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CHEMISTRY
FOR SENIOR SECONDARY SCHOOLS**



Okoye E.I.

**MODEL TEXTBOOK OF
CHEMISTRY**

FOR

SENIOR SECONDARY SCHOOLS

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EBINCO Press

Rev. J.C. Okoye's Compound

Umuabu Village,

Adazi-Enu,

Anambra State,

08068575646

08038816693

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1st Published 2006

Revised Edition 2009

Third Edition 2012

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ISBN 978-076-161-6

Foreword

This latest edition of “Concise Chemistry” is written by E.I. Okoye. The text is a complete coverage of the Senior Secondary School Chemistry Syllabus. It is very comprehensive and will offer maximum help and guidance to the West African Chemistry Students who are preparing for their Senior Secondary School Certificate Examination (S.S.C.E.), those sitting for University/Polytechnic Matriculation Examinations and those in Higher Schools and first year Chemistry students in Tertiary Institutions.

I am therefore recommending this comprehensive textbook to all Students who want to understand the fundamentals of chemistry.

Dr. (Mrs) T.U. Onuegbu

Nnamdi Azikiwe University, Awka.

Preface

Concise Chemistry with workbooks is a complete coverage of the Senior Secondary School Chemistry Syllabus. It will provide a detailed theoretical and practical knowledge required for the Senior Secondary School Certificate Examination (S.S.C.E). It will also offer maximum help and guidance to the first year Chemistry Students in the Tertiary Institutions.

The book is divided into three parts which comprises the SS 1, SS 2 and SS3 Syllabus respectively. The scheme of work is followed strictly and religiously. The author therefore recommends this book to every Chemistry Teacher and Student for use in Senior Secondary Schools in West Africa.

The author's sincere gratitude is extended to Mr. Ideh C. (MSc.), the subject specialist who read the work and made necessary corrections. She also thanked the two English specialists who edited the work. They are: Mrs. Ikegbusi N.G. (M.Ed) and Mr. Nwofor G.O. (B.Ed).

Okoye E.I. Ph.D, MSc., MSc. Ed, BSc. Ed, N.C.E.

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CHAPTER 1

INTRODUCTION TO CHEMISTRY

Natural science is a systematic study of natural concept or phenomena obtained generally through controlled laboratory experimentation and observation. To facilitate study and investigation, natural science is compartmentalized or divided into different disciplines such as biology, physics and chemistry. Biology deals with the study of living things (plants and animals). Physics deals with the study of various forms of energy and matter.

Chemistry is therefore the branch of natural science that deals with the study of the composition of matter, its properties, the changes it undergoes, its reaction with other substances, its uses and synthesis. The term substance refers to any pure element or pure compound. Knowledge of chemistry is very significant in our life because almost everything in existence is produced by the application of the knowledge of chemistry. For example, plastic containers, foam, pomades, cements, soap, cutlery, textile materials, detergent, drugs, food, furniture and some drinks are all manufactured or made through the application of knowledge of chemistry. Chemistry is of vital importance in our daily life. Basic knowledge of chemistry enables us to face the challenges of modern society and to enjoy our lives. Both science and art students should realize the important role chemistry plays in their lives and should endeavour to have at least the basic knowledge of it. This is why we study integrated or basic science in the junior secondary school to show the unification and inter-relationship of scientific knowledge. Today, knowledge of chemistry is used either directly or indirectly in almost all spheres of human activities such as in the processing of cassava, production of shoe polish, hair relaxer and others.

1.1 Scientific Method: The procedure the scientists use in searching for scientific knowledge is known as scientific method. It involves the following steps:

1. Identification of the problem
2. Collection of relevant data
3. Formation of hypothesis
4. Experimentation to test the hypothesis
5. Observation of results of experiment
6. Writing out the results of the experiment.
7. Communicating the result of the experiment to other natural scientists who might wish to repeat the experiment or accept the findings.

Some methods used in the teaching of chemistry includes;

Discussion method, Lecture method, demonstration method, discovery method, field trips, laboratory method and project method.

1.1 **Hypothesis:** It is an idea put forward which has not been proved correct or wrong. It is an intelligent guess of an answer to the problem under investigation. It is an idea or answer, which has no evidence supporting it until a full investigation is carried out.

1.2 **Theory:** A theory is a scientifically accepted idea supported by scientific evidence. If a hypothesis is correctly proved by a scientific investigator or a scientist and other scientists accept the correctness of the hypothesis, then the hypothesis becomes a theory.

1.3 **Law:** A scientific law is a theory which has been proved correct by large number of scientists and from which valid predictions can be made.

1.4 **Careers in Chemistry:** Some of the employment opportunities in the area of chemistry include working in the laboratory as a laboratory technician, in an industry as a research and development chemist, process development chemist, environmental control chemist, analysis and quality control chemist, self reliant chemist, working in the laboratory as a laboratory assistant in the post primary school or institution of higher learning and teaching chemistry in post primary school or institution of higher learning.

Exercises

1. What is chemistry?
2. Of what importance is the subject chemistry?
3. Define “Hypothesis”
4. What employment opportunities are available in the area of chemistry?
5. What methods are used in the teaching of chemistry?
4. What employment opportunities are available in the area of chemistry?
5. What methods are used in the teaching of chemistry?

CHAPTER 2

CHEMICAL INDUSTRIES

2.1 Natural Resources in Nigeria

Some important raw materials available in Nigeria include:

1. Marble, starch, cellulose, animal protein, fats and oil, hides and skin, rubber, brine
2. Marble, aluminum, copper, iron ore, cola, tin ore, limestone, gold, bauxite, lead, zinc, clay, platinum, zinc oxide, zinc chloride, graphite, koalin, mica, columbite etc.
3. **Ethane, ethene, alkanols, alkanoates (esters), natural gas etc.**

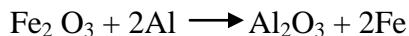
These raw materials are divided into two classes; organic and inorganic chemicals. The raw materials are also known as chemicals in that, they are utilised in the chemical industry in the production of other useful chemicals. A chemical industry is therefore an industry that uses the knowledge of chemistry in the manufacture of useful chemicals from other chemical substances. The raw materials obtained from plants and animals are known as organic chemicals, while the ones obtained or that can be synthesized in the laboratory from inorganic elements are called inorganic chemicals.

Starch is obtained in potatoes and cereal crops, such as maize, rice and wheat. It is used for the production of glucose. Animal and vegetable fats are used as; (i) foodstuffs (ii) production of soap and glycerol. Hides and skin unsuitable for the production of leather are used in the production of glue. Milk is obtained from animal protein.

Metallic Raw Materials

Aluminum being very light is used in house hold utensils, Aeroplane parts and surgical instruments.

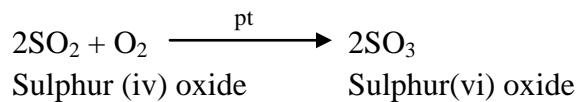
Metallic Aluminum is also used to reduce oxides of metals such as iron, chromium, vanadium and molybdenum, etc.



Due to the lightness of aluminum, it is used in constructing trucks, trains and ships, roofing sheets, stair rails and furniture. Aluminum is used in making cable. Because aluminum cannot react with beverages, it is used in brewery industry in making aluminum cans used in the preservation and storage of beverages. Aluminum is also used in the making of alum used in the purification of water.

Copper is used in the making of alloy called cupronickel. It is used in the making of nails e.g. copper nails. It is also used in the making of doorknobs, furnitures, coins and jewelries.

Platinum is used as a catalyst in a variety of reactions e.g. in the manufacture of sulphur(vi) Oxide.



Platinum is also used in plating other metals.

Zinc oxide is used as a white pigment in the making of white paint.

Zinc chloride is used as a deodorant and can also be used as a wood preservative.

Zinc is used widely for galvanizing iron and steel to prevent rusting. However, galvanized vessels cannot be used for storing food because of the danger of zinc poisoning.

Zinc is used in making alloys such as brass and bronze.

Mercury is used for the manufacture of industrial chemicals. It is also used in the making of some thermometers e.g. ordinary thermometer, clinical thermometer etc.

Limestone is used in the manufacture of Blackboard chalk, glass, cement, calcium hydroxide and steel. Marble and limestone are used as building materials.

2.2 Types of Chemical Industries

Chemical industry has already been defined. Since chemical industries produce a variety of products, they are classified on the basis of their products. Some chemical industries specialize on the production of:

Chemicals such as acids, alkalis, salts and organic chemicals

Chemical products like detergent, drug, soap and cosmetic for our personal use.

Chemical products used as starting materials in other industries that manufacture things like paints and fertilizers.

Chemical products used in the making of plastics and synthetic or artificial fibres.

Chemicals are sub-divided into two based on their quantitative production. They are divided into heavy and fine chemicals. Some basic chemicals manufactured in large quantity in the chemical industries are:

- Tetraoxosulphate (vi) acid: Its chemical formula is H_2SO_4 . It is used in the making of fertilizers, tetraoxosulphate (vi) salts, cleansing of (pickling) steel etc.
- Ammonia: Its chemical formula is NH_3 . It is used in the manufacture of fertilizers, explosives, ammonium hydroxide, ammonium salts etc.

- Sodium hydroxide. Its chemical formula is NaOH. It is used in the making of soap, detergent, sodium salts etc.

These and other chemicals produced in very large quantities are referred to as heavy chemicals.

Drugs and chemicals produced in small amount are called fine chemicals. They are produced to a very high degree of purity. Drugs used in medicine (e.g. paracetamol, antimalaria tablets or injections, multivite tablets or capsule) have to be manufactured in a very pure form.

2.3 Factors that Determines the Site of Chemical Industries

The factors that should be considered in location of chemical industries include:

- (i) Nearness of the chemical industry to the raw material or the easy accessibility of raw materials.
- (ii) Availability of labour, both skilled and unskilled.
- (iii) Abundance of cheap raw materials.
- (iv) Cheap and regular supply of fuel for the processing machines.
- (v) Transport facilities for export of finished products.
- (vi) Markets for the products

2.4 Importance of Chemical Industries to the Individual

Chemical industries provide employment to the educated individuals (people). They provide job to both skilled and unskilled individuals. They get employed in their areas of specialization. The mechanical Engineers will be employed in the industry's mechanic workshop. The laboratory scientist will be employed in their laboratory while the chemists are employed in their quality control section as quality control chemists. Unskilled individual are deployed as messengers and cleaners in the industry. Marketers and business managers are employed in their marketing section for the buying of raw materials and sales of the products. In this way, the standard of these individuals are raised. They use the salary they earned in maintaining themselves and their children by giving them good health care and education.

2.5 Importance of Chemical Industries to the Nation

Chemical industries provide job opportunity to the people of a particular nation. Chemical industry helps in raising the standard of living of a particular nation. It leads to the provision of good roads, electricity and pipe-born water supply in the community. Chemical industry provides better earnings for the nation or community. The products manufactured in the industries are sold to other nations. The nation therefore engages in international trade and earns better income.

2.6 Adverse Effects of Chemicals to Environment

Pollution

Pollution is a phenomenon that occurs when a pollutant is introduced into the atmosphere. If a pollutant is introduced into the air, it causes air pollution. If it is introduced into the water, it causes water pollution and if it is introduced into the soil, it causes soil pollution.

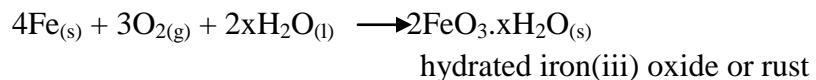
A Pollutant is a substance or a chemical which when introduced into the atmosphere causes atmospheric pollution. It can also be defined as a substance which when introduced into an environment, makes the environment unfit or unhealthy for the life of living things in that environment. Pollutants can be classified into three. These are: Air pollutants, water pollutants and land or soil pollutants. Air pollutants are airborne materials such as dust, smoke and soot, gases such as sulphur (iv) oxide (SO_2), Carbon (ii) oxide (CO), oxides of nitrogen and hydrogen sulphide (H_2S). Others include insecticides, herbicides, etc.

Water pollutants are too many. This is because, some air and soil pollutants can dissolve either directly or indirectly in a body of water to cause water pollution. This means that some air and soil pollutants can also be listed as water pollutants. Some examples of water pollutants are SO_2 , CO , nitrogen (iv) oxide (NO_2), nitrogen (ii) oxide (NO), fertilizers, heat, mercury, lead, synthetic organic compounds such as insecticides, fungicides, herbicides and oil.

Land or soil pollutants include plastics, tyres, old cars, solid waste from kitchens and industries such as papers, cans, vegetable materials, soap and detergents, fertilizers, faeces and dungs etc.

