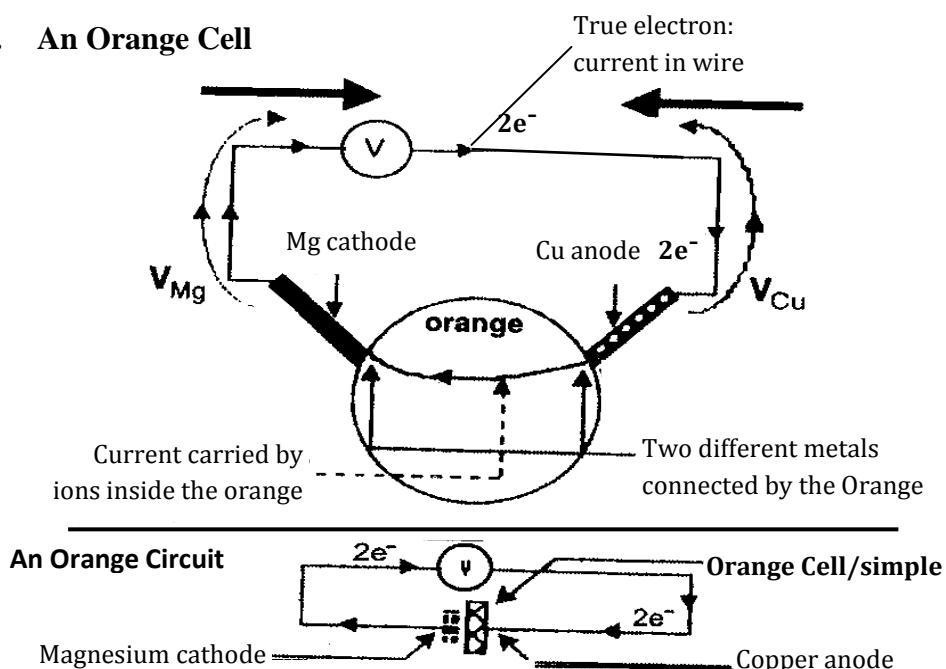


Fig 9.1. An Orange Cell

II. The Electrochemical cell (Daniell cell)

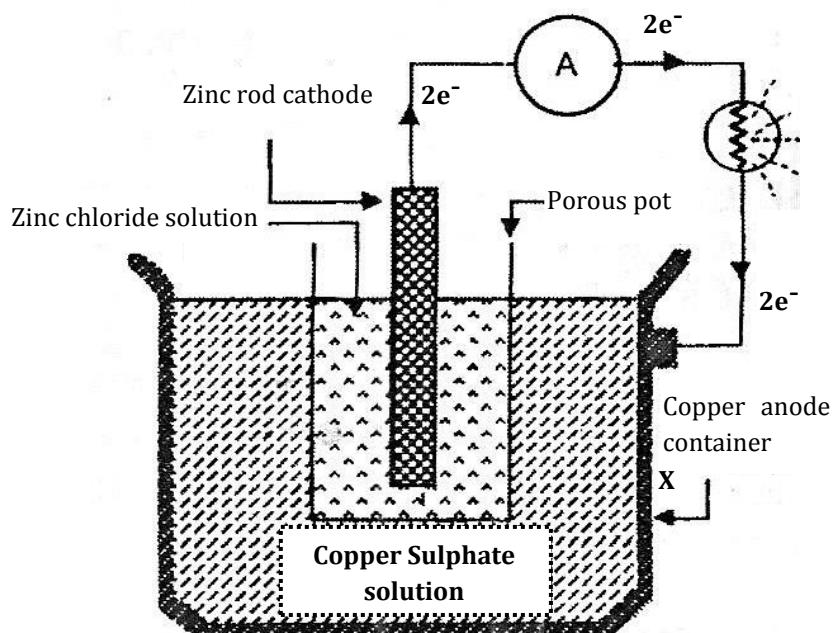


Fig 9.2. The Daniell cell

Results and their explanations

- Zinc rod dissolves according to the half equation $\text{Zn}_{(s)} - 2\text{e}^- \rightarrow \text{Zn}^{2+}_{(aq)}$. The ions pass into the solution and the electrons go to the outside wire which is also called the external circuit.
- Copper metal deposits form on the copper anode container. Copper ions in the CuSO_4 solution gain the two electrons giving a copper metal. Copper coats the copper anode container according to the following half equation

$$\text{Cu}^{2+}_{(aq)} + 2\text{e}^- \rightarrow \text{Cu}^0_{(s)}$$
- Current flows from zinc cathode to the copper anode container.

FUNCTIONS OF THE POROUS POT

- Preventing general mixing of the two liquid electrolytes.
- Allows the passage of ions between the two sections of the cell thereby completing the circuit.

CHIEF DIFFERENCE BETWEEN DISPLACEMENT REACTIONS IN A BEAKER AND AN ELECTROCHEMICAL CELL

- ❖ In the beaker electrons are handed directly while in the Daniell cell the electrons are handed through a wire.

7. VOLTAGE SERIES

$\text{K}_{(s)} - 1\text{e}^- \rightarrow \text{K}^+_{(\text{aq})}$	-2.92	Volts
$\text{Na}_{(s)} - 1\text{e}^- \rightarrow \text{Na}^+_{(\text{aq})}$	-2.71	Volts
$\text{Mg}_{(s)} - 2\text{e}^- \rightarrow \text{Mg}^{+2}_{(\text{aq})}$	-2.37	Volts
$\text{Al}_{(s)} - 3\text{e}^- \rightarrow \text{Al}^{+3}_{(\text{aq})}$	-1.67	Volts
$\text{Zn}_{(s)} - 2\text{e}^- \rightarrow \text{Zn}^{+2}_{(\text{aq})}$	-0.76	Volts
$\text{Fe}_{(s)} - 2\text{e}^- \rightarrow \text{Fe}^{+2}_{(\text{aq})}$	-0.44	Volts
$\text{Sn}_{(s)} - 2\text{e}^- \rightarrow \text{Sn}^{+2}_{(\text{aq})}$	-0.14	Volts
$\text{Pb}_{(s)} - 2\text{e}^- \rightarrow \text{Pb}^+_{(\text{aq})}$	-0.12	Volts
$\text{H}_2_{(\text{g})} - 2\text{e}^- \rightarrow \text{H}^{+2}_{(\text{aq})}$	0.00	Volts
$\text{Cu}_{(s)} - 2\text{e}^- \rightarrow \text{Cu}^{+2}_{(\text{aq})}$	0.34	Volts
$\text{Ag}_{(s)} - 1\text{e}^- \rightarrow \text{Ag}^+_{(\text{aq})}$	0.80	Volts

- Different metals joined together to form cells give different Electromotive Force (EMF). **EMF** is *the force that drives electrons*. Zn and Fe would give $-0.76V - 0.44V = -0.32V$. Zn and Cu would give $-0.76V - +0.34V = -110V$.
 - Metals far apart on the voltage series give large EMF
 - Metals higher in the ECS form the negative cathode that repels the electron current uphill to the positive anode.
 - The minus sign has no effect on the size of the reading. It only shows the direction of the voltmeter needle

8. POLARISATION

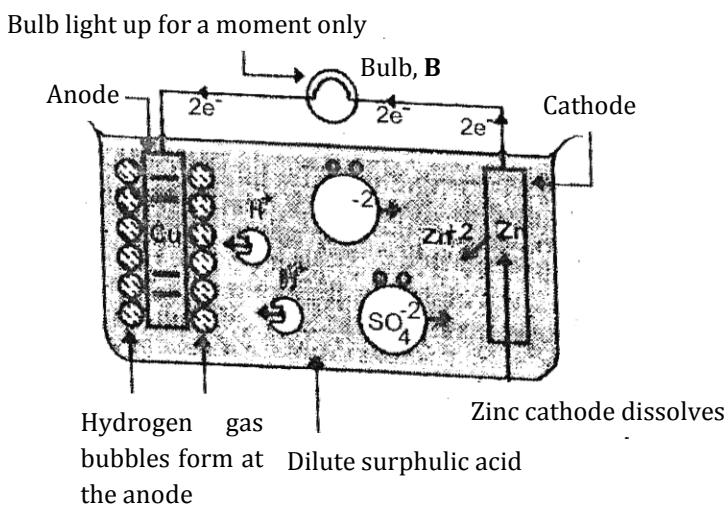


Fig 9.3. Diagram showing Polarisation

When a simple cell is connected up as shown above, the small bulb B glows for a short time and then goes out. Zinc cathode dissolves and hydrogen gas collects at the copper anode.

Polarisation is the effect which causes the current to die away in an electrochemical cell by the formation of an insulating layer of hydrogen gas bubbles on the anode.

DEPOLARISATION: Potassium dichromate (K_2CrO_4) if added to the solution makes the bulb stay alive for sometime. It is a depolariser. It supplies oxygen which combines with hydrogen gas to form water. Potassium dichromate works as an oxidising agent.

9. WET LECLANCHÉ CELL

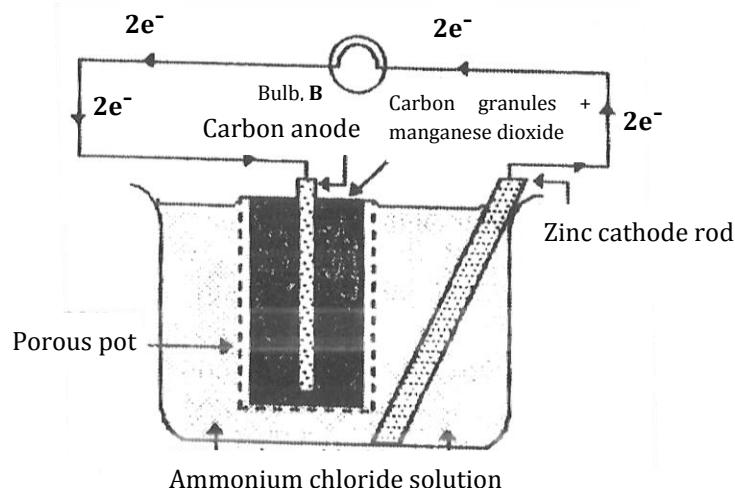


Fig 9.4. A Wet Laclanché Cell

Complete the following:

Electrolyte:

Cathode:

Anode:

Equation for the dissolving of zinc:

Ionisation of ammonium chloride:

Formation of ammonia gas and hydrogen gas

What causes polarisation?

How is depolarisation achieved?