Corrosion of Metals

The Corrosion of metals results from the combined action of atmospheric oxygen and water. It is generally accelerated by the presence of carbon(iv) oxide and gaseous pollutants like sulphur(iv) oxide in the air. These gases pollutants are introduced into the atmosphere by the chemical industries. The corrosion of iron is known as rusting. Iron rusts by combining or reacting with oxygen in the presence of water to form brown hydrated iron (iii) oxide; $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$. This brown compound; $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$ is commonly known as rust.



Drug Abuse

Sources of drugs are mainly from plants. Some of these drugs are medicinal while others are poisonous. Plants are the sources of the most powerful hard drugs in our society. Some of these natural hard drugs obtained from plants are alcohol, Nicotine, Caffeine, Morphine, Cocaine, Reserpine, Heroine, Atropine, Codeine, Quinine, Strychnine and others. Adolescents abuse drugs

by either inhaling, smoking, drinking or chewing them. These hard drugs obtained from plants are used by the chemical industries in synthesising useful medicines.

Common Ways of Misusing Drugs and Their Effects

Adolescents misuse drugs by taking drugs that are not prescribed by doctors or medical practitioners. Sometimes they take underdose of legally prescribed drugs or they take too much of legally prescribed drug hoping that by so doing, their illness will be cured quickly.

The use of drugs among adolescents is on the increase. Some of the most commonly abused drugs include Indian hemp, heroin, cocaine and alcohol. Everyday, young people and adults use drugs, alcohol and other mind - altering substances that damage health and compromise their well-being. The reasons for this are varied and complex. Some people fall prey to the use of substances like alcohol and tobacco through clever advertising strategies designed to encourage consumption of these products. Others use drugs to lessen feelings of failure, disappointment and emotional pain. Some people seek the new experiences and thrills drugs promise. Others use drugs for no other reason than to feel a part of the group. The teen years are known to be a time when young people will want to experiment, without fully thinking through possible consequences. Some people become addicted to drugs. Addiction occurs when the normal functions of the body are altered in such a way that the body requires the continued presence of the drug to function. Addiction can be psychological or physical depending on the drug. The person who is addicted cannot function normally without the drugs of addiction.

Many people cannot function without the stimulant caffeine, found in coffee, tea and chocolate. Some cannot socialize without alcohol or marijuana. Without the drugs, the addicts feel poorly, are anxious and restless and may even neglect themselves. Drug addiction is a serious problem requiring professional help to overcome.

Exercises

1. Name six important raw materials available in Nigeria
2. Briefly define these terms:
 - (i) Chemical Industry
 - (ii) Organic chemical
 - (iii) Inorganic chemical
3. Write two uses of these chemicals or metals:
 - (i) Aluminum (ii) Copper
 - (iii) Platinium (iv) Zinc

(v) Mercury

4. Differentiate between (i) heavy and fine chemicals
(ii) Write two examples of each.
5. Mention two factors that should be considered in location of chemical industries
6. List two importance of chemical industries to:
 - (i) an individual (ii) a community or nation
7. Define these terms (i) pollution (ii) pollutant
(iii) air pollution
8. What are the three classes of pollutants?
(ii) Write three examples of each
9. What are the sources of some of the most powerful hard drugs in our environment or in our society.
10. Write some natural hard drugs that are abused by young people in our society.
11. What are the reasons why adolescents abuse drugs?
12. (a) Write an equation to illustrate Corrosion of metals.
(b) What is the IUPAC nomenclature of ‘rust’?

CHAPTER 3

PARTICULATE NATURE OF MATTER

The concept or idea that matter is made of particles is referred to as the particulate nature of matter. We have seen that matter is made up of particles from the diagrammatic representation of the three states of matter. The fact that particles of matter are in constant motion can be seen from the fact that when a reagent bottle containing ammonia is opened in the chemistry laboratory, the odour will be perceived by somebody standing outside the laboratory after some seconds. Again if hydrogen sulphide gas is generated inside a chemistry laboratory, the odour of the gas can easily be perceived at different distances within the environment, or when a gas cylinder is opened in the laboratory without lighting the bunsen burner, the odour of the gas can also be perceived by people at different distances within the environment. If matter were not made of particles that were in constant motion, the odour of the gas will not be perceived at different distances from the source. Scientists discovered that matter is made up of particles which can be atoms, molecules or ions.

3.1 Atom:

An atom is the smallest particle of an element, which can take part in a chemical reaction. If a piece of solid element like copper or zinc is ground into very tiny pieces, the smallest part of it which can take part in a chemical reaction is called an atom.

3.2 Molecule:

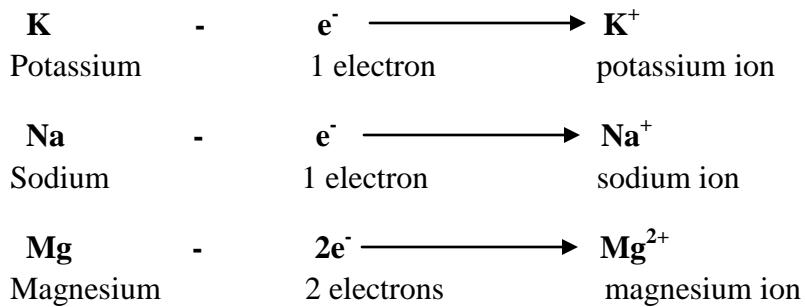
A molecule of a compound or element is the smallest particle of a compound or element which is capable of independent existence. A molecule may be composed of atoms of the same element or atoms of different elements. For example, a nitrogen molecule (N_2) contains two atoms of nitrogen, an oxygen molecule (O_2) contains two atoms of oxygen but a molecule of water (H_2O) contains two atoms of hydrogen and one atom of oxygen.

3.3 Ion:

An ion is an electrically charged atom or group of atoms. It is formed as a result of the loss or gain of electrons. The electrons lost or gained are equal to the valence of the ions. There are two types of ions. These are:

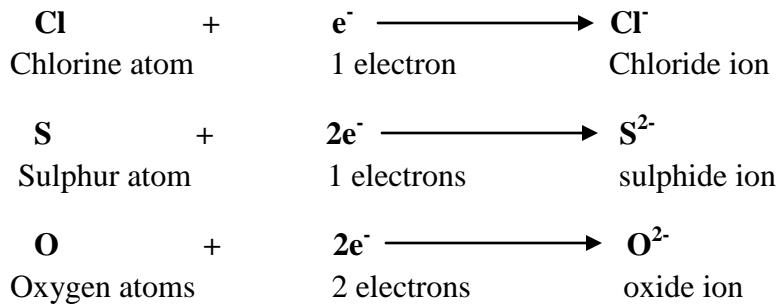
- i Positively charged ion or cation
- ii Negatively charged ion or anion.

A cation is formed when an atom or group loses electrons. For example



Cations include ammonium ion (NH_4^+), hydrogen ion (H^+) and metallic ions.

An anion is formed when an atom or group gains electrons. For example;

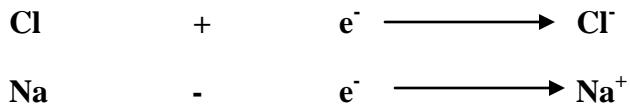


Anions include hydroxide ion and acid radicals. Examples of acid radical are:



3.4 Neutral Atom:

A neutral atom is an element in which the number of protons in the nucleus is equal to the number of electrons surrounding the nucleus. If an atom gains or loses an electron or electrons, it becomes an ion. So an ion may be defined again as an atom, which has lost or gained an electron or electrons. The loss or gain of electron can be caused by the chemical environment of the atom eg in electrolysis or by the absorption of energy eg the excited state of an atom results in the formation of negatively charged ion but the loss of electron results in the formation of positively charged ion as shown below.



In any electrically neutral substance the number of positive charge on the cation must be equal to the number of negative charge on the anion. For example, in (K^+Cl^-) the number of charge on the potassium ion is 1 and that on the chloride ion is 1, as a result, potassium chloride (KCl) is electrically neutral.

3.5 Atomicity:

Atomicity of an element is the number of atoms in one molecule of the element.

Elements whose molecule contains one atom is said to be mono-atomic e.g copper (Cu), sodium (Na).

Element whose molecule contains two atoms are said to be diatomic e.g O_2 . Element whose molecule contains three atoms are tri-atomic eg O_3 . Element whose molecule contains four atoms are said to be tetra-atomic while a molecule containing more than four atoms is said to be poly-atomic eg S_8 .

ELEMENT	MOLECULAR FORMULA	ATOMICITY
Carbon	C	1
Copper	Cu	1
Sodium	Na	1
Hydrogen	H_2	2
Oxygen	O_2	2
Nitrogen	N_2	2
Ozone	O_3	3
Phosphorus	P_4	4
Sulphur	S_8	8

3.6 Dalton's Atomic Theory:

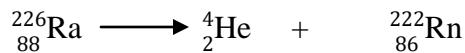
Dalton's atomic theory is the idea put forward to explain atom. It was first postulated in the year 1808 by a scientist called John Dalton.

Dalton's atomic theory states that:

1. Matter is made up of small indivisible particles called atoms.
2. Atoms of a given element are all exactly alike.
3. Atoms cannot be created or destroyed.
4. Atoms combine in small whole numbers.

3.7 Recent Modification of Dalton's Atomic Theory:

1. Number one postulate is no longer acceptable because radioactive elements are dividing spontaneously. The atomic nucleus is giving out particles and so producing two less complex atoms, eg radium disintegrates to produce two noble gases, helium and radon.



1. Number two postulate is no longer acceptable because of isotopy. Isotopy is a phenomenon that occurs when an element can exist in two or more forms having similar chemical properties but different atomic masses. Thus chlorine has isotope ${}^{35}_{17}\text{Cl}$ and ${}^{37}_{17}\text{Cl}$. Both have the same number of electrons in their outermost shells. So they have the same atomic number and exhibit the same chemical properties, but the 37-isotope has two extra neutron in the nucleus and so is the higher atom. Potassium also have isotope ${}^{39}_{19}\text{K}$ and ${}^{41}_{19}\text{K}$.
2. Number three postulate is only acceptable and applicable to chemical reactions, in which atoms react as whole units. The changes that occurs as a result of atomic fission, destroy atoms of the element involved. The nuclei are broken into smaller units forming simpler atoms (atoms are created). The nucleus of the uranium isotope U=235, can absorb a neutron and break up into two unequal atoms with mass numbers of 95 and 140.
3. Number four theory is still acceptable for most elements in inorganic chemistry. In organic chemistry, the statement is no longer acceptable because carbon can form very complex organic compounds that do not react in small whole numbers. Examples of complex organic compounds are:

Name	Molecular Formula
Heptane	C ₇ H ₁₆
Octane	C ₈ H ₁₈
Nonane	C ₉ H ₂₀
Decane	C ₁₀ H ₂₂

3.8 Atomic Structure:

An atom is made up of three fundamental particles. These are electron, proton and neutron. An electron is negatively charged. It has a mass of 1/1840 Unit. It is found in the shell or orbit. A proton is positively charged. It has a mass of 1 unit. It is found in the nucleus. A neutron has no charge. It has a mass of 1 unit. It is found in the nucleus.

Characteristics of the fundamental particles of an atom.

Name of Particle	Charge	Mass based on carbon-12 Isotope	Location in the atom
Electron	-1 (negative)	1/1840	In the shell or orbit
Proton	+1 (positive)	1	Inside the nucleus
Neutron	0 (neutral)	1	Inside the nucleus

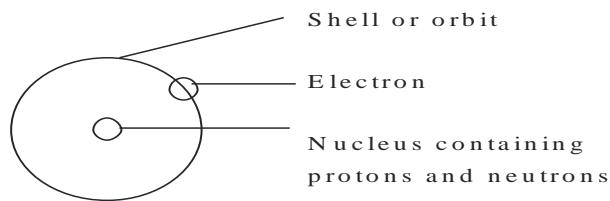


Fig. 3.1 structure of an atom.

3.9 Atomic Number:

Atomic number is the number of proton in the nucleus of an atom. For instance, if an atom has 11 protons, the atomic number is 11. In an atom, the number of proton is equal to the number of electron. An atom is therefore said to be electrically neutral because of the equality of electrons and protons.

3.10 Mass Number:

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

$$\text{Mass no.} = \text{no. of protons} + \text{no. of neutrons}$$

Mass number is also called an atomic mass. An atom with 10 protons and 12 neutrons has an atomic mass or mass number of 22.

An element X with mass number 14 and the atomic number of 7 has the symbol ${}_{7}^{14}\text{X}$

Thus an element ${}_{7}^{14}\text{X}$ has mass number of 14 and atomic number of 7.

3.11 Isotopy:

Isotopy is the occurrence of atoms of the same element having the same atomic number but different mass number. These atoms of the same element having the same atomic number but different mass number are called isotopes. The difference in the mass number of these isotopes are due to difference in the number of neutron in their nuclei.