10. DRY LECLANCHE CELL – A DRY CELL

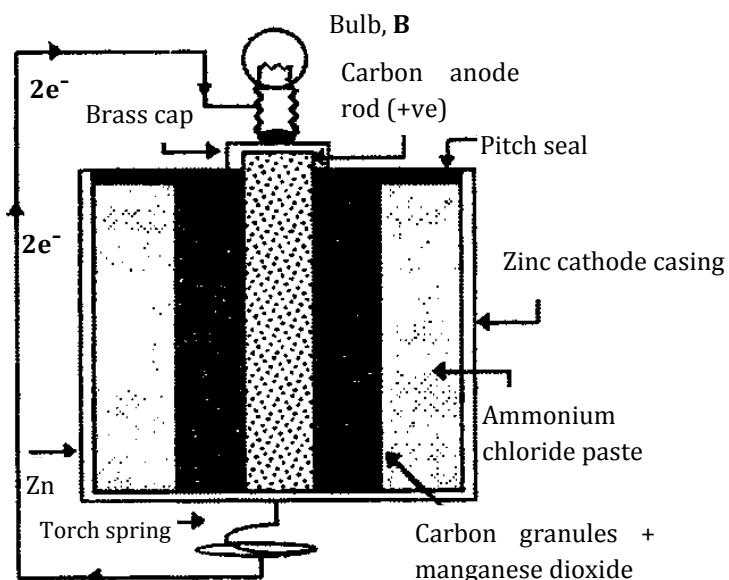


Fig 9.5. The Dry Laclanché Cell or a Torch battery

ANODE: carbon rod with brass cap at the top.

CATHODE: Zinc casing. Dissolves and releases electrons to the external circuit.
The electrons arrive at the carbon electrode

DEPOLARISER: Manganese dioxide MnO_2 .

FUNCTION OF CARBON GRAPHITE GRANULES:

- Helping the manganese dioxide to conduct electric current.

ELECTROLYTE: Ammonium chloride jelly.

FUNCTIONS OF THE PITCH SEAL

- ✓ Keeping contents in place
- ✓ Stopping the drying out of the moist mixture inside.

ADVANTAGES OF THE DRY CELL

- ✓ Small, efficient and easy to use in many different situations.

11. RUSTING

A process in which iron reacts with water in the presence of oxygen.

CONDITIONS FOR RUSTING

Study carefully the experimental setups below

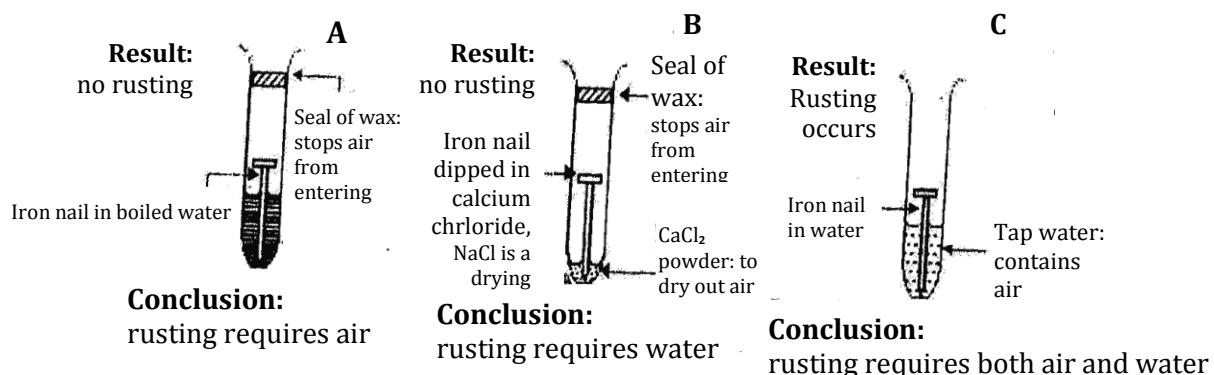


Fig 9.6. Experimental setup of the Conditions for Rusting

- ✓ Presence of water
- ✓ Presence of oxygen

Sea water and acid solutions act as catalysts because they are electrolytes.

***Where there is cold water, oxygen is also present because the oxygen is dissolved in the water.

PROTECTION OR PREVENTION AGAINST RUSTING

1. Painting – Paint is hydrophobic because it is an alkane. Therefore it prevents water and oxygen from being in contact with iron.
2. Covering the metal by oil or grease.
3. Electroplating.
4. Galvanising.

GALVANISING

It is covering iron metal with zinc.

Consider covering iron with tin: In case of a puncture rusting takes place quickly.

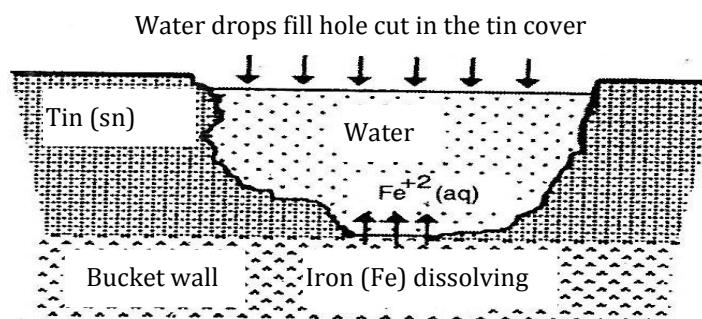
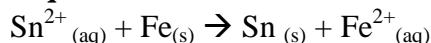


Fig 9.7. A puncture in a tin made of iron covered with tin

EXPLANATION

- The two metals are in contact and the electrolyte water forms a connection. This sets up an electrochemical cell.
- The iron cathode dissolves according to the half equation $\text{Fe}_{(s)} - 2e^- \rightarrow \text{Fe}^{2+}_{(aq)}$
- At the tin anode: tin ions gain the electrons and form tin metal atoms according to the half equation $\text{Sn}^{2+}_{(aq)} + 2e^- \rightarrow \text{Sn}^0_{(s)}$

Overall equation



Iron here is used up to form tin.

Consider covering iron with zinc: In case of a puncture rusting does not take place.

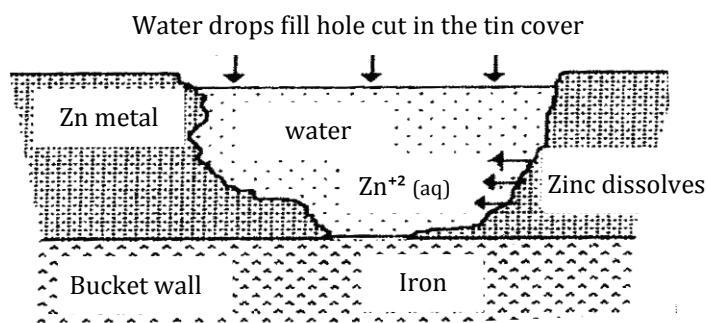
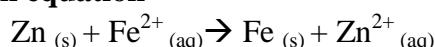


Fig 9.8. A puncture in a tin made of zinc covered iron

EXPLANATION

- The two metals are in contact and the electrolyte water forms a connection. This sets up an electrochemical cell.
- The zinc cathode dissolves according to the half equation $\text{Zn}_{(s)} - 2e^- \rightarrow \text{Zn}^{2+}_{(aq)}$
- At the iron anode: iron ions gain the electrons and form tin metal atoms according to the half equation $\text{Fe}^{2+}_{(aq)} + 2e^- \rightarrow \text{Fe}^0_{(s)}$

Overall equation



Iron ions once formed would be forming iron metal atoms. In fact iron ions do not form at all.

12. SACRIFICIAL PROTECTION

When an iron pipe has been put under soil or water, the pipe corrodes. It is expensive to be removing and changing pipes.

In sacrificial protection an iron pipe is connected with a metal higher up in the electrochemical series e.g. Mg or Al.

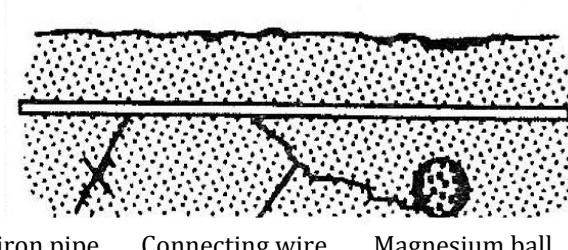


Fig 9.9. An illustrating of sacrificial protection

A magnesium ball is connected to an iron pipe. The magnesium ball dissolves and loses two electrons according to the half equation $Mg_{(s)} - 2e^- \rightarrow Mg^{2+}_{(aq)}$. The electrons go to the iron pipe through the wire. The iron pipe dissolves and the ions produced are forced to accept the two electrons forming iron metal according to the half equation $Fe^{2+}_{(aq)} + 2e^- \rightarrow Fe_{(s)}$

In this way the iron pipe is prevented from rusting. The large ball of Mg can protect the iron pipe many Km long hence it is sacrificial.

Chapter ten

PROTON TRANSFER

O. INTRODUCTION

A proton is a hydrogen ion H^+ . Kumbukirani kuti (remember) a hydrogen atom ionises by losing an electron to remain with a proton only.

1. LOWRY BRONSTED THEORY OF ACIDS AND BASES.

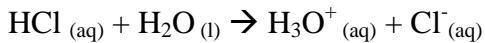
- ✓ An acid is a proton donor. It can be a molecule or an ion.
- ✓ A base is a molecule or an ion that accepts a proton.
- ✓ Consider hydrogen chloride gas reacting with ammonia to form ammonium chloride. $HCl_{(g)} + NH_3_{(g)} \rightarrow NH_4Cl_{(s)}$
 - a. Let us look at what happens to the $HCl_{(g)}$
 $HCl_{(g)} - H^+_{(aq)} \rightarrow Cl^-_{(aq)}$
HCl is donating a proton. According to the Lowry Bronsted theory it is acting as an acid.
 - b. Looking at what happens to the NH_3
 $NH_3_{(g)} + H^+_{(aq)} \rightarrow NH_4^+_{(aq)}$
Ammonia here is accepting a proton and is therefore acting as a Lowry Bronsted base.
- ✓ Consider heating ammonium chloride in a closed container
 $NH_4Cl_{(s)} + Heat \rightarrow NH_3_{(g)} + HCl_{(g)}$. We may rewrite it as
 $NH_4^+Cl^- + Heat \rightarrow NH_3_{(g)} + HCl_{(g)}$
 NH_4^+ donates a proton and is therefore an acid according to the Lowry Bronsted Theory. Cl^- is a proton acceptor. Therefore it is acting as a Lowry Bronsted theory base.

2. HYDROLYSIS

Meaning: reaction with water.

A. FORMATION OF THE HYDRONIUM ION

Ionic equation of ionisation of hydrogen chloride gas in water

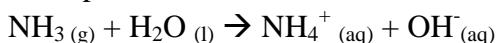


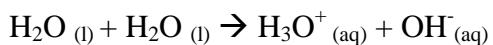
In this reaction a water molecule accepts a proton to form a hydronium ion (H_3O^+). Therefore the water molecule is a Lowry Bronsted base.

B. FORMATION OF A HYDROXIDE ION

With bases water can donate protons and so act as a Lowry Bronsted acid.

Examples:



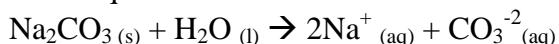


Water can act both as Lowry Bronsted base and Lowry Bronsted acid.

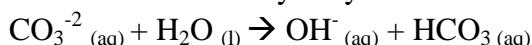
Water can donate protons to itself. Therefore water is an amphoteric/amphiprotic substance.