Examples of elements and their isotopes:

Hydrogen: It has three isotopes

- i. Protium ${}_1^1\text{H}$
- ii. Deuterium ${}_1^2\text{H}$
- iii. Tritium ${}_1^3\text{H}$

Chlorine: It has two isotopes; chlorine 35 and chlorine 37. They are represented as follows:



Exercises:

1. What do you understand by these terms?
 - I. Particulate nature of matter
 - II. Atom, III. Molecules, IV. Ion, V. Atomicity of an element.
2. What are the four postulates of John Dalton in the year 1808?
3. What are the recent modification of two of the Dalton's atomic theory?
4. What are the three fundamental particles of an atom? State their masses and the charge carried by each one of them.
5. What is an isotope? Give two examples of it.
6. Calculate the mass number of these atoms.
 - i. An atom with 10 electrons, 10 protons and 14 neutrons.
 - ii. An atom with 14 electrons, 14 protons and no neutron.

CHAPTER 4

SYMBOLS, VALENCIES AND CHEMICAL FORMULAE

Symbol of an element is a simple method of representing an element. Chemical formula is a simple system of representing molecules of elements or compounds. Element are divided into metals and non-metals. Physical and chemical difference between metals and Non-metals is discussed in chapters 2.16 and 2.17 of this book. Some examples of metals are:

Metals	Symbols
Potassium	K
Sodium	Na
Calcium	Ca
Magnesium	Mg
Zinc	Zn
Iron	Fe
Lead	Pb
Copper	Cu
Mercury	Hg
Silver	Ag
Gold	Au

Some examples of non-metals are

Non-metals	Symbols
Oxygen	O
Sulphur	S
Chlorine	Cl
Nitrogen	N
Iodine	I
Hydrogen	H
Carbon	C
Bromine	Br

4.1 Elements and their Valencies

Valency: The valency of an element is the number of hydrogen atoms that can combine with or replace one atom of the element. Valency of an element is also called the combining power of the element or the oxidation number of the element. Some elements have more than one valency and they are said to have variable valences.

Element	Symbol	Valency
Potassium	K	1
Sodium	Na	1
Calcium	Ca	2
Magnesium	Mg	2
Aluminum	Al	3
Zinc	Zn	2
Iron	Fe	2 or 3
Lead	Pb	2 or 4
Copper	Cu	1 or 2
Mercury	Hg	1 or 2
Silver	Ag	1
Hydrogen	H	1
Chlorine	Cl	1
Bromine	Br	1
Iodine	I	1
Oxygen	O	2
Nitrogen	N	3 or 5
Sulphur	S	3, 4 or 6
Carbon	C	2 or 4
Phosphorus	P	3 or 5

4.2 Radicals:

A radical is a charged group of atoms. It behaves as a single unit. Eg SO_4^{2-} etc

Ion: An ion is an atom or group of atoms that is charged e.g Ca^{2+} , SO_4^{2-} etc. Read chapter 4.3 of this book for more detailed explanation.

Radical	Formula	Valency
Tetraoxosulphate (vi)	SO_4^{2-}	-2
Trioxonitrate (v)	NO_3^-	-1
Trioxocarbonate (iv)	CO_3^{2-}	-2
Hydroxyl	OH^-	-1
Hydrogen trioxocarbonate (iv)	HCO_3^-	-1
Hydrogen tetraoxosulphate (vi)	HSO_4^-	-1
Ammonium	NH_4^+	+1
Trioxochlorate (vi)	ClO_3^-	-1

4.3 Formula of Compound

Compounds and radicals are represented by chemical formula. A chemical formula is a short hand method of representing a compound. A symbol is to an element as a chemical formula is to a compound. A chemical formula of a compound tells us:

- i. The different types of element that make up the compound.
- ii. The number of atoms of each type of element present in the compound.

4.4 Binary Compounds:

Binary compounds are compounds containing only two elements and they always end in (ide) eg

NaCl	=	Sodium Chloride
NaH	=	Sodium hydride
H ₂ S	=	Hydrogen sulphide
AgI	=	Silver iodide

4.5 How to Write Formula of Compounds:

Valenceis of elements help us to write the formula of compounds. When two or more elements combine chemically to form a compound, the valency of each element will tell us the number of atoms of X that will combine with Y:

Example 1:

If X has valency of +1 and Y has valency of 1,

both X and Y have equal valences. Therefore 1 atom of X will combine with 1 atom of Y and the formula of the compound formed by X and Y is XY

The rule to be followed in writing a chemical formula of a compound is as follows:

Rule 1:Write the symbols of the elements involved or Write the symbol and formula of the element and the radical involved.

Rule 2:The valency or the oxidation numbers of the elements are written down or write down the oxidation numbers of the element and the radical involved.

Rule 3:Interchange the oxidation numbers, omitting the signs. Write them below and to the right of the symbols.

Example 2:

Hydrogen has valency of +1, oxygen has valency of -2. When hydrogen combines with oxygen to form water, the formula of the compound water is H_2O .

$$\text{H} = +1$$

$$\text{O} = -2 \quad \text{Formula of water} = \text{H}_2\text{O}.$$

Example 3:

Aluminum has valency of +3, Oxygen has valency of -2.

$$\text{Al} = +3$$

$$\text{O} = -2$$

The formula of aluminum oxide is Al_2O_3 .

Example 4:

Write the formula of calcium oxide, given that the valency of:

$$\text{Ca} = +2$$

$$\text{O} = -2$$

Since their valencies are equal, it means that one atom of each will combine to form a compound CaO and not Ca_2O_2 . When the valency of one element is a multiple of the valency of the other element, the simple ratio of their valences are the number of their atoms present in the compound.

For instance, the valency of:

$$\text{C} = -4$$

$$\text{O} = -2$$

The formula of carbon (iv) oxide is not C_2O_4 but CO_2 because the simplest ratio of 2:4 is 1:2.

Example 6:

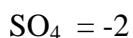
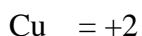
When a metal combines with a non-metal, you write the metal first before the non-metal or radical. E.g. Write the formula of compound formed between potassium and chlorine.

$$\text{K} = +1$$

$$\text{Cl} = -1$$

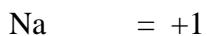
Formula = KCl.

- a. Write the formula of copper (ii) tetraoxosulphate (vi)



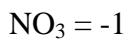
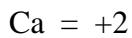
They have equal valences. Therefore the formula is CuSO_4

- b. Write the formula of sodium trioxocarbonate (iv)



The formula is Na_2CO_3

- a. Write the formula of calcium trioxonitrate (v).



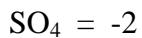
The formula is $\text{Ca}(\text{NO}_3)_2$

Here, we place the NO_3 in a bracket before the subscript 2, to indicate that NO_3 is an entity and there are two of that entity in the compound.

Example 7:

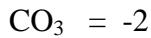
When hydrogen combines with a radical, an acid is formed. When a metal combines with a radical, a salt is formed. Common acid radicals are: SO_4^{2-} , CO_3^{2-} , NO_3^{-} , Cl^-

- a. $\text{H} = +1$



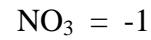
Formula = H_2SO_4 Tetraoxosulphate (vi) acid

- b. $\text{H} = +1$



Formula = H_2CO_3 Trioxocarbonate (iv) acid

- c. $\text{H} = +1$



Formula = HNO_3 Trioxonitrate (v) acid

d. H = +1

Cl = -1

Formula = HCl Hydrochloric acid

4.6 Types O Chemical Formula: Empirical, Molecular and Structural Formulae:

Empirical Formula:

Empirical formula is the simplest formula of a compound. It is the smallest ratio of different atoms present in one molecule of a compound. It shows the relative numbers of each kind of atom in a molecule of a compound. The empirical formula of these compounds are:

Compound	Empirical Formula
Butane	C_2H_5
Ethanoic acid	CH_2O
Ethane	CH_3

Calculation of Empirical Formula of Compound:

Example 1:

What is the empirical formula of an organic compound, which contains 80% Carbon and 20% hydrogen. This means that in every 100g of the organic compound, the masses of carbon and hydrogen are 80g and 20g respectively. (C = 12, H = 1)

Element	Carbon	Hydrogen
Composition	80	20
Divide by the atomic mass of each element	$\frac{80}{12} = 6.7$	$\frac{20}{1} = 20$
Divide by the smallest number	$\frac{6.7}{6.7} = 1$	$\frac{20}{6.7} = 3$

The ratio of carbon to hydrogen = 1:3

Therefore empirical formula = CH_3

Example 2:

What is the empirical formula of a molecule which contains 40.4% calcium, 11.8% carbon and 47.8% Oxygen? ($\text{Ca} = 40$, $\text{C} = 12$, $\text{O} = 16$)

Element	Calcium	Carbon	Oxygen
Percentage Composition	40.4	11.8	47.8
Ratio of number of atoms = % composition	40.4	11.8	47.8
Relative atomic mass	40	12	16
: Ratio of number of atoms 0.98	= 1.01	= 0.98	= 2.98
	$\frac{1.01}{0.98} = 1$	$\frac{0.98}{0.98} = 1$	$\frac{2.98}{0.98} = 3$

The empirical formula is CaCO_3

Molecular Formular:

Molecular formular of a compound is a whole number multiple of its empirical formula. It is the formula of a compound, which expresses the actual number of each kind of atom present in a molecule of the compound. Molecular formula is derived from the empirical formula of the compound.

The molecular formular of these compounds are:

Compound	Molecular Formula	Empirical Formula
Butane	C_4H_{10}	C_2H_5
Ethanoic acid	$\text{C}_2\text{H}_4\text{O}_2$	CH_2O
Ethane	C_2H_6	CH_3

Calculation of Molecular Formula of Compounds:

In order to determine the molecular formula, we need to know the relative molecular mass of the compound.

Example 1:

What is the molecular formula of a compound which on analysis, its empirical formula is found to be CH_2 and its relative molecular mass is known to be 42. ($\text{C} = 12$, $\text{H} = 1$).

The relative molecular mass = 42

$$\text{CH}_2 = 12 + (1 \times 2) = 14.$$

Therefore, the formula mass of (the empirical formula) is 14.

$$\frac{\text{The relative molecular mass}}{\text{The formula mass}} = \frac{42}{14} = 3$$

$$3 \times \text{CH}_2 = 3 \times \text{C}_2\text{H}_4 = \text{C}_3\text{H}_6$$

Example 2:

An organic compound on analysis yielded 2.0g of carbon, 0.34g of hydrogen and 2.67g of Oxygen. If the molar mass or the molecular mass of the compound is 180g, calculate the molecular formula of the compound. ($\text{C} = 12$, $\text{H} = 1$, $\text{O} = 16$)

Element	Carbon	Hydrogen	Oxygen
Composition by mass	2.0g	0.34g	2.67g
Divide by the atomic mass of each element.	$\frac{2.0}{12} = 0.17$	$\frac{0.34}{1} = 0.34$	$\frac{2.67}{16} = 0.17$
Divide by the smallest number	$\frac{0.17}{0.17} = 1$	$\frac{0.34}{0.17} = 2$	$\frac{0.17}{0.17} = 1$

$$\text{Ratio} = \text{C} : \text{H} : \text{O}$$

$$1 : 2 : 1$$

The empirical formula is CH_2O

Let the molecular formula of the compound be $(CH_2O)_n$, where n is a whole number.

Molecular mass of the compound is 180

$$\therefore (CH_2O)_n = 180$$

$$\therefore C_n + H_{2n} + O_n = 180$$

$$12n + 2n + 16n = 180$$

$$30n = 180$$

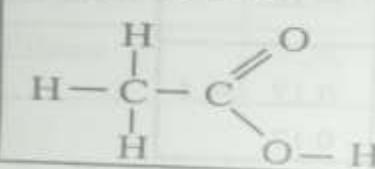
$$n = \frac{180}{30} = 6$$

$$n = 6$$

Therefore the molecular formula is $(CH_2O)_6 = C_6H_{12}O_6$

Structural Formula:

Structural formula of a compound is a formula which shows the arrangement of the atoms of a compound in space. It is the formula that shows the special arrangement of atoms of a particular compound. In organic chemistry, a molecular formula may represent more than one substance. A structural formula which is more informative than the molecular formula is therefore used to represent a particular organic substance. It shows how and which atoms are linked together. A structural formula of an organic compound can also be written in a condensed form. Eg.