C. WHY IS SODIUM CARBONATE WHICH CONTAINS NO HYDROXIDE IONS A BASE?

Ionic Equation of ionisation of sodium carbonate



The carbonate ions hydrolyse as follows



Production of hydroxide ions makes the solution basic.

D. WHY IS POTASSIUM HYDROGEN SULPHATE WHICH CONTAINS NO HYDRONIUM IONS AN ACID?

Potassium hydrogen sulphate ionises in water to produce hydrogen sulphate ions.

The hydrogen sulphate ions go on to hydrolyse and produce hydronium ions.

Exercise

Write down the molecular formula of potassium hydrogen sulphate

Show how it ionises in water

Show how the hydrogen sulphate ions go on hydrolysing to produce hydronium ions.

3. STRONG ACIDS AND WEAK ACIDS

According to the ionic theory,

Strong acids ionise almost completely in water to produce many hydrogen ions in water. Examples: hydrochloric acid, sulphuric acid and nitric acid.

Weak acids ionise partially in water and produce few hydrogen ions. Examples: organic acids like ethanoic acid, carbonic acid.

According to Lowry Bronsted theory

A strong acid gives away protons very easily to form many hydrogen ions.

The ionic theory and Lowry Bronsted theory agree on the description of strong acids.

4. STRONG BASES AND WEAK BASES

According to the ionic theory

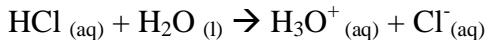
Strong bases almost ionise completely to produce many hydroxide ions in water. E.g. $\text{NaOH}_{(aq)}$, $\text{KOH}_{(aq)}$. Weak bases ionise partially to produce few hydroxide ions in water.

According to the Lowry Bronsted theory strong bases accept protons well and hold onto them very strongly.

5. CONJUGATE ACIDS AND CONJUGATE BASES

A conjugate acid is a base that has accepted a proton in a form of a hydrogen ion.

A conjugate base is an acid that has donated a proton in a form of a hydrogen ion.



$\text{H}_3\text{O}_{(\text{aq})}^+$ is a conjugate acid. We find it by adding a H^+ to its conjugate base. $\text{Cl}_{(\text{aq})}^-$ is a conjugate base. We find it by subtracting a H^+ ion from its conjugate acid.

6. STRENGTH OF AN ACID.

A strong acid has many hydrogen ions, donates protons easily and ionises almost completely in water.

Ways of measuring the strength of an acid

1. USING UNIVERSAL INDICATOR – Add two drops of the universal indicator and observe the colour change. Red means a strong acid, orange means an acid of intermediate strength, yellow means a weak acid.
2. pH METER - Dip the electrodes of the pH meter in the given acid. Strong acids show a pH of 1 and 2. pH 3 and 4 means an acid of intermediate strength. pH 5 and 6 means a weak acid.
3. CONDUCTIVITY CIRCUIT. Place the acid in the beaker of a conductivity circuit. Observe ammeter reading. (Rather vaguely one may observe brightness of a bulb). High ammeter reading means strong acid and low ammeter reading means weak acid.

Factors to keep constant when comparing the strength of two acids

- ✓ The two acids must have same concentration
- ✓ Use the same base for both acids

7. CONCENTRATION

This is the mass of an acid dissolved per litre.

$$\text{Concentration} = \frac{\text{Mass}}{\text{Volume}} = \text{Molarity} \times \text{RFM}.$$

Concentration is experimentally found out by titrating.

8. PROPERTIES OF ACIDS

1. Have sour taste
2. Change the colour of indicators – give examples.
3. React with carbonates to give off water, salt and CO_2 .
4. React with metals higher than the hydrogen half equation in the electrochemical series to produce hydrogen gas.
5. They are electrolytes because they ionise in water to produce hydrogen ions.

NB: Non metal oxides and hydroxides are acids.

9. PROPERTIES OF BASES

1. Have a bitter taste
2. Have a pH higher than 7.
3. React with acids to produce salt and water. Name this reaction.....
4. Have a slippery feel like soap.
5. They are electrolytes because they ionise in water to produce hydroxide ions.

NB: Nitrogen compounds, metal oxides and hydroxides are bases.

Chapter eleven

RATES OF REACTION

O. INTRODUCTION

A rate of anything is always to do with time. Therefore the rate of reaction is also known as SPEED OF REACTION.

1. CASE STUDY EXPERIMENTS

- ✓ The reaction between sodium theosulphate and hydrochloric acid to produce sodium chloride, water, sulphur dioxide and sulphur.



- ✓ The reaction between calcium carbonate and hydrochloric acid to produce calcium chloride, carbon dioxide and water.



2. WAYS OF MEASURING THE RATE OF A REACTION

- i. Rate of production of a gas.
- ii. Observing change in intensity of colour.
- iii. Formation or disappearance of a precipitate.
- iv. Measuring the changing mass of the reaction.
- v. Measuring the increase in volume of a gas.

3. HOW TO INCREASE THE RATE OF A CHEMICAL REACTION

- i. Increasing concentration.
- ii. Increasing the temperature.
- iii. Adding a catalyst.
- iv. Increasing the surface area.
- v. Increasing the pressure of a gas.

NB: other factors affecting reaction rates are light and nature reactants.

4. THE COLLISION THEORY AND THE FACTORS AFFECTING THE RATE OF CHEMICAL REACTIONS.

Collision between reactant particles causes chemical bonds to be broken and allow new ones to form. For this to happen, particles must collide with ACTIVATION ENERGY. Therefore reaction rate speeds up with an increase in the number of successful collisions per second.

I. CONCENTRATION

A more concentrated solution has more particles of the dissolved substance in a given volume of solution. Different concentrations of $\text{HCl}_{(\text{aq})}$ show differences in the steepness of the slope. As the concentration of HCl is increased the steepness of the slope is also increased. This means that increasing the concentration of a liquid (gas/solid) increases the rate of reaction.

Explanation: By increasing the concentration of one or both reactants more particles meet and collide. This produces more collisions per second and the reaction speeds up

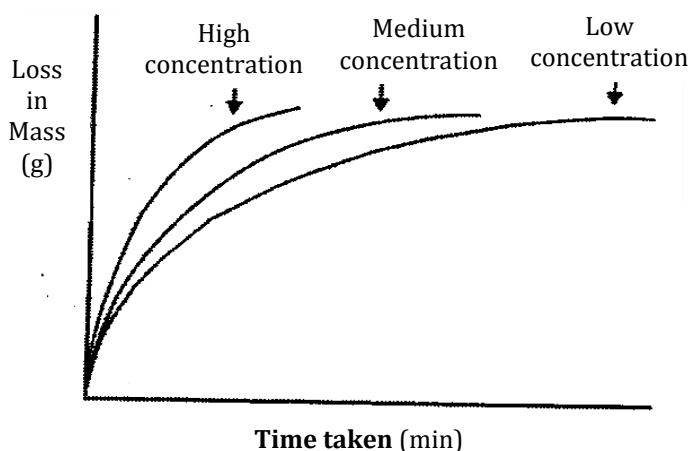


Fig 11.1. Rate reaction and Concentration

(graph of loss of mass against time taken)

II. PRESSURE OF A GAS

For gases the concentration becomes the pressure. High pressure causes more particles to meet and collide. This produces more collisions per second and the reaction speeds up.

III. SURFACE AREA OF A SOLID – PARTICLE SIZE EFFECT.

The surface area of a solid can be increased by decreasing its particle size. Large lumps of a solid have a smaller surface area than the solid when powdered. A large surface area leads to more collisions per second between particles. The rate of reaction increases as the surface area of a solid increases.

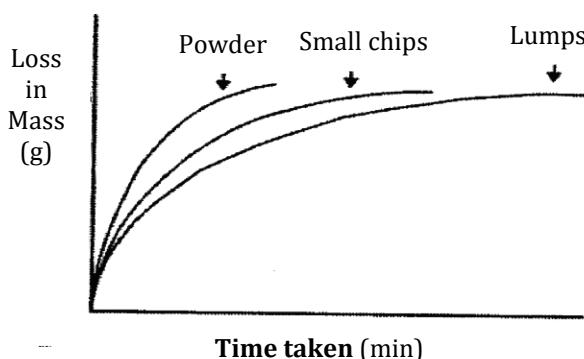


Fig 11.2. Rate reaction and particle size

(graph of loss of mass against time taken)

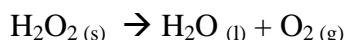
IV. TEMPERATURE

Temperature increases the kinetic energy of the particles hence the particles move faster and faster and there are more collisions per second. This increases the reaction rate.

V. CATALYSTS

A Catalyst is a substance that speeds up a chemical reaction without itself being used up or changed.

- ✓ Hydrogen peroxide (H_2O_2) decomposes slowly to give water and oxygen gas if left in a dark bottle at room temperature. The reaction becomes faster in the presence of catalyst manganese dioxide or potassium iodide.



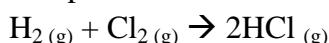
Explanation: When a catalyst is added to a solution the activation energy is lowered meaning that more collisions will have the required energy and so the number of successful collisions will increase the chemical reaction

NOTE: ACTIVATION ENERGY refers to the energy enough to speed up a chemical reaction.

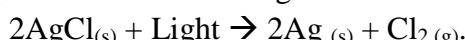
VI. LIGHT

Light affects the speed of some chemical reactions.

- ✓ At room temperature hydrogen and chlorine react slowly to hydrogen chloride. In the presence of ultraviolet rays from the sun the reaction takes place rapidly at room temperature.



- ✓ Solid silver chloride breaks down in the presence of light to form black solid silver particles and chlorine gas.



VII. NATURE OF THE REACTANTS

Reactants that have weak bonds will break more easily and so will react at low activation energy and lower temperatures. Alkenes for example have a weak bond and so react quickly in addition reactions. Organic reactions are generally slow because a lot of bonds need to be broken before a reaction can take place.

Chapter twelve

REVERSIBLE REACTIONS AND REACTIONS AT EQUILIBRIUM

0. INTRODUCTION

Reversible reactions are those chemical reactions which go in both directions.

SYMBOL: Double arrow

Examples:

- a. Heating ammonium chloride in a closed container.
 $\text{NH}_4\text{Cl} + \text{heat} \rightleftharpoons \text{NH}_3 + \text{HCl}$
- b. Haber process – Industrial manufacture of ammonia.
- c. Heating dinitrogen tetroxide (N_2O_4)
- d. Cooling nitrogen dioxide (NO_2)

1. CLOSED AND OPEN SYSTEMS

If ammonium chloride solid powder is heated in a crucible, all the ammonium chloride decomposes to form ammonia gas and hydrogen chloride gas. Nothing remains in the crucible. This is an open system.

If ammonium chloride is heated in a closed container the NH_4Cl decomposes into NH_3 and HCl gas. The ammonia gas and HCl gas once formed combine to form ammonium chloride powder. There is forward and backward reaction.

2. EQUILIBRIUM

Equilibrium is a point in a chemical reaction where the forward and backward reactions occur at the same rate. Since the process continues it is said to be DYNAMIC EQUILIBRIUM.