Structural Formula	Condensed Formula
Ethanoic acid 	Ethanoic acid CH ₃ COOH

Compound	Structural Formula	Molecular Formula	Empirical Formula
Butane	 $\begin{array}{ccccccc} & \text{H} & \text{H} & \text{H} & \text{H} & & \\ & & & & & & \\ \text{H} & - & \text{C} & - & \text{C} & - & \text{C} & - & \text{C} & - & \text{H} \\ & & & & & & \\ & \text{H} & \text{H} & \text{H} & \text{H} & & \end{array}$	C_4H_{10}	C_2H_5
Ethanoic acid	 $\begin{array}{ccccc} & \text{H} & & \text{O} & \\ & & & \diagup & \\ \text{H} & - & \text{C} & - & \text{O} \\ & & & \diagdown & \\ & \text{H} & & \text{O} & - & \text{H} \end{array}$ <p>Condensed Structural Formula is CH_3COOH</p>	$\text{C}_2\text{H}_4\text{O}_2$	CH_3O
Ethane	 $\begin{array}{ccccc} & \text{H} & & \text{H} & \\ & & & & \\ \text{H} & - & \text{C} & - & \text{C} & - & \text{H} \\ & & & & & \\ & \text{H} & & \text{H} & & \end{array}$	$\boxed{\text{C}_2\text{H}_6}$	$\boxed{\text{CH}_3}$

Exercises:

- Define these terms and give an example of each.
 - Symbol
 - Chemical formula
 - Valency
 - Radical
 - Binary compound
- Write the formula and the IUPAC name of the compounds formed by these pairs of substances.
 - $\text{Al} = +3, \text{Cl} = -1$
 - $\text{K} = +1, \text{O} = -2$
 - $\text{Na} = +1, \text{SO}_4 = -2$
 - $\text{H} = +1, \text{CO}_3 = -2$
 - $\text{Ca} = +2, \text{NO}_3 = -1$
- a. What do you understand by these terms
 - Empirical formula?
 - Molecular formula?
 - Structural formula?
- Write an example of each using a particular compound to show their relationship.

4. What is the empirical formula of a compound which contains 40.1% nitrogen, 14.3% hydrogen and 45.6% oxygen? (N = 14, H = 1, O = 16).
5. What is the empirical formula of an inorganic compound which is found to contain 42.5% chlorine and 57.5% oxygen? If the molecular mass of the compound is 167, what is its molecular formula? (Cl = 35.5, O = 16)

CHAPTER 5

THE LAWS OF CHEMICAL COMBINATION

The four laws of chemical combination are:

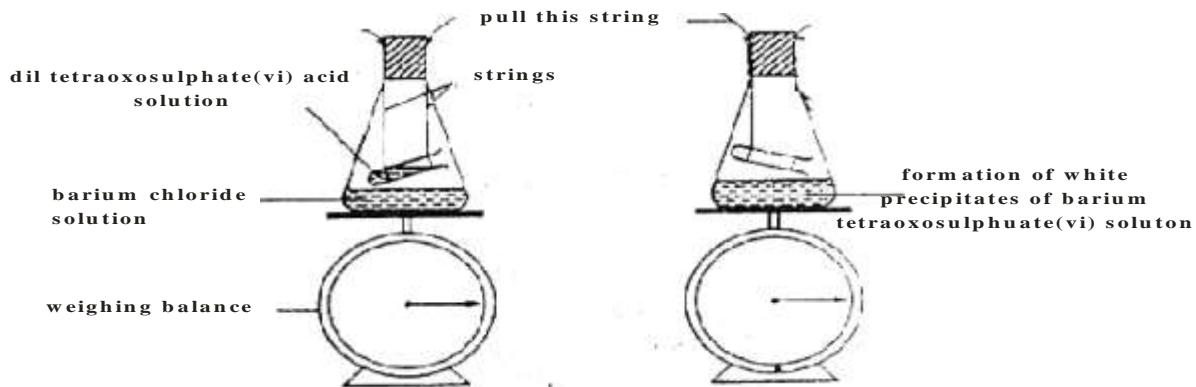
1. The law of conservation of matter or the law of indestructibility of matter.
2. The law of definite proportion or the law of constant composition.
3. The law of multiple proportions.
4. The law of reciprocal proportion

5.1 The Law of Conservation of Matter or Mass.

This law is also known as the law of indestructibility of matter. It states that matter can neither be created nor destroyed in the course of a chemical reaction, but changes from one form to another.

Verification of the Law of Conservation of Mass

The commonest reagents in the chemistry laboratory that can be used to verify this law are dilute tetraoxosulphate (vi) acid and barium chloride solution.



Method:

1. Pour small quantity of dilute tetraoxosulphate(vi) acid into the conical flask.
2. Suspend a small test tube filled with barium chloride solution in the conical flask by means of two threads. See the above diagram.
3. Cork the flask.

4. Weigh the flask and its contents with a chemical balance and record the mass.
5. Pour the barium chloride solution into the dilute tetraoxosulphate (vi) acid by pulling the thread attached to the bottom end of the small test tube.
6. Shake the flask carefully to ensure complete reaction.
7. Then weigh the flask and its contents again. Record the mass.

Equation of the reaction



Observation

When the two reactants come into contact, there is a visible reaction. A white precipitate of barium tetraoxosulphate (vi) is formed. This showed that a chemical change has occurred. The masses of the flask and its content are the same. It is therefore inferred that since the content are the same, before and after the reaction, the masses of the reactants and the products are equal. This experiment is a proof of the above law.

5.2 The Law of Definite Proportion or Constant Composition.

This law states that all pure samples of the same chemical compound (no matter the method of preparation) contained the same elements and combined in the same proportion by mass.

Consider a pure water for instance, it does not matter whether its source is a river, well, sea, rain, spring etc., on electrolysis of acidified water, it must always contain hydrogen and oxygen in the ratio of 2:1.

Verification of the law of definite proportion

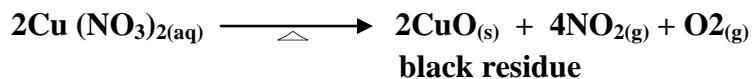
This law can be verified in the laboratory by:

1. Using two or three different methods to prepare pure samples of the same compound, and
2. Analysing the samples to find out whether they contain the same elements and are combined in the same proportion by mass.

Method:

Sample A is prepared by dissolving copper metal completely in Conc HNO_3 . The resultant green solution of $\text{Cu}(\text{NO}_3)_2$ formed is evaporated to dryness. Application of heat is continued till the residue decomposes to give black solid which is CuO . Keep this copper (ii) oxide dry in a desiccator.





Sample B is prepared by heating (strongly) copper (ii) trioxocarbonate (iv) in a crucible till it decomposes to black residue (copper (ii) oxide) and carbon (iv) oxide. Store the black residue in a desiccators



Analysi

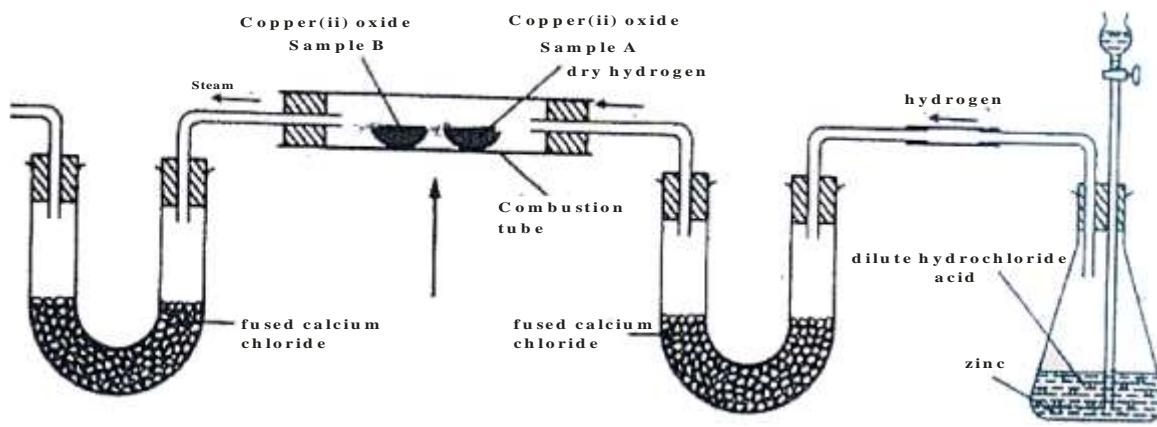
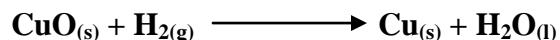


Fig. 5.2

Method

1. Samples A and B are placed in two heated porcelain boats in a combustion tube.
2. A generated stream of dried hydrogen is made to pass through them. Each of the samples is reduced to copper. The steam formed by the reaction between hydrogen and oxygen is absorbed by the fused calcium chloride. The mass and percentage of copper in each sample is calculated when the copper residue is cooled.



	RESULT	Sample A	B
a	Mass of porcelain boat	= 7.5g	8.5g
b	" " " " + CuO	= 12.5g	13.5g
c	" " CuO	= 5g	5g
		Sample A	B
d	Mass of porcelain boat + Cu residue	= 11.5g	12.5g
e	Mass of Oxygen = (b-d) = 12.5-11.5	= 1g	1g
	Mass of copper (Cu) = (e-e)	= 4g	4g

Percentage of copper present =

$$\text{Sample A} = \frac{4}{5} \times \frac{100}{1} = 80\%$$

$$\text{Sample B} = \frac{4}{5} \times \frac{100}{1} = 80\%$$

Percentage of oxygen present =

$$\text{Sample A} = \frac{1}{5} \times \frac{100}{1} = 20\%$$

$$\text{Sample B} = \frac{1}{5} \times \frac{100}{1} = 20\%$$

Therefore % of Cu:O in CuO is



Conclusion:

Pure samples of copper (ii) oxide prepared using different methods contain the same elements copper and oxygen and are present in a definite proportion or in a constant composition.

5.3 The Law of Multiple Proportion

The law states that if two elements X and Y combine to form more than one compound, then the masses of X which separately combine with a fixed mass of Y must be in simple multiple ratio.

Some examples of paired elements that can obey this law are:

1. Copper and oxygen can combine to form more than one compound which are (a) black copper (ii) oxide (CuO) and red copper (i) oxide Cu_2O .
2. Iron and chlorine combine to form two compounds-yellow iron (iii) Chloride, FeCl_3 and green iron (ii) chloride, FeCl_2 .
3. Carbon and oxygen combine to form CO and CO_2 .
4. Iron and oxygen combine to form brown Fe_2O_3 and black FeO .

Verification of the law of multiple proportion.

Method:

1. Weigh two clean crucible labeled X and Y.
2. Place a known quantity of Cu_2O in X and CuO in Y.
3. Put them in a combustion tube and heat very strongly.
4. Reduce them to pure copper by passing a stream of hydrogen through the tube.

Result:

It was found that the different masses of copper which combine with a fixed mass of oxygen to form Cu_2O and CuO is in the ratio of 2:1.

Again, it was found that in the formation of CO and CO_2 , the different masses of oxygen which combine with a fixed mass of carbon is in the ratio of 1:2 ie CO and CO_2 .

12g of carbon combine with 16g of oxygen to form CO and 32g of oxygen to form CO_2 . therefore, the several masses of oxygen that combine with fixed mass of carbon is $16: 32 = 1:2$.

5.4 The Law of Reciprocal Proportions:

This law states that the masses of different elements, A, B, C, which combine separately with a fixed mass of another element D, are the same as, or simple multiples of the masses in which A, B, C themselves combine with one another.

5.5 Calculation Based on the Laws of Chemical Combination:

Example 1:

Three oxides of nitrogen contain 53.5%, 69.6% and 36.4% oxygen respectively. What law of chemical combination is illustrated by this data?

% of oxygen obtained on analysis 36.4%, 53.5%, 69.6%

Ratio = 36.4 : 53.5 : 69.6

$$1 : 1.5 : 1.9$$

Ratio = 1 : 2 : 2

The law illustrated is the law of multiple proportion.

Example 2:

Two oxides of lead, A and B weighing 0.446g and 0.717g respectively, on reduction with hydrogen left 0.414g and 0.621g of lead respectively, show that these figures are in accordance with the law of multiple proportion.

		SAMPLE A	SAMPLE B
a	Mass of lead oxide	0.446g	0.717g
b	Mass of lead	0.414g	0.621g
c	∴ Mass of oxygen (a-b)	0.032g	0.098g

% by mass of lead in the two samples

Sample A

$$\frac{0.414}{0.446} \times \frac{100}{1}$$
$$= 92.8\%$$
$$\frac{0.621}{0.717} \times \frac{100}{1}$$
$$= 86.6\%$$

$$\text{ratio} = 92.8 : 86.6 = 1:1$$

% by mass of oxygen in the two samples

Sample A

$$\frac{0.032}{0.446} \times \frac{100}{1}$$
$$= 7\%$$
$$\frac{0.098}{0.717} \times \frac{100}{1}$$
$$= 13.7\%$$

$$\text{Ratio} = 7 : 13.7$$

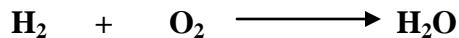
$$= 1 : 2$$

The mass of lead is fixed. Therefore the masses of oxygen, which combine with a fixed mass of lead is in the ratio of 1:2.

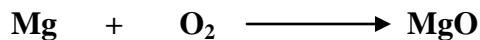
5.6 Chemical Equation

Chemical equation is a shorthand method of representation a chemical reaction.

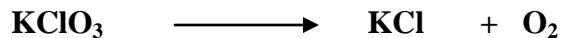
1. Hydrogen combines with oxygen to form water OR Hydrogen burns in air to form water. This statement can be represented thus:



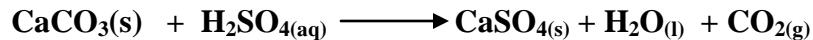
2. Magnesium burns in air to form magnesium oxide.



3. Potassium trioxochlorate (v) decomposes on heating to form potassium chloride and oxygen.



4. Calcium trioxocarbonate (iv) reacts with dilute tetraoxosulphate (vi) acid to form Calcium tetraoxosulphate (vi), water and carbon (iv) oxide.



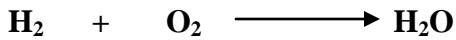
Writing Chemical Equation

1. In writing chemical equation, we represent an element by its chemical symbol and a compound by its chemical formula.

Example:

When oxygen reacts with hydrogen to form water, we write the equation thus:

Hydrogen combines with oxygen to form water



2. The reactants are written on the left hand side of the arrow while the products are written on the right.
3. The reactants and the products are linked by an arrow. The “+” sign on the left side of the arrow means ‘react with’, while the “+” sign on the right side means ‘and’.

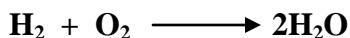
Balancing of Chemical Equation

The rules that must be maintained in balancing chemical equations are:

1. The total number of atoms of any element on the left hand side must be equal to the total number of atoms of that element on the right hand side.
2. Equations must always be balanced through the use of coefficients and not by changing the subscript numbers written in the formulae.
3. Common gases such as oxygen, hydrogen, nitrogen and chlorine are diatomic in their free state eg O₂, H₂, N₂ and Cl₂.

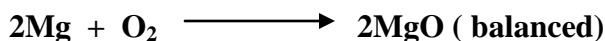


Count the number of hydrogen atoms on the left side and then on the right side. There are two atoms of hydrogen on both sides of the equation. So it is balanced, but there are two atoms of oxygen on the left and one atom on the right. Then, to balance it, we write 2 in front of H₂O.



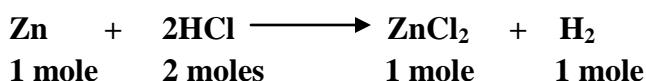
Now, the number of hydrogen is 2 on the left side of the equation and 4 on the right side. To balance hydrogen, write 2 in front of H₂ on the left side.

The equation is now balanced. ie.



5.7 Calculating Reacting Masses and Volumes using Chemical Equations.

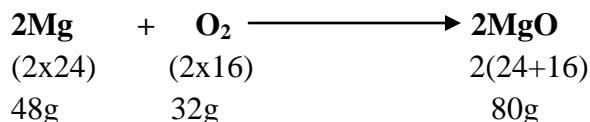
When zinc reacts with HCl_(aq), hydrogen gas is given off.



therefore 1 mole of Zn reacted with 2 moles of hydrochloric acid to form 1 mole of ZnCl₂ and 1 mole of hydrogen.

Example 1:

Calculate the mass of magnesium oxide formed when magnesium burns in oxygen to form the oxide. (Mg = 24, O = 16)



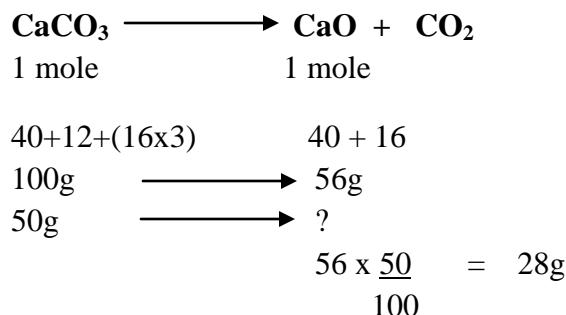
48g of Mg will react with 32g of O₂ to form 80g of MgO

Procedure:

1. Write the chemical equation using chemical symbols and formula.
2. Balance the chemical equations
3. Do the calculation.

Example 2:

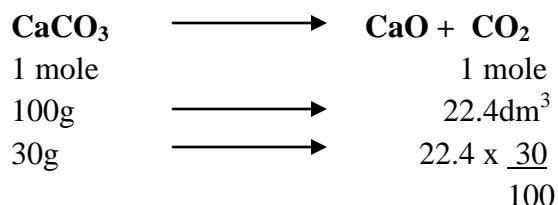
Calculate the mass of calcium oxide produced when 50g of calcium trioxocarbonate (iv) is heated. (Ca = 40, C = 12, O = 16)



Example 3

Calculate the volume of carbon (iv) oxide liberated when 30g of calcium trioxocarbonate (iv) is heated. (Ca = 40, C = 12, O = 16)

NOTE: The molar volume of any gas liberated at standard temperature and pressure (S.T.P) is 22.4 dm³.

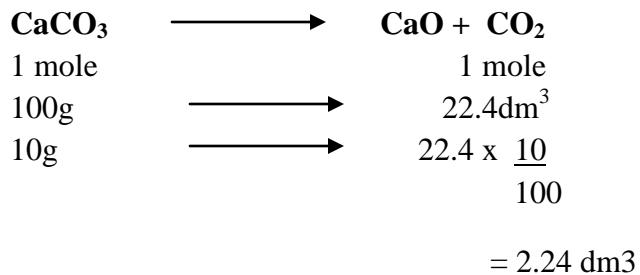


$$= 6.72 \text{ dm}^3 \text{ of CO}_2$$

Example 4:

Calculate the volume of carbon (iv) oxide liberated at 27°C and 750 mm Hg when 10g of calcium trioxocarbonate (iv) is heated.

(Ca = 40, C = 12, O = 16, Molar volume of a gas at S.T.P = 22.4 dm³)



convert the 2.24 dm³ at S. T. P. to the volume at 270°C and 750 mm Hg using the general gas equation. i.e.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$\text{Where } P_1 = \text{S.P} = 760 \text{ mm Hg}$$

$$V_1 = 2.24 \text{ dm}^3$$

$$T_1 = \text{S.T} = 273 \text{ K}$$

$$P_2 = 750 \text{ mm Hg}$$

$$V_2 = ?$$

$$T_2 = (27 + 273) \text{ K} + 300 \text{ K}$$

$$\therefore V_2 = P_1 V_1 T_1$$

$$P_2 T_2$$

$$= \frac{760 \times 2.24 \times 300}{750 \times 273} = 2.494 \text{ dm}^3$$

The volume of CO₂ liberated = 2.494 dm³

Exercise:

1. Write down in your exercise book, the four laws of chemical combination.
2. State the law of
 - a. Conservation of matter or the law of indestructibility of matter.
 - b. Definite proportion or constant composition.
3. 1.1g of an oxide of A contained 0.5g of A and 4g of another oxide of A contained 1.6g of A.
show these masses are in accordance with the law of multiple proportion.
1. a. What do you understand by the term “Chemical Equation”?
 b. Write two examples to show what a chemical equation is?
2. Calculate the volume of carbon (iv) oxide liberated when 20g of calcium trioxocarbonate (iv) is heated. (Ca = 40, C = 12, O = 16)
3. Calculate the mass of copper (ii) oxide produced when 63.5g of copper is completely burnt in excess oxygen. (Cu = 63.5, O = 16)
4. Calculate the percentage of calcium in calcium trioxocarbonate (iv). Formula = CaCO_3
(Ca = 40, C = 12, O = 16)

CHAPTER 6

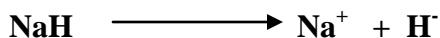
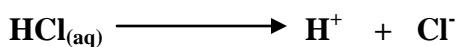
THE PERIODIC TABLE

Periodic table is a table of elements arranged in order of their increasing atomic numbers. There are about 109 known elements in the world. Modern periodic table is created by a man called Mendeleeff. The periodic table has horizontal rows called periods. A group is a vertical column in the periodic table containing elements with the same number of electrons in the outermost shell. Elements in the same group have similar chemical properties. The number of the group indicates the number of electrons in the outermost shell of the element in the group. In other words, all the elements in each group have the same valence electrons. Group 4 shows that element in this group have 4 electrons in their outermost shell. A group is a family of elements with similar chemical properties. There are groups 1, 2, 3, 4, 5, 6, 7, and 0.

GROUP PERIOD	1	2	3	4	5	6	7	0
1	1 H Hydrogen							2 He Helium
2	3 Li Lithium	4 Be Beryllium	5 B Boron	6 C Carbon	7 N Nitrogen	8 O Oxygen	9 F Fluorine	10 Ne Neon
3	11 Na Sodium	12 Mg Magnesium	13 Al Aluminum	14 Si Silicon	15 P Phosphorus	16 S Sulfur	17 Cl Chlorine	18 Ar Argon
4	19 K Potassium	20 Ca Calcium	31 Ga Gallium	32 Ge Germanium	33 As Arsenic	34 Se Selenium	35 Br Bromine	36 Kr Krypton
5	37 Rb Rubidium	38 Sr Strontium	49 In Indium	50 Sn Tin	51 Sb Antimony	52 Te Tellurium	53 I Iodine	54 Xe Xenon
6	55 Cs Cesium	56 Ba Barium	81 Tl Thallium	82 Pb Lead	83 Bi Bismuth	84 Po Polonium	85 At Astatine	86 Rn Radium
7	87 Fr Francium	88 Ra Radium						

Fig 6.1 Periodic table

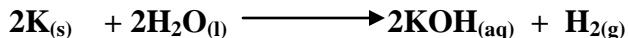
Hydrogen has a unique position in the periodic table because its chemistry resembles that of the alkali metals and halogens.



Hydrogen is sometimes placed under group 1 because it has one electron.

6.1 Group 1 Elements:

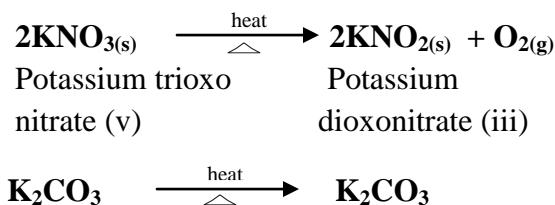
These are Lithium, Sodium, Potassium, Calcium, Francium. Group 1 elements are called the alkali metals because they react vigorously with cold water forming alkali and liberating hydrogen. Their oxides dissolve in water to form strong alkalis.



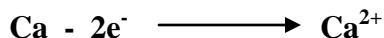
They have one electron in their outermost shell and exhibit valency of 1. Because they are very reactive and can react with air and water, they are stored under oil or paraffin. They have low melting point and low densities. They are good conductors of heat and electricity. They ionize by loss of their one valence electron.



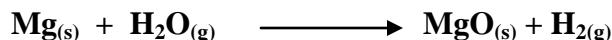
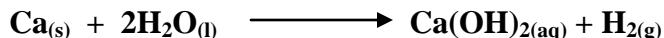
Their trioxonitrate (v) salts decompose with difficulty on heating. Sodium and potassium trioxocarbonate (iv) salts are the only soluble carbonates that do not decompose on heating.



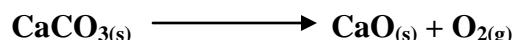
Group II Elements: are Beryllium, Magnesium, Calcium, strontium and barium. They are called the alkaline earth metals. They have two electrons in their outermost shell and ionize by the loss of the two valence electrons.



They react with steam to liberate hydrogen as follows:



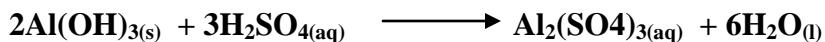
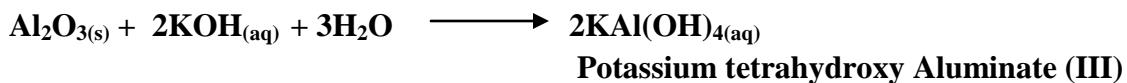
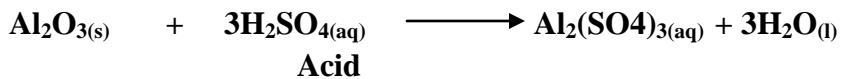
Their trioxonitrate (v) salts and trioxocarbonate (iv) salts decompose readily on heating as follows”



Group III Elements are boron, aluminum, gallium etc. The only most common element of these is aluminum. It has three electrons in its outermost shell and ionizes by the loss of the 3 valence electrons.



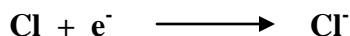
Al and its oxide cannot react with water and so does not form an alkali. Al_2O_3 and $\text{Al}(\text{OH})_3$ exhibit both acidic and basic character. They are said to be amphoteric in nature.



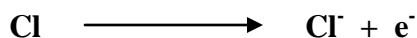
On heating, aluminum trioxonitrate (v) decomposes.