SYMBOL: \rightleftharpoons

CHARACTERISTIC OF A REACTION AT EQUILIBRIUM

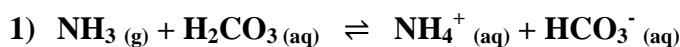
- ✓ Constancy of macroscopic properties. MACROSCOPIC PROPERTIES are those properties that can ordinarily be observed e.g. colour, pressure and volume.

3. LE CHATELIER'S PRINCIPLE

The Le Chatelier's Principle states that

- ✓ Whatever is done to a reaction at equilibrium, the reactions will tend to do the opposite.
- ✓ If a change is applied to system in equilibrium the equilibrium will tend to move in a direction to oppose that change.

Examples of chemical reactions, changes applied and what the reaction does



Ammonia and carbonic acid reacting to produce ammonium ions and hydrogen carbonate ions.

THE CONSTANT MACROSCOPIC PROPERTY: Smell of ammonia

CHANGE APPLIED: Adding more carbonic acid

EFFECT: Equilibrium shifts to the right hand side to produce more ammonium ions and hydrogen carbonate ions. Ammonia gas can no longer be smelt. It is used up.



Potassium chromate reacts with sulphuric acid to produce potassium dichromate, potassium sulphate and water.

Change applied	Colour change	Change in position of equilibrium
Increase concentration of K_2CrO_4	Increase in intensity of orange colour	Equilibrium shifts to the right hand side to produce more $\text{K}_2\text{Cr}_2\text{O}_7$
Increase concentration of $\text{K}_2\text{Cr}_2\text{O}_7$	Increase in intensity of yellow colour	Equilibrium shifts to the left hand side to produce more K_2CrO_4
Decreasing concentration of K_2CrO_4	Decrease in intensity of orange colour	Equilibrium shifts to the left hand side to produce more K_2CrO_4

Table 12.1. Effects of changes applied to a reaction at equilibrium

Chapter thirteen

EXOTHERMIC AND ENDOOTHERMIC REACTIONS (HEATS OF REACTION)

0. INTRODUCTION

In chemical reactions there is involvement of energy during bond breaking and bond making. When methane and oxygen gas react to form carbon dioxide and water, there is energy required for breaking the bonds in the reactants. As the products form there is liberation of energy.

1. EXOTHERMIC REACTIONS

DEFINITION: It is a type of chemical reaction in which heat energy is given out to the surrounding. The container feels warm – it heats up. Temperature rises.

EXAMPLES

1. Mixing sulphuric acid with water
2. Dissolving NaOH pellets in water
3. Reaction between a strong acid and a strong base.
4. Respiration
5. Reaction between magnesium metal ribbon and $\text{HCl}_{(\text{aq})}$
6. Reaction between magnesium and oxygen gas

HEAT ENERGY DIAGRAM TO SHOW EXOTHERMIC REACTIONS

- The picture model is that of a brick falling from the top of the house naturally. The loss of potential energy is likened to the loss of heat to the surrounding.
- Heat of reaction is the value of heat change. Heat Change is also enthalpy change.

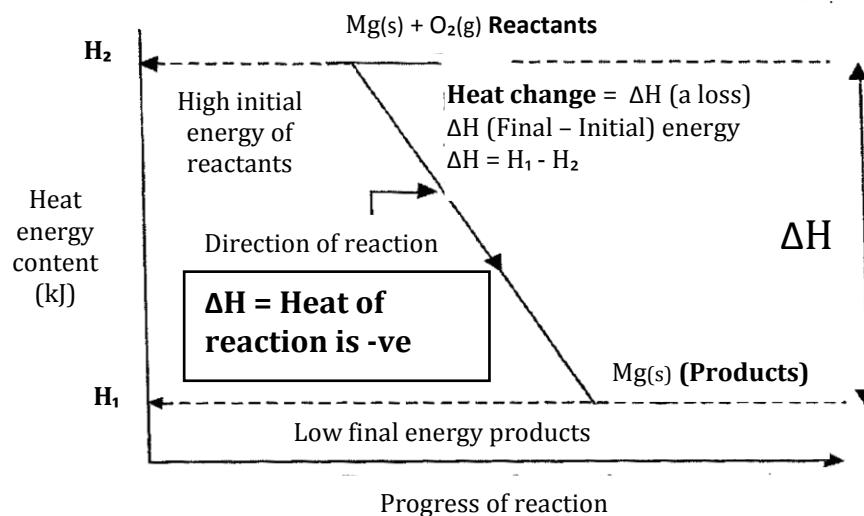


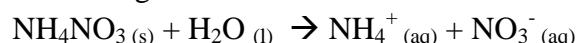
Fig 13.1. Heat energy diagram for exothermic reactions

2. ENDOOTHERMIC REACTIONS

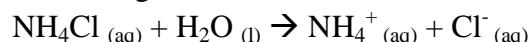
DEFINITION: It is a type of chemical reaction in which heat energy is taken in from the surrounding.

EXAMPLES

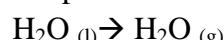
1. Dissolving ammonium nitrate in water



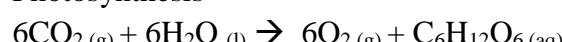
2. Dissolving ammonium chloride in water



3. Evaporation



4. Photosynthesis



5. Dissolving potassium nitrate in water

EXERCISE:

Write down the ionisation of potassium nitrate in water.

HEAT ENERGY DIAGRAM TO SHOW ENDOOTHERMIC REACTIONS

The picture model is that of a book being raised onto a shelf.

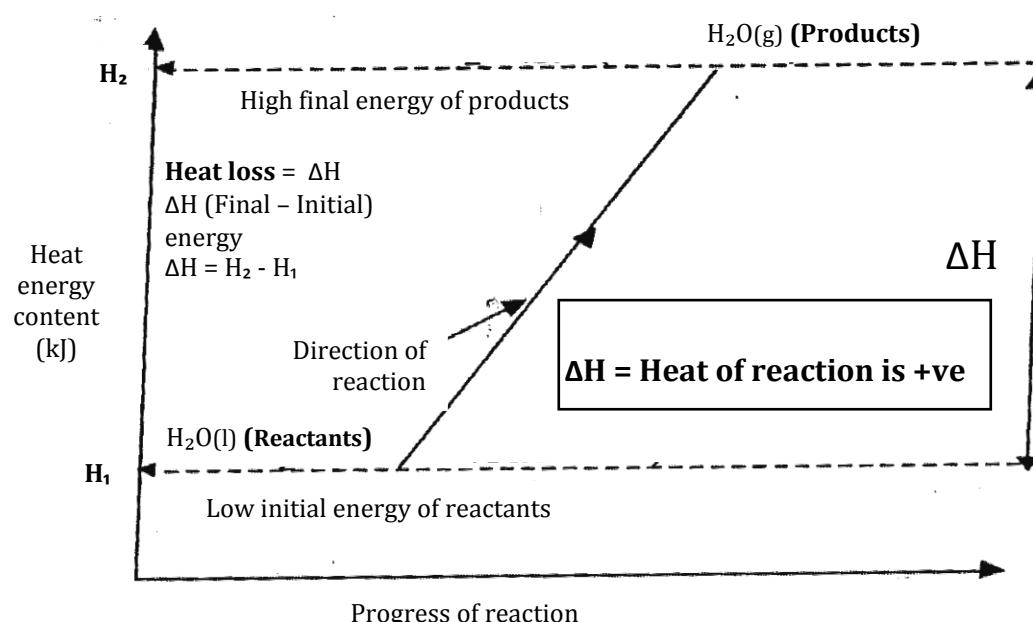


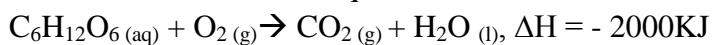
Fig 13.2 diagram heat energy for endothermic reactions

3. WHY IS BOND MAKING EXOTHERMIC WHILE BOND BREAKING ENDOOTHERMIC?

Bond making releases heat to the surrounding while bond breaking requires heat from the surrounding.

4. WRITING DOWN CHEMICAL EQUATIONS FOR EXOTHERMIC REACTIONS

- i. Writing down the equation in full and the heat of reaction is written separately using a comma at the end of the equation.

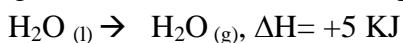


- ii. Writing down the heat of reaction as part of the products.

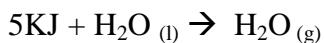


5. WRITING DOWN CHEMICAL EQUATIONS FOR ENDOTHERMIC REACTIONS

- a. Writing down the equation in full and the heat of reaction is written separately using a comma at the end of the equation.



- b. Writing down the heat of reaction as part of the products.



6. HEAT AND TEMPERATURE

Temperature and heat are related but they are not the same. Supplied with the same amount of heat, water and a needle register differences in temperature rise.

Temperature is measured in Kelvins while heat is measured in joules.

Heat is a form of energy while temperature is the degree of hotness or coldness.

Chapter fourteen

ORGANIC CHEMISTRY

O. INTRODUCTION

DEFINITION: *Organic chemistry is a branch of science that deals with organic (carbon) compounds.*

CATEGORISATION OF SUBSTANCES

Substances are categorised either as PURE SUBSTANCES or MIXTURES (Impure substances). The other way is to categorise substances either as ELEMENTS or COMPOUNDS.

CATEGORISATION OF COMPOUNDS

- I. Inorganic Compounds – These do not contain a carbon atom in their chemical formula.
- II. Organic Compounds (Carbon Compounds) – These contain the element carbon in their chemical formula. Exceptions are CO_2 , CO and metal carbonates e.g. calcium carbonate and lithium carbonate which are simple compounds of carbon.

CATEGORISATION OF ORGANIC COMPOUNDS

- I. Hydrocarbons – These are organic compounds which contain the elements carbon and hydrogen only
- II. Oxycarbons – These are organic compounds containing the element oxygen.
- III. AMINES – These are organic compounds with the $-\text{NH}_2$ - functional group.

Why are there more organic compounds than all the inorganic compounds put together?

Carbon has a valency of 4. It can make 4 bonds. Carbon atoms join up with other carbon atoms to form long chains of carbon atoms inside molecules of different compounds.

FUNCTION OF ORGANIC CHEMISTRY

Collecting together families of carbon compounds in homologous series. It is enough to study the properties of one or two members to know the general properties of all homolog members. This reduces the tedious task of studying the properties of each organic compound separately.

HOMOLOGOUS SERIES: It is a collection of organic compounds that have similar chemical properties with related molecular formulae and structures. Members of a homologous series are homologs.

GENERAL PROPERTIES OF HOMOLOGS

- I. All homolog members conform to a general molecular formula.
- II. All homologous series have a common system of naming (NOMENCLATURE). The prefix indicates the number of carbon atoms in that compound.
- III. A homologous series has a functional group. A functional group is a special atom or group of atoms joined together. The functional group determines chemical properties of the members of the homologous series.
- IV. All members show similar chemical properties (reactions). In homologous series what is responsible for chemical properties is the FUNCTIONAL GROUP.
- V. Homolog members next to each other by CH_2 .
- VI. General methods of preparation are known and this can be applied to any member.
- VII. The physical properties of the members change gradually in the same direction along the series.