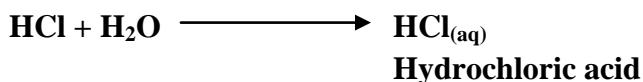
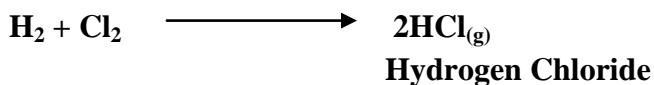
Group VII Elements: are called the halogens. They are fluorine, chlorine, bromine and iodine. They are non-metals. They have seven electrons in their outermost shell and ionize by gain of one electron.



Or



Halogen exists as diatomic molecules eg Fluorine (F_2), Chlorine (Cl_2), Bromine (Br_2) and Iodine (I_2). They are very reactive non-metals. Halogens have similar chemical and physical properties. Their hydrides which are gases at room temperature, dissolve in water to form acids, eg.



Group O Elements are the inert gases or noble gases or rare gases. They are: Helium, Argon, Neon, Krypton, Xenon and Radon. They are inert or highly unreactive because they have their outer shell completely filled with electrons.

The transition elements are all metals. They are called transition metals because they use electrons from both the inner (penultimate) and outer shells for chemical combination. They have variable valencies and are therefore used as catalysts e.g. Iron (ii) and Iron (iii).

Periodicity is the regular patterns or trends in which the physical and chemical properties of elements occur when the elements are arranged according to their atomic numbers in the periodic table.

Periodic Law states that the properties of element are the periodic functions of their atomic number.

6.2 The Noble Gases

The noble gases are also called the rare gases. They are chemically inert, or very unreactive because, they have duplet or octet arrangement. They constitute about 1% by volume of atmospheric air. The symbols and electronic configuration of three rare gases are as follows:

Rare gases	Symbol	Atomic No.	Electronic configuration
Helium	He	2	1 s ²
Neon	Ne	10	1 s ² 2 s ² 2 p ⁶
Argon	Ar	18	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s.

Fig 6.1 Increase in energy level of atomic orbitals

Other rare gases are, krypton, xenon, and radon.

Argon is the most abundant of all the rare gases in the atmosphere. With the exception of helium, the outer quantum shell of other rare gases are filled with eight electrons.

Uses of Noble Gases

1. An invisible ray (radiation) produced by radon is used in curing some forms of cancer.
2. Helium is used in observation balloons because of its lightness and non-combustibility.
3. Neon, argon and helium can be made to produce coloured light when they are subjected to low pressure and a high voltage current is passed through them.

6.3 Arrangement of Electrons in the Main and Sub-energy Levels

In an atom main shells are lettered K, L, M, N, etc. it is lettered outwards from the nucleus and they correspond to main energy level 1, 2, 3, 4, etc. The maximum number of electrons in a main shell or in each main energy level is given by $2 n^2$, where n stands for the number of the main energy level.

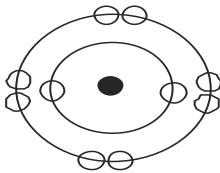


Fig 6.2 Arrangement of electrons in the main shells.

Thus for the main energy level (shell) 1, the maximum number of electrons it can accommodate is $2n^2$ ie

$$\begin{aligned} 2 \times n \times n \\ 2 \times 1 \times 1 &= 2 \text{ electrons} \end{aligned}$$

For the main shell 2, the maximum number of electrons it can carry is $2n^2$ ie

$$\begin{aligned} 2 \times n \times n \\ 2 \times 2 \times 2 &= 8 \text{ electrons etc.} \end{aligned}$$

Each main shell is made up of one or more sub-shells or orbitals. An orbital is the space around the nucleus where there is the probability of finding electrons. There are s, p, d, and f types of orbitals. There are three p-orbitals which are $2p_x$, $2p_y$ and $2p_z$ orbitals. The energy of electrons in a given sub-shell is different from the energy of electrons in other sub-shells. The sub-shell with the lowest energy is called the S-sub-shell. The next higher energy sub-shell is called the p-orbital. The next higher energy sub-shell is called the d-sub-shell while the next higher energy sub-shell or orbital is called the f-orbital.

There are limited number of electrons which can be accommodated in each sub-shell. For s-orbital, the maximum number of electrons it can accommodate is two (2). The p-orbital can carry a maximum of six (6) electrons: the d-orbital can contain a maximum of ten (10) electrons while the f-orbital can accommodate a maximum of fourteen (14) electrons.

Orbitals with the same energy are called degenerate orbitals. The p-orbital consists of three degenerate orbitals, called P_x , P_y and P_z . the d-orbital consists of five degenerate orbitals while the f-orbital consists of seven degenerate orbitals. The arrangement of electrons in the sub-shells of some elements are shown below:

Main shell (n) →			K 1	L 2	M 3
Subshell →			s	s, p	s, p, d
Element	Symbol	Atomic Number			
Sodium	Na	11	2	2 6	1
Magnesium	Mg	12	2	2 6	2
Aluminum	Al	13	2	2 6	2 1
Silicon	Si	14	2	2 6	2 2
Phosphorus	P	15	2	2 6	2 3
Sulphur	S	16	2	2 6	2 4
Chlorine	Cl	17	2	2 6	2 5
Argon	Ar	18	2	2 6	2 6
Potassium	K	19	2	2 6	2 6 1
Calcium	Ca	20	2	2 6	2 6 2

6.4 Electronic Configuration

Electronic configuration of an atom is the arrangement of electrons in the orbitals or sub-shells of the atom. The physical and chemical properties of an atom depend on the arrangement of its electrons. Orbitals with lower energy are filled with appropriate number of electrons before orbitals with higher energy. For example, the 2s-orbital is filled completely before the 2 p-orbitals. The arrangement of the orbitals according to their increasing energy is 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, etc.

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s,
Increase in energy level of atomic orbitals →

Fig 6.3 Energy levels of atomic orbitals or subshells

1s is the orbital that has the least energy while 7s is the one that has the highest energy.

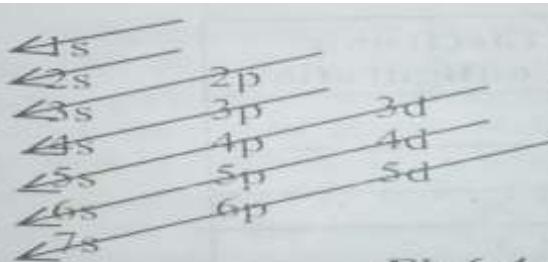


Fig 6.4 order of filling orbitals:

Electronic Configuration of the first 20-elements:

Electrons in an atom occupy orbital in order of increasing energy. The orbital or shell nearest to the nucleus is the one with the lowest energy. Electron can move from one energy level to another. For electron to move from one energy level to another with higher energy, the electron must absorb energy from the surroundings. The amount of energy to be absorbed must be equal to the difference in energy between the two energy levels.

In the same way for an electron to move from an orbital at a higher energy level, to an orbital at a lower energy level, the electron must emit an amount of energy equal to the energy difference between the two shells. Atoms of each element have characteristic electronic configuration peculiar to that element. Writing of electronic configuration of atoms of some elements in their ground state can easily be done by filling in electrons to the orbitals starting with the lowest energy subshell. The ground state of an element is its most stable state. The electronic configuration of an atom of an element with the atomic number 7 is $1s^2\ 2s^2\ 2p^3$. This implies that the subshells are listed in order of increasing energy and the numbers 2 and 3 are added to show the number of electrons on each orbital. An element is said to be in its ground state or most stable state when its electrons are occupying orbitals with the lowest available energy. The ground state electronic configuration of the first twenty element are as follows:

Element	Symbol	Atomic Number	Electronic Configuration
Hydrogen	H	1	$1s^1$
Helium	He	2	$1s^2$
Lithium	Li	3	$1s^2\ 2s^1$
Beryllium	Be	4	$1s^2\ 2s^2$
Boron	B	5	$1s^2\ 2s^2\ 2p^1$
Carbon	C	6	$1s^2\ 2s^2\ 2p^2$
Nitrogen	N	7	$1s^2\ 2s^2\ 2p^3$
Oxygen	O	8	$1s^2\ 2s^2\ 2p^4$
Florine	F	9	$1s^2\ 2s^2\ 2p^5$
Neon	Ne	10	$1s^2\ 2s^2\ 2p^6$
Sodium	Na	11	$1s^2\ 2s^2\ 2p^6\ 3s^1$
Magnesium	Mg	12	$1s^2\ 2s^2\ 2p^6\ 3s^2$
Aluminum	Al	13	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
Silicon	Si	14	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$
Phosphorus	P	15	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$
Sulphur	S	16	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
Chlorine	Cl	17	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
Argons	Ar	18	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
Potassium	K	19	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
Calcium	Ca	20	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

Energy Level	Number of Orbitals	Maximum Number of electrons
n=1	$1^2 = 1$	$2 \times 1 = 2$
n=2	$2^2 = 4$	$2 \times 4 = 8$
n=3	$3^2 = 9$	$2 \times 9 = 18$
n=4	$4^2 = 16$	$2 \times 16 = 32$
n=5	$5^2 = 25$	$2 \times 25 = 50$
n=6	$6^2 = 36$	$2 \times 36 = 72$
n=7	$7^2 = 49$	$2 \times 49 = 98$
n=8	$8^2 = 64$	$2 \times 64 = 128$

Fig. 6.7

In writing the electronic configuration, the numbers before the letters denote the main energy levels while the letters denote the energy sub-shells or orbitals. The figure on top of the letter show the number of electrons in each orbital.

6.5 Relative Atomic Mass Based on Carbon 12:

The mass of an atom is an extremely very small and inconvenient fraction. It is therefore preferable to express atomic masses by comparing them with the mass of a standard atom. Nowadays, the atom chosen as the standard for comparison is carbon-12. Carbon 12 is the most common isotope of carbon, which has a mass of 12 units. Relative atomic mass of an element is the number of times one of the element is heavier than one-twelfth ($^{1/12}$) of one atom of carbon 12 isotope. Relative atomic mass of an element has no unit because it is a ratio but atomic mass of an element is expressed in gram.

$$\text{R. At. Mass} = \frac{\text{Mass of one atom of an element}}{\text{Mass of } ^{1/12} \text{ of one atom of carbon 12.}}$$

The relative atomic mass of the first twenty elements are shown below:

NAME	SYMBOL	RELATIVE ATOMIC MASS
Hydrogen	H	1.008
Helium	He	4.003
Lithium	Li	6.94
Beryllium	Be	9.012
Boron	B	10.81
Carbon	C	12.01
Nitrogen	N	14.01
Oxygen	O	16.00
Florine	F	19.00
Neon	Ne	20.18
Sodium	Na	22.99
Magnesium	Mg	24.31
Aluminum	Al	26.98
Silicon	Si	28.09
Phosphorus	P	30.97
Sulphur	S	32.06
Chlorine	Cl	35.45
Argons	Ar	39.95
Potassium	K	39.10
Calcium	Ca	40.08

From the definition of relative atomic mass, one might think that all elements would have whole number masses. In fact relative atomic masses are rarely found to be exactly whole numbers. This is because most elements have two or more naturally occurring isotopes, and the relative atomic mass takes into account the abundant of each isotope. For example, chlorine has approximately 75% of an isotope with mass number 35, and 25% of an isotope with mass number 37. therefore, the relative atomic mass of chlorine is found to be 35.5.

Solution:

$$\begin{array}{ccc}
 75\% & & 25\% \\
 75 & : & 25 \\
 \text{Ratio} = & 3 & : 1 \\
 \text{Total Ratio} = & 3 + 1 = 4 \\
 \text{Isotopes} & 35 & 37 \\
 & (35 \times 3) + (37 \times 1) \\
 & 105 + 37 = 142 \\
 \therefore \text{Average or the relative atomic mass} & = \frac{142}{4} & = 35.5
 \end{array}$$

Question 1

An element X has two isotopes $^{20}_{10}\text{X}$, $^{22}_{10}\text{X}$, present in the ratio of 1:3, what would their relative atomic mass be?

Answer

The two isotopes = X and X

Their ratio = 1 : 3

Total ratio = 1 + 3 = 4

∴ Relative atomic mass of X is $\frac{(\text{mass of 1}^{\text{st}} \text{ isotope} \times 1) + (\text{mass of 2}^{\text{nd}} \text{ isotope} \times 3)}{\text{total ratio}}$

$$= \frac{(20 \times 1) + (22 \times 3)}{4} = \frac{20 + 66}{4} = 21.5$$

Question 2:

The mean atomic mass of chlorine is 35.5. Chlorine has two isotopes of mass 35 and 37. What is the percentage composition of the isotope of mass number 35?

Answer

Let the number of atoms be 100

“ “ “ “ ^{35}Cl “ A atoms

“ “ “ “ ^{37}Cl “ 100 -A atoms

A atoms have 35 units each and (10-A) atoms have 37 units each.