1. ALKANES (PARAFFINS)

Alkanes are a homologous series of saturated hydrocarbons with a general formula $\text{C}_n\text{H}_{2n+2}$.

Why hydrocarbons? They contain hydrogen and carbon atoms only.

Why saturated? Their bonding ability is completely used up. Therefore they cannot add any more atoms.

A. FUNCTIONAL GROUP

Alkanes have no functional group. They only have C–C and C–H single bonding. The single bond is very strong.

B. STUDY THE TABLE BELOW – Copy and Fill in the missing gaps

FORMULA	MO. STRUCTURE& CO. FORMULA	NAME

Table 14.1 The molecular formula, structure, condensed formula and name of alkane

- ✓ Take two condensed formulae of alkane members next to each other.
Observe the increase in CH_2 .

A CONDENSED FORMULA is a chemical formula that also gives how atoms are attached to each other in the structural formula.

C. PHYSICAL PROPERTIES OF ALKANES

Physical properties change gradually when moving from one alkane member to the next.

EXPLANATION: Passing from one alkane molecule to the next there is addition of CH_2 . Molecular sizes and masses increase. The strength of Intermolecular forces increases. Physical properties are determined by intermolecular forces.

I. Melting Points and Boiling Points

These changes of state require heat energy. The heat separates the molecules. The greater the IMF holding the molecules together the greater the heat needed to pull the molecules apart. As the number of carbon atoms increases the molecular size increases too. This leads to increase in the strength of IMF and makes melting and boiling points high.

II. State at room temperature

All substances which have their melting and boiling points below room temperature are gases at room temperature. Liquids have their melting points below room temperature and their boiling points above room temperature. Solids have both melting and boiling points above room temperature. Alkanes exist as solids, liquids and gases at room temperature.

III. Density

Density refers to how closely particles are packed together in matter. As the number of carbon atoms increases density in alkanes increases too. Compared to water alkanes are less dense.

IV. Viscosity

The viscosity of alkanes increases with an increase in the number of carbon atoms. Compared to water alkanes are more viscous.

V. Solubility

The solubility rule is ‘**LIKE DISSOLVES LIKE**’.

Alkanes are hydrocarbons. Hydrocarbons are unlike water and so alkanes are insoluble in water. They are referred to as HYDROPHOBIC meaning ‘Water hating’. When alkanes are mixed with water two layers form.

D. CHEMICAL PROPERTIES OF ALKANES

Alkanes are inert or unreactive because they have no functional group. They are also called PARAFFINS meaning ‘little affinity.’

Alkali metals are stored in liquid alkanes e.g. paraffin and ligroin to protect them from oxidation by air.

The strong bonds in alkanes need high temperatures to break when taking part in chemical reactions. Alkanes react best with substances which are themselves very reactive like oxygen gas, O₂ and chlorine gas; Cl₂. Alkanes undergo combustion and substitution reactions.

I. Combustion/Burning

It is a reaction between a fuel and oxygen gas to produce heat energy in addition to other products. It is a strong exothermic reaction. A fuel is a substance which burns well. Combustion is complete when oxygen is in excess and incomplete when oxygen is not enough.

COMPLETE	INCOMPLETE
<ul style="list-style-type: none">• Sufficient O₂• Produces blue flame• Flame is very hot• Products are CO and water vapour• Does not contain soot; $2\text{C}_2\text{H}_{6(\text{g})} + 7\text{O}_{2(\text{g})} \rightarrow 4\text{CO}_{2(\text{g})} + 6\text{H}_2\text{O}_{(\text{g})}$	<ul style="list-style-type: none">• Insufficient O₂• Produces yellow flame• The flame is less hot• Products include CO and C• Contains soot; $2\text{C}_2\text{H}_{6(\text{g})} + 6\text{O}_{2(\text{g})} \rightarrow 3\text{CO}_{2(\text{g})} + 6\text{H}_2\text{O}_{(\text{g})} + \text{C}_{(\text{s})}$ $\text{C}_4\text{H}_{10(\text{g})} + 3\text{O}_{2(\text{g})} \rightarrow \text{CO}_{2(\text{g})} + 5\text{H}_2\text{O}_{(\text{g})} + \text{CO}_{(\text{g})} + \text{C}_{(\text{g})}$

Table 14.2. Characteristics of complete and incomplete combustion.

PROBLEM

Write a balanced equation for the combustion of butane and complete the table.

NOTES ON COMBUSTION

A flame is necessary to begin the combustion of alkanes. The high temperatures must provide activation energy for breaking chemical bonds.

Once started the reaction continues without any help. It is a strong exothermic reaction and so the making of new bonds gives out large amounts of energy.

II. SUBSTITUTION REACTIONS

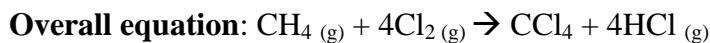
Chlorine (halogens) reacts with alkanes in the presence of ultraviolet light from the sun in substitution reactions. The ultraviolet light provides energy to break the Cl-Cl bond in Cl₂.

GENERAL EQUATION:



STAGES

- i. $\text{CH}_4 \text{(g)} + \text{Cl}_2 \text{(g)} \rightarrow \text{CH}_3\text{Cl} + \text{HCl(g)}$
Methane reacts with chlorine gas to produce monochloromethane and hydrogen chloride gas.
- ii. $\text{CH}_3\text{Cl} + \text{Cl}_2 \text{(g)} \rightarrow \text{CH}_2\text{Cl}_2 + \text{HCl(g)}$
Monochloromethane further reacts with chlorine gas to produce dichloromethane and hydrogen chloride gas
- iii. $\text{CH}_2\text{Cl}_2 + \text{Cl}_2 \text{(g)} \rightarrow \text{CHCl}_3 + \text{HCl(g)}$
Dichloromethane further reacts with chlorine to produce trichloromethane (chloroform) and hydrogen chloride gas.
- iv. $\text{CHCl}_3 + \text{Cl}_2 \text{(g)} \rightarrow \text{CCl}_4 + \text{HCl(g)}$
Trichloromethane reacts with chlorine to produce tetrachloromethane and hydrogen chloride gas



Trichloromethane is commonly called chloroform

Tetrachloromethane is commonly called carbon tetrachloride.

USES OF HALOALKANES (HALOGEN ALKANES)

- i. Used as fire extinguishers because they do not burn in air
- ii. Used as anaesthetics e.g. chloroform
- iii. Used as solvents e.g. carbon tetrachloride

2. ALKENES

Alkenes are a homologous series of unsaturated hydrocarbons with a general formula C_nH_{2n} .

A. FUNCTIONAL GROUP: Carbon – Carbon double bond. C=C

There is no member of alkene family with one carbon atom because there must be two carbon atoms present for there to be a double bond.

Why unsaturated?

The bonding ability of alkenes is not completely used up. In chemical reactions the double covalent bond easily breaks and more bonds can form and other atoms can be added on.

B. STUDY THE TABLE BELOW – Fill in the missing gaps

FORMULA	MO. STRUCTURE& CO. FORMULA	NAME

Table 14.3. Molecular formula, structure, condensed formula and name of alkene

Take two condensed formulae of alkane members next to each other. Observe the increase in CH_2 .

C. PHYSICAL PROPERTIES – Refer to notes on physical properties of alkanes

Solubility: Hydrocarbons are unlike water. Therefore alkenes are insoluble in water.

State the observation made when an alkene is mixed with water

D. CHEMICAL PROPERTIES OF ALKENES

I. Combustion – Refer to notes on combustion of alkanes

- ✓ Give an equation of complete combustion of propene.
- ✓ In terms of products differentiate between complete combustion and incomplete combustion of alkenes.
- ❖ Compared to alkanes with similar number of carbon atoms, when alkenes burn, they do so with a sootier flame because alkenes have a higher percentage of carbon

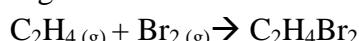
II. Additional reactions

In additional reactions the double bond breaks and the alkene can add extra atoms.

- Alkenes reacting with halogens

GENERAL EQUATION: Alkene + halogen \rightarrow haloalkane

E.g. Ethene reacts with bromine to yield bromoethane



- Reacting Alkenes with steam yields an alkanol.



- c. HYDROGENATION: Alkenes react with hydrogen to produce saturated hydrocarbons.
- $$\text{C}_2\text{H}_4 \text{ (g)} + \text{H}_2 \text{ (g)} \xrightarrow{140^\circ\text{C} \& \text{nickel}} \text{C}_2\text{H}_6 \text{ (g)}$$

INDUSTRIAL USE OF HYDROGENATION

- ✓ Converting unsaturated vegetable oils to saturated fats in the manufacture of margarine.

- d. Polymerisation (*refer to notes on polymerisation*)

E. BROMINE TEST: TEST FOR CARBON-CARBON DOUBLE BOND

- Put 15 drops of bromine solution (Br_2CCl_4) in a test tube. Add two drops of the unknown (test liquid).
- POSITIVE RESULT: The test liquid decolourises bromine i.e. Bromine solution changes from reddish brown to colourless.
- CONCLUSION: The test liquid has a $\text{C}=\text{C}$ or is an alkene or is an unsaturated hydrocarbon

F. SOURCES OF ALKENES

- ✓ Cracking of alkanes: Cracking is the splitting of large hydrocarbons into more useful smaller hydrocarbons at high temperatures involving a catalyst. E.g. $\text{C}_{12}\text{H}_{26} \xrightarrow{\text{silica \& aluminium oxide}} \text{C}_5\text{H}_{12} + \text{C}_4\text{H}_8 + \text{C}_3\text{H}_6$
- ✓ Dehydration of alcohols
Ethene can be obtained by removing a water molecule from ethanol.
 $\text{C}_2\text{H}_5\text{OH} \text{ (l)} \rightarrow \text{C}_2\text{H}_4 \text{ (g)} + \text{H}_2\text{O} \text{ (l)}$

3. ALKANOLS

Alkanols are a homologous series of Oxycarbons with a general formula $\text{C}_n\text{H}_{2n+1}\text{OH}$.

A. FUNCTIONAL GROUP: OH

OH functional group called the hydroxyl group.

B. STUDY THE TABLE BELOW – Fill in the missing gaps

FORMULA	MO. STRUCTURE& CO. FORMULA	NAME

Table 14.4. Molecular formula, structure, condensed formula and name of alkanol

Take two condensed formulae of alkane members next to each other. Observe the increase in CH₂.

C. PHYSICAL PROPERTIES OF ALKANOLS

POINTS TO NOTE:

- As we go along a homologous series there is an increase in molecular sizes because the number of carbon atoms increase. This makes IMF become stronger.
- Alkanols have stronger IMF than hydrocarbons because of the presence of the special hydrogen bonding, OH.
- Along the alkanol series the relative size of the OH group gets smaller as the molecules get bigger

I. Melting points and boiling points

Going along the alkanol homologous series, M. Pts and B. Pts increase –Explain why?

Comparing alkanols with hydrocarbons of similar molecular masses alkanols have higher M. Pts and B. Pts – Suggest why?

The presence of the special hydrogen bonding OH makes alkanols to have stronger intermolecular forces than hydrocarbons with similar molecular masses.

Work out the molecular masses of ethanol and propane?

Why is ethanol a liquid while propane is a gas at room temperature?

II. Viscosity

Glycerol (CH₂OHCHOHCH₂OH) has more than one OH group making it viscous.

III. Solubility

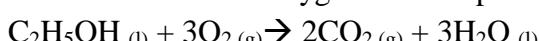
Alkanols are soluble in water. The presence of OH makes alkanols to be like water.

Going along the alkanol homologous series solubility decreases because the OH proportion decreases in an alkanol.

D. CHEMICAL PROPERTIES OF ALKANOLS

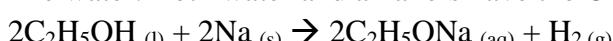
I. Combustion

Alkanols react with oxygen in air to produce carbon dioxide and water.



II. Reaction with sodium metal

Alkanols react with sodium to produce hydrogen gas and an alkaline solution just like water. Both water and alkanols have the OH group.



Ethanol reacts with sodium to produce sodium ethoxide and hydrogen gas. This is a replacement reaction.

III. Oxidation

Oxidation of alkanols produces alkanoic acids. Alkanals or Aldehydes are the intermediate products.

✓ OXIDATION: Alkanals are easily oxidised while ketones are not easily oxidised. Oxidation is a chemical property of both alkanols and alkanals.

Alkanols are partially oxidised to alkanals and completely oxidised to alkanoic acids.

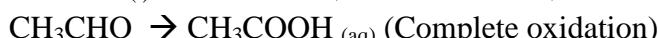
GENERAL EQUATION



OXIDISING AGENTS:

- Acidified potassium dichromate ($\text{K}_2\text{Cr}_2\text{O}_7$)
- Air in the presence of a catalyst

Example



IV. Esterification

(Notes will be found on page 85)

V. Dehydration

Is the removal of a water molecule in ethanol molecules to produce alkenes. Ether is an intermediate product.

GENERAL EQUATION



Dehydrating agent: $\text{H}_2\text{SO}_4_{(aq)}$

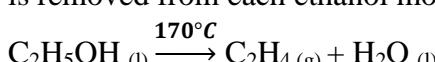
➤ PARTIAL DEHYDRATION

Partial dehydration occurs at a temperature of 140°C . One water molecule is removed from two molecules of ethanol to produce ether.



➤ COMPLETE DEHYDRATION

Complete dehydration occurs at a temperature above 170°C . One water molecule is removed from each ethanol molecule to form ethene.



E. TEST FOR ALKANOLS

I. SOLUBILITY TEST

- Add 15 drops of water to a test tube.
- Add at least 10 drops of the test liquid.

POSITIVE RESULT: One layer forms

INTERPRETATION: The sample liquid could be an alkanol.

II. SODIUM TEST

- Add 15 drops of test liquid in a test tube.

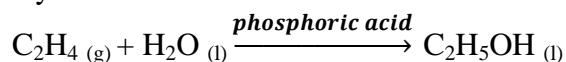
- Place a match head amount of freshly cut sodium metal

POSITIVE RESULT: Sodium metal disappears. Bubbles of a gas form. The gas produces a pop sound when ignited. The liquid formed turns phenolphthalein solution to pink colour.

INTERPRETATION: The test liquid is either an alkanol.

F. SOURCES OF ALKANOLS

- ✓ Hydration of alkenes



- ✓ Fermentation. **Fermentation** is a chemical process in which glucose breaks down into ethanol and carbon dioxide gas in the presence of yeast.



IMPORTANT NOTES

- Yeast is a living organism that contains the enzyme zymase.
- Fermentation stops when the ethanol concentration reaches 12% because the yeast is killed.
- Wines and beers are solutions of ethanol and water.
- To produce more concentrated drinks the alcohol is distilled.
- Rectified spirits have 96% alcohol concentration.
- Absolute alcohol contains less than 0.5% water.

The Traditional Method for the Preparing Ethanol (*kachasu*)

Maize bran (madeya), water and sugar are mixed in a clay pot (mtsuko). Yeast is added to the mixture to encourage fermentation to occur. The clay pot top is covered with a black plastic bag. The mixture is kept for three days so that fermentation completes. The mixture is then distilled using local distillation apparatus as shown below:

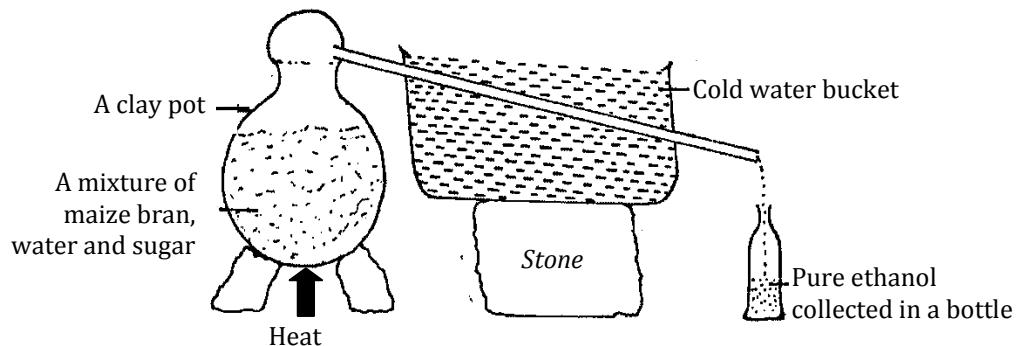


Fig 14.2. A local distillation apparatus

The mixture in the clay pot is heated using firewood. The ethanol vaporises upon reaching the temperature of 78°C. The ethanol vapour moves up the clay pot into the delivery tube. As the ethanol vapour passes through the condenser the cold water in the condenser cools it to form liquid ethanol which is collected in the receiving container.

G. INDUSTRIAL USE OF ETHANOL

1. Motor fuel blending agent.
2. Solvent.
3. Manufacturing of cosmetics.
4. Manufacturing ethyl ethanoate.

H. DANGERS OF ALCOHOL CONSUMPTION

- i. Damaging the liver causing cirrhosis.
- ii. Mood changes leading violent behaviour.
- iii. Slowing down reactions leading to road accidents.

4. CARBOXYLIC/ALKANOIC ACIDS

They are a homologous series of oxycarbons with a general molecular formula $C_nH_{2n+1}COOH$.

A. FUNCTIONAL GROUP

The functional group is COOH the carboxylate group.

B. STUDY THE TABLE BELOW – Fill in the missing gaps

FORMULA	MO. STRUCTURE& CO. FORMULA	NAME

Table 14.5. Molecular formula, structure, condensed formula and name of alkanoic acids

Take two condensed formulae of alkane members next to each other. Observe the increase in CH_2 .

C. SOURCES OF ALKANOIC ACIDS

I. NATURAL SOURCES

- a. Insect sting e.g. bee sting has methanoic acid (HCOOH)
- b. Citrus fruits have citric/ethanoic/acetic acid.
- c. Tomatoes
- d. Human sweat has butanoic acid.

II. ARTIFICIAL SOURCES

- a. Oxidation of alkanols
- b. Distillation of sodium ethanoate with concentrated acid.
$$\text{CH}_3\text{COONa} + \text{H}_2\text{SO}_4 \text{ (aq)} \rightarrow \text{CH}_3\text{COOH} \text{ (aq)} + \text{NaHSO}_4$$
- c. Hydration of aldehydes

D. PHYSICAL PROPERTIES

NB: Carboxylic acids have stronger IMF than alkanols and hydrocarbons.

EXPLANATION: The presence of the electron withdrawing C=O makes the OH in carboxylic acids more polarised.

I. Solubility

Alkanoic acids are soluble in water because of the presence of the OH.

II. Viscosity

Alkanoic acids have high viscosity

III. Density

Relative to hydrocarbons and alkanols, alkanoic acids have high density which increases on going along the homologous series.

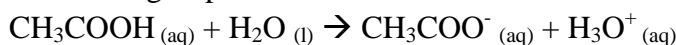
IV. Strong odour/choking smell

V. Acid physical properties: sour taste and pH less than 7.

E. CHEMICAL PROPERTIES

I. Acid reactions

Alkanoic acids are weak acids. They can donate hydrogen ions from the functional group COOH.



Ethanoic acid reacts with water to produce ethanoate ion and hydronium ion.

TYPICAL ACID REACTIONS INCLUDE

- a. Neutralisation

General equation of neutralisation

Acid + Base \rightarrow Salt and Water

Example: $\text{CH}_3\text{COOH} \text{ (aq)} + \text{NaOH} \text{ (aq)} \rightarrow \text{CH}_3\text{COONa} \text{ (s)} + \text{H}_2\text{O} \text{ (l)}$

Ethanoic acid reacts with sodium hydroxide to produce sodium ethanoate and water.

- b. Reaction with magnesium to produce a magnesium salt and hydrogen gas. $2\text{CH}_3\text{COOH} \text{ (aq)} + \text{Mg} \text{ (s)} \rightarrow (\text{CH}_3\text{COO})_2\text{Mg} + \text{H}_2 \text{ (g)}$
- c. Turns blue litmus red or in general changing the colour of indicators.
- d. Reaction with carbonates to produce carbon dioxide gas

II. Esterification

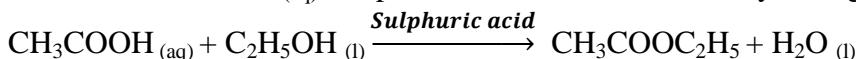
This is a chemical reaction between alkanoic acids and alkanols to produce esters and water.

It is a condensation reaction.

GENERAL EQUATION

Alkanoic acid + Alkanol \rightarrow Ester and water

CATALYST: H_2SO_4 (aq). Sulphuric acid functions as a dehydrating agent.



Ethanoic acid reacts with ethanol to produce ethyl ethanoate and water.

Esters have a pleasant fruit smell

- Octylethanoate has an orange smell.
- Butylethanoate has a banana smell
- Ethylbutanoate has a pineapple smell.

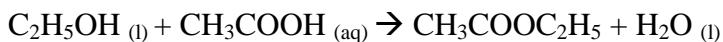
USES OF ESTERS

1. For making perfume.
2. Artificial fruit flavouring
3. Substitute for cotton in the textile industry
4. For making yacht sails
5. Bonding region in glass fibre
6. For making food roasting bags.

F. PREPARATION OF ETHYL ETHANOATE

MATERIALS: Ethanol, ethanoic acid, concentrated sulphuric acid.

CHEMICAL EQUATION



The reaction is at equilibrium

Excess alcohol must be used to shift the reaction to the right.