∴ Total mass of atoms = 35.5

Total number of atoms

$\frac{(.35 \times A) + (100-A) \times 37}{100 \text{ atoms}} = 35.5$

$\frac{35A + 3700 - 37A}{100 \text{ atoms}} = 35.5$

$(35-37)A + 3700 = 3550 \text{ atoms}$

$-2A = -150$

$$\begin{aligned}2A &= 150 \\ A &= \frac{150}{2} = 75 \text{ atoms}\end{aligned}$$

In %, we have $\frac{75 \text{ atoms}}{100 \text{ atoms}} \times \frac{100}{1} = 75\%$

Question 3

Carbon has two existing (main) isotopes $^{13}_6\text{C}$ and $^{12}_6\text{C}$, with abundances of 1.11% and 98.89% respectively. Calculate the relative atomic mass of carbon.

Answer

Relative atomic mass = R. a. m.

$$\begin{aligned}\text{R. a. m.} &= \text{mass due to C} + \text{mass due to C} \\ &= \frac{13 \times 1.11}{100} + \frac{12 \times 98.89}{100} \\ &= 12.01\end{aligned}$$

6.6 Relative Molecular Mass:

Relative molecular mass of a compound or an element is the number of times one molecule of the compound or element is heavier than ^{12}C of the mass of one atom of carbon 12 isotope.

$$\text{R.M Mass} = \frac{\text{Mass of one molecule of element or compound}}{\text{Mass of } 1/12 \text{ of one atom of carbon-12}}$$

The relative molecular mass of a substance is obtained by adding together the relative atomic masses of all the atoms of the elements present in one molecule of the substance. Relative molecular mass has no unit.

Example 1: calculate the relative molecular mass of water (H_2O).

$$(\text{H} = 1, \text{O} = 16)$$

Solution: Relative molecular mass of H_2O

$$\begin{aligned}&= (1 \times 2) + (16) \\ &\quad 2 + 16 = 18\end{aligned}$$

Example 2: Calculate the relative molecular mass of Sodium chloride. (Na = 23, Cl = 35.5)

Solution: Relative molecular mass of NaCl = $23 + 35.5 = 58.5$

6.7 Energy Levels in Atomic System

The Four Quantum Numbers:

These are represented with the letters n, l, m and s. The four quantum numbers are:

- i. **The principal quantum number, 'n':** gives the main or major energy levels. 'n' can have any positive value other than zero. ie $n = 1, 2, 3, 4, 5$ etc.
- ii. **The secondary or azimuthal quantum number 'l':** Within the main energy level, there are sub energy levels (sub shells). The total number of sub-shells is equal to the principal quantum number. The azimuthal quantum number 'l' has values from 0 to $(n=I)$. If the value of 'n' is 1 (first energy level), l has only one value 0. If $n = 2$ (second energy level), l has two values, 0 and 1. Since l has one value when $n = 1$ and two values when $n = 2$, it could be said that there is one sub shell in the first shell (k) and 2 sub shells in the second shell (l). The principal quantum number therefore gives the main shell or energy level while the azimuthal quantum number gives the sub shells or sub energy levels.
- iii. **Magnetic quantum number, 'm':** is the additional lines that appeared in the spectrum when a strong magnetic field was applied to the source of spectrum. For each azimuthal quantum number 'l' there are $(2l + 1)$ values of m.
- iv. **The spin quantum number, 's':** An electron can spin either clock wise or anti clockwise. The spin quantum number 'S' therefore has two values which are equal and opposite. The value of s is + and - .

An atomic orbital is the region in space within which electron can be found. These atomic orbitals have different shapes, sizes and orientations.

An atom:

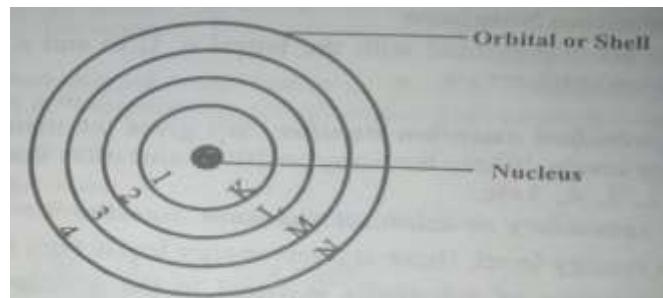


Fig. 6.5

SHAPE OF S AND P ORBITALS

S - orbital

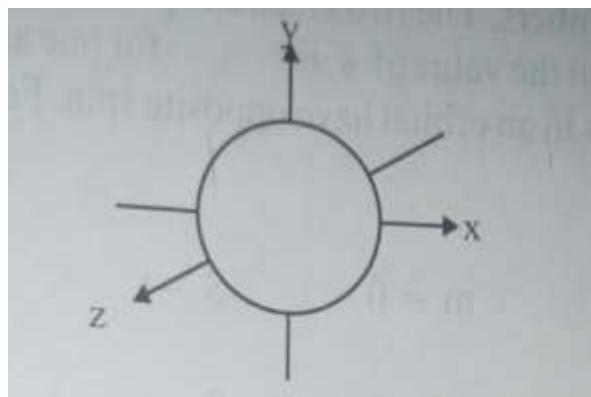


Fig. 6.6

Energy Level	Number of Orbitals	Maximum Number of electrons
n=1	$1^2 = 1$	$2 \times 1 = 2$
n=2	$2^2 = 4$	$2 \times 4 = 8$
n=3	$3^2 = 9$	$2 \times 9 = 18$
n=4	$4^2 = 16$	$2 \times 16 = 32$
n=5	$5^2 = 25$	$2 \times 25 = 50$
n=6	$6^2 = 36$	$2 \times 36 = 72$
n=7	$7^2 = 49$	$2 \times 49 = 98$
n=8	$8^2 = 64$	$2 \times 64 = 128$

Fig. 6.7

- iii. An orbital can carry a maximum of only two electrons and orbitals of the same type within a principal quantum number possess equal energies. For instance the three 2p - orbitals in the same L - shell all have equal energies. The same type of orbitals in different quantum numbers have different energies. For example, the 2s - orbital in the L-shell, have higher energy than the 1s - orbital in the K-shell.
- iv. Within a principal quantum number, different sub-shells or orbitals do not have equal energies. For instance 2s orbital in L-shell does not have equal energy with the 2p orbitals in the same L-shell. The order of the increasing energy content of the orbitals is as follows:

$$1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d \dots \text{etc.}$$

Pauli's Exclusion Principle: states that no two electrons have identical four quantum numbers. The two electrons present in an orbital have the same n, l and m but the value of s is + for one

and - for the other. The two electrons in an orbital have opposite spin. For two electrons in s orbital

$$n = 1 \ l = 0 \quad m = 0 \quad S = +$$

$$n = 1 \ l = 0 \quad m = 0 \quad S = -$$

Hund's rule: states that the electrons fill the orbitals of an atom in such a way that they first occupy singly orbitals of equivalent energies before they start being paired up in them.

Filling of electrons in orbitals

Based on the rules laid down by Pauli and Hund, electrons are filled in the orbitals of the first ten elements in the periodic table.

Element	K-shell		L-Shell				M - Shell					
	1	1s	2	2s	2px	2py	2pz	3	3s	3px	3py	3pz
1 H		↑										
2 He		↑↓										
3 Li		↑↓		↑								
4 Be		↑↓		↑↓								
5 B		↑↓		↑↓		↑						
6 C		↑↓		↑↓		↑						
7 N		↑↓		↑↓		↑						
8 O		↑↓		↑↓		↑↓		↑				
9 F		↑↓		↑↓		↑↓		↑↓				
10 Ne		↑↓		↑↓		↑↓		↑↓		↑↓		

Electronic configuration: is the arrangement of electrons in the orbitals.

The electronic configuration of

$$C = 1s^2 \quad 2s^2 \quad 2p^2$$

$$N = 1s^2 \quad 2s^2 \quad 2p^3$$

$$F = 1s^2 \quad 2s^2 \quad 2p^5$$

$$Ne = 1s^2 \quad 2s^2 \quad 2p^6$$

$$Cl = 1s^2 \quad 2s^2 \quad 2p^6 \quad 3s^2 \quad 3p^5$$

$$K = 1s^2 \quad 2s^2 \quad 2p^6 \quad 3s^2 \quad 3p^6 \quad 4s^1$$

$$Ca = 1s^2 \quad 2s^2 \quad 2p^6 \quad 3s^2 \quad 3p^6 \quad 4s^2$$

Exercises:

1. Define the following:
 - a. Periodic table
 - b. Periodicity
2. State periodic law
3. Write two examples of the following
 - a. Group I elements
 - b. Group II elements
 - c. Group III elements
4. (a) Define electronic configuration
(b) Write the electronic configuration of Na, Al, P and O. Their atomic numbers are 11, 13, 15 and 8 respectively.
5. (a) What do you understand “Relative atomic mass” to mean? (b) Has it any unit? Why?
6. An element, Y, has two isotopes Y_1 , Y_2 , present in the ratio of 3:1. What would the relative atomic mass of Y be?
7. The average or relative atomic mass of an element X, is 21.5. X has two isotopes of mass 20 and 22. What is the percentage composition of the isotope of mass 20?
8. What are the four quantum numbers?
9. (a) What is an atomic orbital?
(b) Draw the S - orbital
10. Which orbital has the least energy?
11. (a) What is electronic configuration?
(b) Write the electronic configuration of these elements:

CHAPTER 7

CHEMICAL COMBINATION OR BONDING

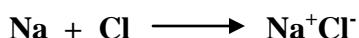
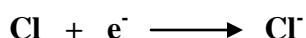
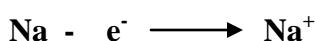
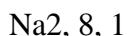
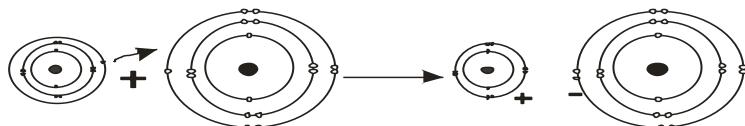
Atoms of elements enter into chemical combination to form stable substances (molecules or compounds). The reason for their entering into the combination is for them to attain the inert octet or duplet structure possessed by the noble gases. The types of chemical bonding that held the chemical substances together are:

- a. Electrovalent or ionic bonding.
- b. Covalent bonding
- c. Co-ordinate covalent or dative bonding
- d. Metallic bonding
- e. Hydrogen bonding
- f. Van der Waal forces

7.1 Electrovalent Bonding:

This is also known as ionic bonding. It is a bonding that occurs between metal and non-metal. It involves the transfer of valence electron or electrons from metal to non metal. The metal loses its valence electron while the non-metal accepts it and the resultant ions are attracted together by a strong force called electrostatic force. The metal acquires positive charge or charges because it contains excess proton or protons in its nucleus while the non-metal acquires negative charge or charges because it has gained extra electron or electrons.

Thus, sodium combines with chlorine to form sodium chloride



Sodium combines with oxygen to form sodium oxide

$$\begin{array}{llll} \text{Na} & = & 11 & 2, 8, 1 \\ \text{O} & = & 8 & 2, 6 \end{array}$$

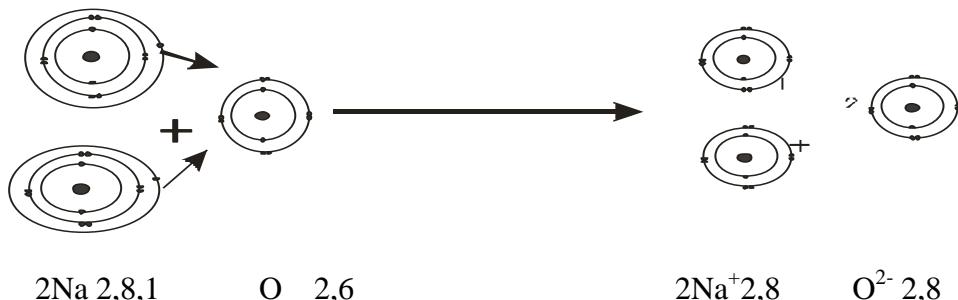
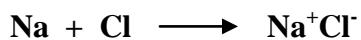
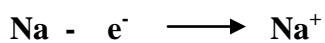
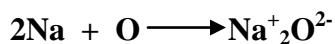
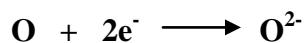
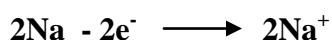


Fig. 7.2



Calcium combines with oxygen to form calcium oxide

$$\begin{array}{llll} \text{Ca} & = & 20 & 2, 8, 8, 1 \\ \text{O} & = & 8 & 2, 6 \end{array}$$