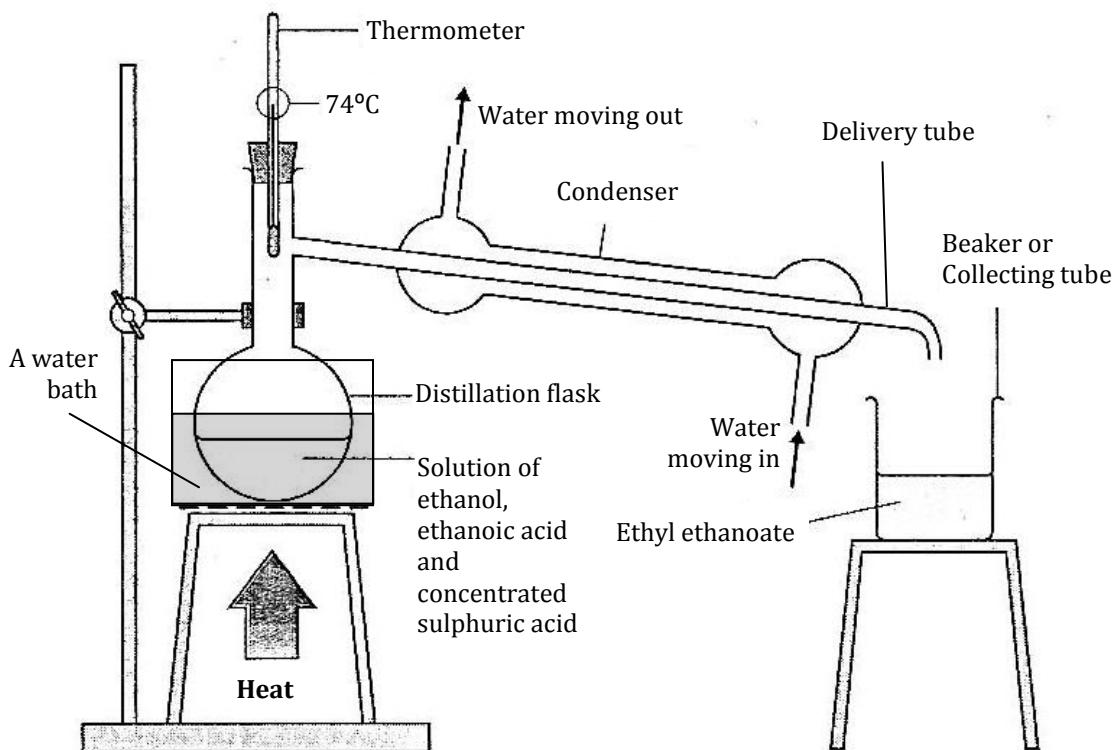


Fig 14.3. Experimental set up for the preparation of ethyl ethanoate.

NOTES

- Since the reaction is at equilibrium, the distillate contains excess ethanol and ethanoic acid.
- To remove excess acid, add sodium hydroxide to neutralise it. This produces sodium ethanoate and water.
 $\text{CH}_3\text{COOH}_{(\text{aq})} + \text{NaOH}_{(\text{l})} \rightarrow \text{CH}_3\text{COONa}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})}$
- To remove excess ethanol, add anhydrous calcium chloride. The products are calcium ethoxide and hydrochloric acid.
 $\text{C}_2\text{H}_5\text{OH}_{(\text{l})} + \text{CaCl}_2_{(\text{s})} \rightarrow \text{C}_2\text{H}_5\text{OCO}_{(\text{aq})} + \text{HCl}_{(\text{aq})}$
- Ethyl ethanoate has the least density. To separate it from impurities, run off the aqueous the lower layer using a separating funnel.

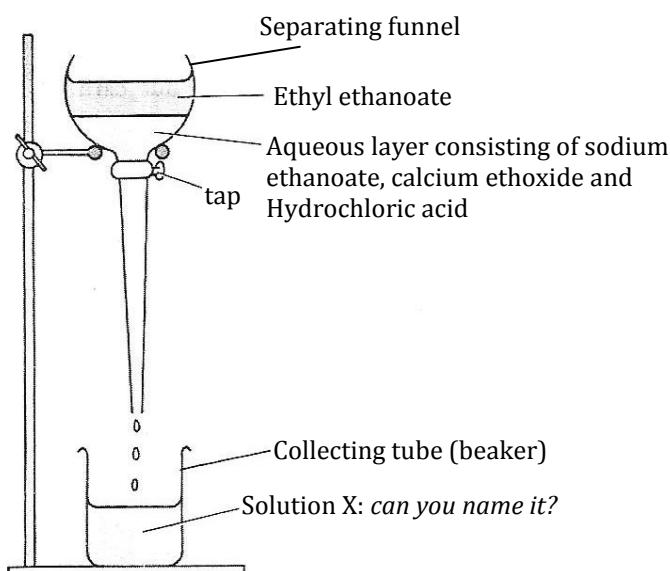


Fig 14.4. Separation of ethyl ethanoate

- Thereafter the ethylethanoate is distilled at 74°C to obtain pure ester.

PHYSICAL PROPERTIES OF ETHYL ETHANOATE

1. Pleasant fruit smell.
2. Slightly soluble in water.
3. Colourless
4. Volatile

CHEMICAL PROPERTIES OF ETHYL ETHANOATE

- i. Burns with bright flames (combustion)
- ii. Hydrolyses slowly to carboxylic acids and alcohols.

USES OF ETHYL ETHANOATE

- a. As a solvent for adhesives e.g. cement
- b. For saponification
- c. Making perfume.

5. ISOMERS AND CONFORMATIONS

ISOMERISM is a situation where two or more molecules with the same molecular formula have different structural formulas.

ISOMERS are molecules with the same molecular formula but different molecular structures.

Here are the three isomers of pentane

Fig 14.5 diagram showing isomers of pentane

Draw the two isomers of butane

CONFORMATIONS are molecules having the same molecular formulas and the same structural formulas but different shapes.

Here are the three conformations of 2 methyl butane

Fig 14.6. Conformations of 2 methyl butane

Draw five conformations of pentane.

We use carbon skeletons in order to see structures of organic compounds easily. Straight chain carbon skeletons have no junctions.

RULES IN NAMING ISOMERS

- I. Write down the carbon skeleton
- II. Choose the parent chain in the carbon skeleton.

IMPORTANCE OF PARENT CHAIN

- Gives the basic name of the compound
- Helps in numbering
- Locating accurately the number and position of the junction in the compound

DEFINITION OF PARENT CHAIN

- For alkanes it is the longest chain of carbon atoms
- For organic compounds with functional group it is the longest chain of carbon atoms that includes the functional group.

HOW TO NUMBER THE PARENT CHAIN

- a. For alkanes C₁ is the carbon atom at the end of the chain nearest to the first junction in the parent chain.
- b. For organic compounds with functional group C₁ is the carbon atom at the end of the carbon chain nearest to the functional group

III. Count up the number of carbon atoms in the parent chain and give a basic name. Show the position of the functional group by the lowest number of the carbon atom holding the functional group.

IV. Count up the number of carbon atoms in each branched group and give its alkyl name

ALKYL RADICAL	ALKYL NAME
---------------	------------

CH ₃	Methyl
-----------------	--------

C ₂ H ₅	_____
-------------------------------	-------

_____	Propyl
-------	--------

GENERAL FORMULA OF ALKYL GROUP is C_nH_{2n+1}

EXERCISE 14a

1. Name the molecules with the following carbon skeletons.
2. Draw the carbon skeleton of the molecules with the following names
 - a. 3, 4, diethyl hex-1-ene
 - b. 2, methyl, 3 ethyl hexan-2-ol

6. CYCLIC/RINGED/CLOSED CHAIN ORGANIC COMPOUNDS

Straight chain organic compounds have carbon atoms connected end to end. (i.e. in lines of C atoms). They are also called ALIPHATIC COMPOUNDS.

Cyclic compounds have carbon atoms joined in rings.

Examples of cyclic compounds

Fig 14.7 diagrams showing cyclic compounds

NOTES

- ✓ Cycloalkanes and open chain alkenes with the same number of carbon atoms are isomers. They have the same molecular formula but different structures.
- ✓ Isomers belonging to different homologous series have the different chemical properties because they have different functional groups.
- ✓ Isomers belonging to the same homologous series have the same chemical properties. Their physical properties are similar but differ slightly because the different structures cause a slight difference in their Intermolecular forces.

7. ALDEYDES/ALKANALS AND KETONES

A. FUNCTIONAL GROUP

It is C=O, carbonyl group functional group.

B. REPRESENTING ALKANALS AND KETONES IN TERMS OF ALKYL RADICAL, R.

ALKANALS: RHC=O

KETONES: RC=OR

Fig 14.8. Representation of alkanals and ketones in terms of alkyl radical, R.

C. STUDY THE TABLE BELOW (FOR ALKANALS)– Fill in the missing gaps

FORMULA	MO. STRUCTURE& CO. FORMULA	NAME

Table 14.6. Molecular formula, structure, condensed formula and name of alkanals

Take two condensed formulae of alkane members next to each other. Observe the increase in CH₂.

STUDY THE TABLE BELOW (FOR KETONES) – Fill in the missing gaps

FORMULA	MO. STRUCTURE& CO. FORMULA	NAME

Table 14.7. Molecular formula, structure, condensed formula and name of ketones

Take two condensed formulae of alkane members next to each other. Observe the increase in CH₂.

NOTES

- ✓ For ketones there is no H atom joined to the C atom holding the C=O.
- ✓ The simplest ketone is propanone.

D. GENERAL MOLECULAR FORMULA

ALDEHYDES: C_nH_{2n+1}CHO

KETONES: C_nH_{2n+1}COR (R is the alkyl radical and is usually CH₃)

E. PHYSICAL PROPERTIES

Both have IMF of intermediate strength. The IMF is stronger than that of hydrocarbons and weaker than that of alkanols and alkanoic acids.

I. Solubility

Both are soluble in water because of the presence of the C=O which is more polarised.

F. CHEMICAL PROPERTIES

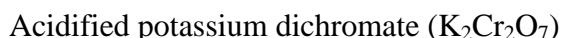
- ✓ COMBUSTION: Both burn in air to form CO₂ and H₂O.
- ✓ OXIDATION: Alkanals are easily oxidised while ketones are not easily oxidised. Oxidation is a chemical property of both alkanols and alkanals.

Alkanols are partially oxidised to alkanals and completely oxidised to alkanoic acids.

GENERAL EQUATION

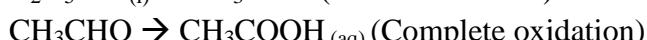


OXIDISING AGENTS:



Air in the presence of a catalyst

Example



- ✓ React with 2, 4 DNPH to produce a red precipitate.

G. HOW TO PREPARE ETHANAL

How to prepare ethanal

MATERIALS: Ethanol, acidified potassium dichromate and distillation apparatus.

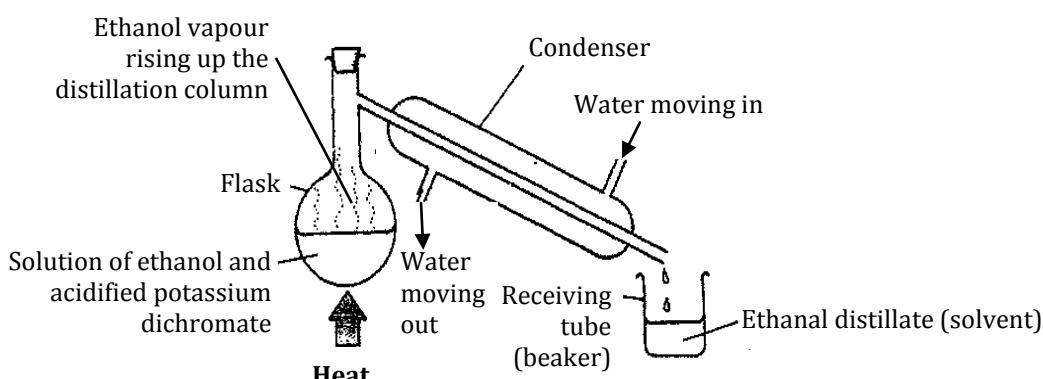


Fig 14.1. Experimental set up of preparation of ethanal.

CONDITION: Excess ethanol. Too much acidified potassium dichromate will mean oxidation will be complete to produce ethanoic acid and not ethanol.

On heating, ethanol vapour forms in the mixture flask. The heat vapourises the ethanal faster than ethanol because ethanal has a lower boiling point of 21°C than ethanol. Ethanal vapour rises up the distillation column and finally enters the condenser. There it is cooled and condenses to form liquid ethanal and the ethanal distillate is collected in the collection flask/beaker (receiving tube).

PROPERTIES OF ETHANAL

- ✓ Miscible in water
- ✓ Has a boiling point of 21°C.
- ✓ Colourless
- ✓ Has a choking smell.

H. BRADY'S TEST: TEST FOR C=O FUNCTIONAL GROUP

Both alkanals and ketones react with 2, 4 DNPH to produce a red precipitate.

PREPARATION

- ✓ Place a match head amount of 2,4 dinitrophenylhydrazine (2,4 DNPH)
- ✓ Add two drops of ethanol to moisten it
- ✓ Add two drops of concentrated sulphuric acid and a clear yellow liquid forms.
- ✓ Add one drop of the unknown

POSITIVE RESULT: A red precipitate forms

INTERPRETATION: The test liquid has a carbonyl functional group. It is either an alkanal or a ketone.

I. FEHLING'S TEST: DISTINGUISHING ALKANALS FROM KETONES

PREPARATION

- ✓ Add 15 drops of fresh Fehling's solution to a test tube.
- ✓ Add 15 drops of the unknown to the test tube
- ✓ Heat the contents gently for about a minute.

POSITIVE RESULT: A red precipitate forms.

INTERPRETATION: The test liquid is an alkanal.

8. POLYMERISATION

DEFINITION: A chemical reaction in which small molecules join together to form a long chain molecule.

MONOMERS are the small molecules that join up together to form long chain molecules.

POLYMERS are the long chain molecules produced by joining together many small molecules.

A. TYPES OF POLYMERISATION

I. ADDITION

Monomers usually of the same type join with each other to form polymers under great heating intense pressure.

II. CONDENSATION POLYMERISATION

This is a reaction whereby a small molecule usually water is eliminated between the different monomers resulting into the formation of a bond between the two large monomers.

B. TYPES OF POLYMERS

I. NATURAL/BIOPOLYMERS

➤ PROTEINS

MONOMERS: Amino acids

FORMATION: Condensation polymerisation

The amino acids join together in long chains linked by a peptide link.

It takes a minimum of 100 amino acids for a molecule to be a protein

Examples of amino acids

Fig 14.10. diagrams showing examples of amino acids

TWO GROUPS OF PROTEINS

a) FIBROUS PROTEINS

- ✓ Have linear molecules
- ✓ Insoluble in water
- ✓ Resistant to acids and alkalis

Examples: collagen in tendons and muscles, elastin in arteries, keratin in nails, hairs, horns and feathers.

b) FICOBULAR OR GLOBULAR PROTEINS

- ✓ Have complicated three dimensional structures
- ✓ Soluble in water
- ✓ Easily affected by acids, alkalis and temperature rise

Examples are casein in milk, albumen in egg white, haemoglobin in red blood cells and enzymes.

USES

- ✓ Growth and repairing worn out cells
- ✓ Structural materials in hair, feathers, nails and horns.
- ✓ Basic material for enzymes
- ✓ Making up the muscles fibres for locomotion.

Fig 14.11 diagram showing molecular structure of proteins

➤ DEOXYRIBONUCLEIC ACID (DNA)

MONOMER: Nucleotides

DNA is a member of nucleic acids

FUNCTION

- ✓ Controls protein synthesis in cells
- ✓ Makes up genes which carry hereditary information.

➤ CELLULOSE

MONOMER: Glucose

The cellulose is found in the walls of plant cells

USES

- ✓ Cellulose from wood is used for making paper.
- ✓ Cellulose from cotton is used for making clothes

➤ RUBBER

Rubber is a naturally occurring hydrocarbon polymer.

➤ SILK

More from Macmillan Chemistry

II. SYNTHETIC POLYMERS

➤ POLYTHENE

MONOMER: Ethene

FORMATION: Many ethane molecules join together in long chains under great heating and intense pressure. The ethane molecules break one of their double bonds releasing free bonds which join with each other forming polythene.

Fig 14.12. Molecular formation and structure of polythene

PROPERTIES

- ✓ Easily moulded into different shapes
- ✓ Good insulators
- ✓ Does not corrode
- ✓ Durable
- ✓ Not affected by weather

USES

- ✓ Making plastic bags
- ✓ Wrapping films
- ✓ Making fizzy drink bottles
- ✓ Electrical insulation
- ✓ Fabric coating
- ✓ Making milk bottle crates

➤ POLYVINYLCHLORIDE

MONOMER: monochloroethene ($\text{CH}_2=\text{CHCl}$)

FORMATION: Millions of monochloroethene join together in long chains supplied with intense heat and great pressure. The monomers break their double bonds and release free bonds which join with other monochloroethene molecules to form PVC.

PROPERTIES

- ✓ Strong
- ✓ Hard
- ✓ Less flexible than polythene

USES

- ✓ Electrical insulation
- ✓ Making pipes and garden hose
- ✓ Making raincoats
- ✓ Guttering
- ✓ Making records

➤ POLYSTYRENE

MONOMER: Phenylethene ($\text{C}_6\text{H}_5\text{CH}=\text{CH}_2$) commonly called Styrene.

FORMATION: Under great heating and intense and intense pressure many phenylethene molecules break their double bonds releasing free bonds which join with each other forming polystyrene.

Fig 14.13 formation and molecular structure of polystyrene

PROPERTIES

- ✓ Light
- ✓ Brittle

USES

- ✓ Making packaging
- ✓ Making toys
- ✓ Insulation
- ✓ Stems of ball point pens
- ✓ Making expanded polystyrene: This is made by blowing air into the plastic while still in liquid form.

➤ PERSPEX

MONOMER: Methyl-2, methylpropanoate.

FORMATION: Under great pressure and intense heating many methyl-2, methylpropanoate molecules break their double bonds releasing free bond which join with each other forming Perspex.

Fig 14.14. Formation and molecular structure of perspex

PROPERTIES

- ✓ Transparent

USE

- ✓ As a glass substitute

➤ POLYTETRAFLUROETHENE (*Teflon*)

MONOMER: Tetrafluroethene ($\text{CF}_2=\text{CF}_2$)

FORMATION: Under great pressure and intense heating many tetrafluroethene molecules break their double bonds releasing free bonds which join with each other forming Teflon.

PROPERTIES

- ✓ Anti-stick.
- ✓ Can withstand very high temperatures.
- ✓ Resistant to chemical attack
- ✓ Very stable
- ✓ Low friction

USES

- ✓ Making non stick coatings on pans.
- ✓ Insulation
- ✓ Making soles (insulation) of electric irons.

Synthetic polymers formed by condensation polymerisation are:

➤ **NYLON/POLYAMIDE**

MONOMER: 1, 6 diaminohexane and hexanedioic acid.

FORMATION: Under great heating and intense pressure many 1, 6 diaminohexane and hexanedioic join together to form nylon. They are joined by amide links giving the name polyamide.

Fig 14.15 formation and molecular structure of nylon

PROPERTIES

- ✓ Hard
- ✓ Pliable
- ✓ Mouldable

USES

- ✓ Making shirts
- ✓ Making jumpers
- ✓ Making tights
- ✓ Making curtain rails
- ✓ Making parachutes
- ✓ Making racket strings for lawn tennis and badminton

➤ **TERYLENE**

It belongs to a group of polyesters

MONOMERS: Ethane 1, 2 diol HO (CH₂)₂OH and Benzene 1, 4 dicarboxylic acid HOOC (C₆H₄) COOH

FORMATION: Under great heating and intense pressure the monomers join together to form Terylene joined by many ester links.

Fig 14.16. Formation and molecular structure of Terylene

- ❖ Terylene is softer than nylon.

USES

- ✓ Clothes
- ✓ Fibres
- ✓ Seat belts
- ✓ Plastic bottles
- ✓ Video tapes

C. FACTORS AFFECTING THE PROPERTIES OF POLYMERS

- a. Chain length: As chain length increases the strength of the IMF increases as well. Boiling points and melting points increase too.
- b. Branching: A highly branched polymer has lower tensile strength and melts more easily.
- c. Cross linking: Increases the rigidity and makes the polymer brittle.

9. PLASTICS

A. ADVANTAGES OF PLASTICS

- ✓ Cheaper
- ✓ flexible
- ✓ light
- ✓ easily coloured
- ✓ water proof
- ✓ do not corrode
- ✓ do not conduct electricity

B. DISADVANTAGES OF PLASTICS

- ✓ They are non biodegradable and so cause environmental and sanitation problems.

C. KINDS OF PLASTICS

I. Thermoplastics (thermosoftening plastics)

- ✓ Melt when heated and harden again when cooled because their polymer chains have no cross links.
- ✓ Have long linear chains with weak IMF
- ✓ Flexible, can stretch and melt at relatively low temperatures

EXAMPLES: polythene, PVC, polystyrene, nylon, Terylene and Teflon.

ADVANTAGE: Can be recycled

II. Thermosetting plastics (thermosets)

- ✓ Do not melt at high temperatures but just break and form dust.
- ✓ Are rigid and brittle

EXPLANATION: Have polymer chains linked to give a cross linked structure. The chains are held firmly and no softening takes place upon heating.

EXAMPLES: Melamine for table ware, phenolic for Formica and saucepan handles

D. PLASTIC WASTE MANAGEMENT

- ✓ Recycling
- ✓ Producing degradable plastics
- ✓ Having shopping bags made of paper and not plastics
- ✓ Incineration: Burning the waste plastics and using the energy for heating purposes in industries or in thermal power production.

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Samuel Kalea,