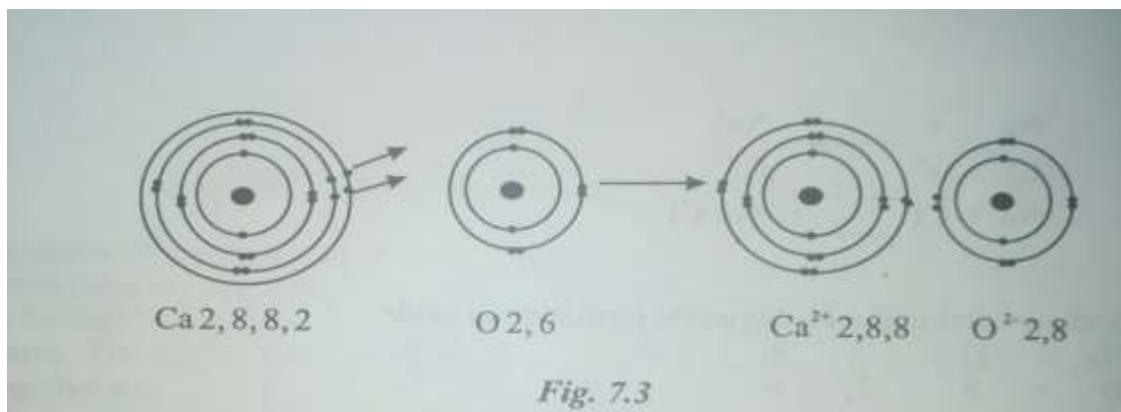
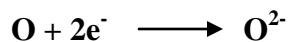
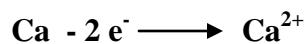


Fig. 7.3



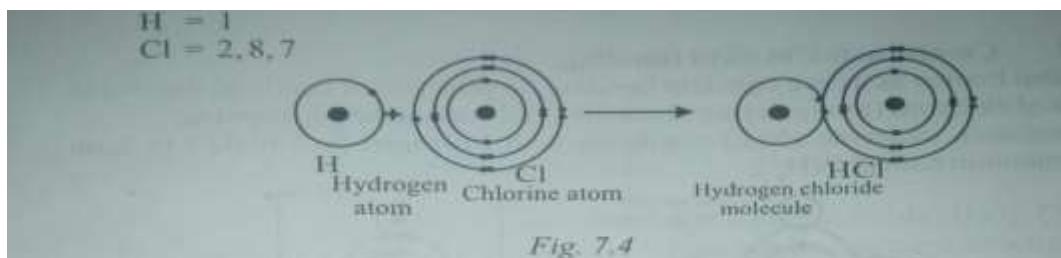
Characteristics of electrovalent compounds.

1. Electrovalent compounds consist of aggregates of ions (charged particles).
2. They have high melting and boiling points.
3. They are soluble in water and other polar solvents.
4. They are strong electrolytes
5. They are insoluble in organic solvents because they are inorganic compounds.
6. They are good conductors of heat and electricity.

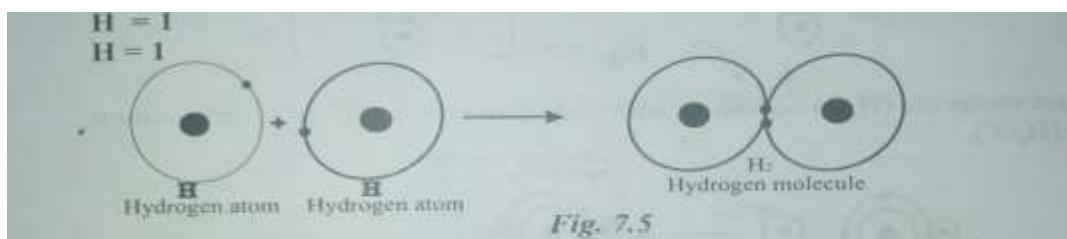
7.2 Covalent Bonding:

In this type of bonding, there is sharing of a pair of electrons. Each reacting atom contributes an electron to be shared by the two atoms, so that each has the stable duplet or octet structure of rare gases. This is the type of bonding that occurs between the non-metals. Diatomic molecules of elements are formed by covalent combination. Eg HCl, H₂ and Cl₂ molecules.

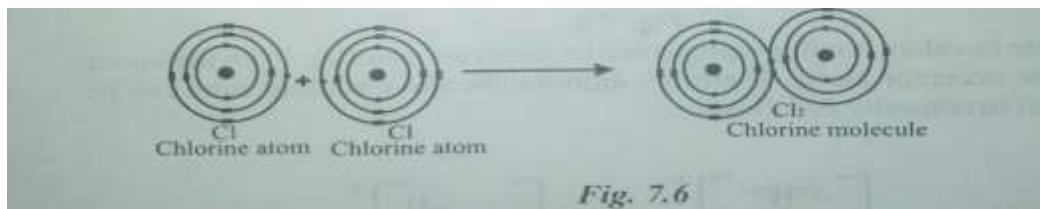
Hydrogen combines with chlorine to form hydrogen chloride gas.



Hydrogen atom combines with another hydrogen atom to form hydrogen molecule.



Chlorine atom combines with another chlorine atom to form chlorine molecule.



Characteristics of covalent compounds.

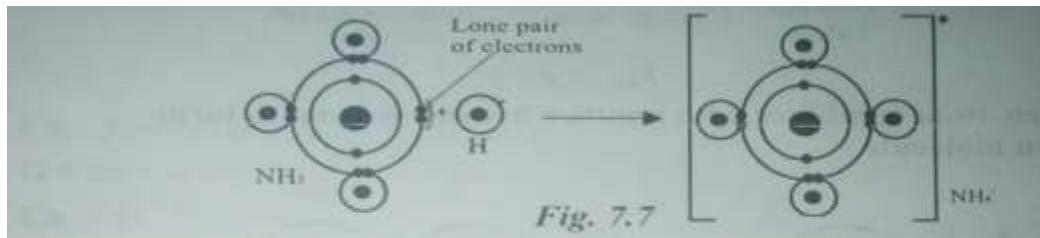
1. Covalent compounds consist of molecules. They are referred to as molecular compounds.
2. They have low melting and boiling points.
3. They are insoluble in water and other polar solvents.
4. They are non-electrolyte
5. They are usually soluble in organic solvents such as benzen and ethanol
6. They are non conductors of heat and electricity.

7.3 Co-ordinate Covalent Bonding:

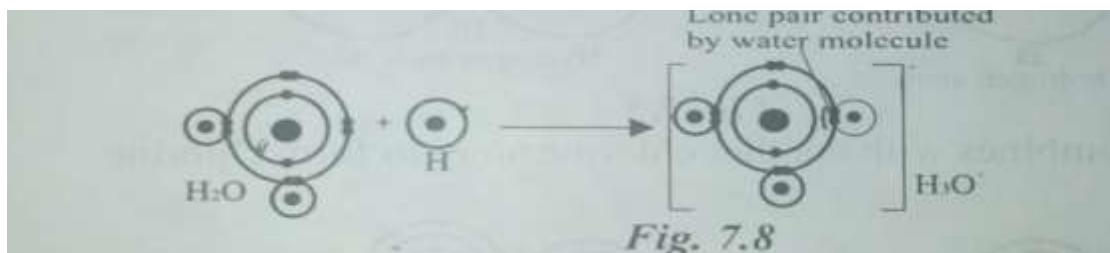
Is also known as dative covalent bonding. In this type of bonding, the shared pair of electrons (lone pair) are donated by only one of the participants.

For dative covalent bonding to occur, one of the reacting substances must posses a lone pair, Ammonia and water molecules possess lone pairs which enable them to form dative covalent bond with hydrogen proton (H^+).

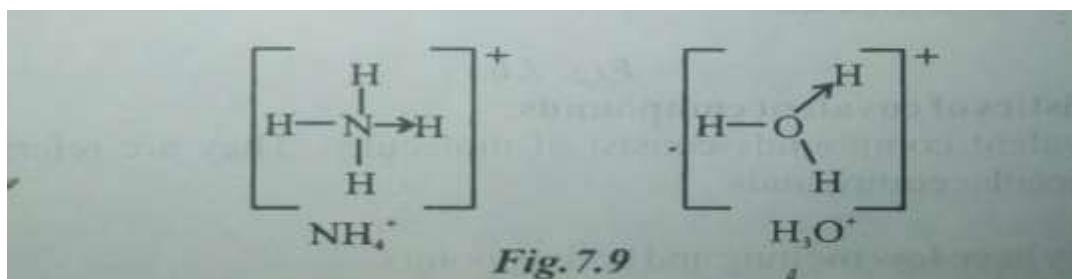
Ammonia molecule (NH_3) combines with hydrogen proton (H^+) to form ammonium radical (NH_4^+).



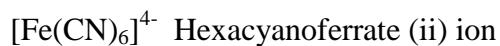
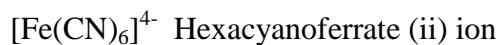
Water molecule (H_2O) combines with hydrogen proton (H^+) to form oxonium ion (H_3O^+).



Co-ordinate covalent bond is represented by an arrow pointing from the donor atom to the acceptor atom. Therefore ammonium and oxonium ions can be represented structurally as follows:



These complexes are held together by co-ordinate covalent bonding:



The substances that act as electron donor are water molecule (H_2O), Cyanide ion (CN^-), and ammonia molecule (NH_3). They can be represent structurally as follows:

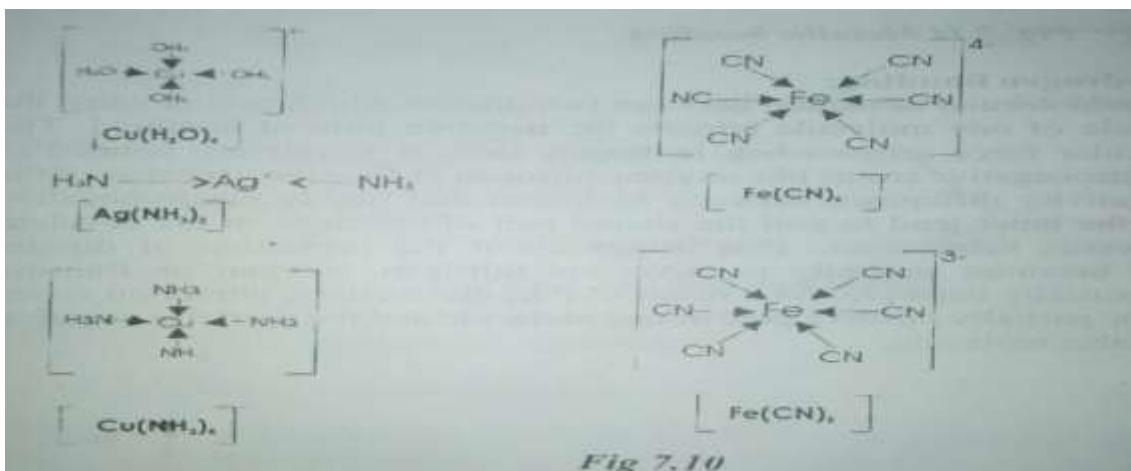


Fig. 7.10

7.4 Metallic Bonding:

It is an attractive force that held metal atoms together in crystal lattice. In a crystal lattice, the valence electrons of metal tend to separate from its orbital and move at random forming an electron cloud. The metallic ions formed tend to repel one another but are held together by the attractive force which exists between the electron cloud and the positively charged metallic ions. This force of attraction is referred to as metallic bond. This metallic bond increases with an increase in the number electron cloud and the positively charged metallic ions. This force of attraction is referred to as metallic bond. This metallic bond increases with an increase in the number of valence electrons for metals in a particular group. For instance, the decreasing order of metallic bond in the crystal lattice of sodium, magnesium and aluminum is as follows:

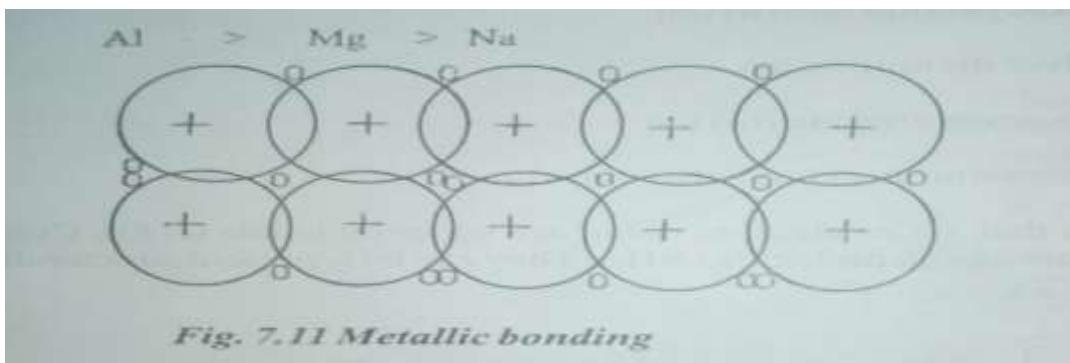
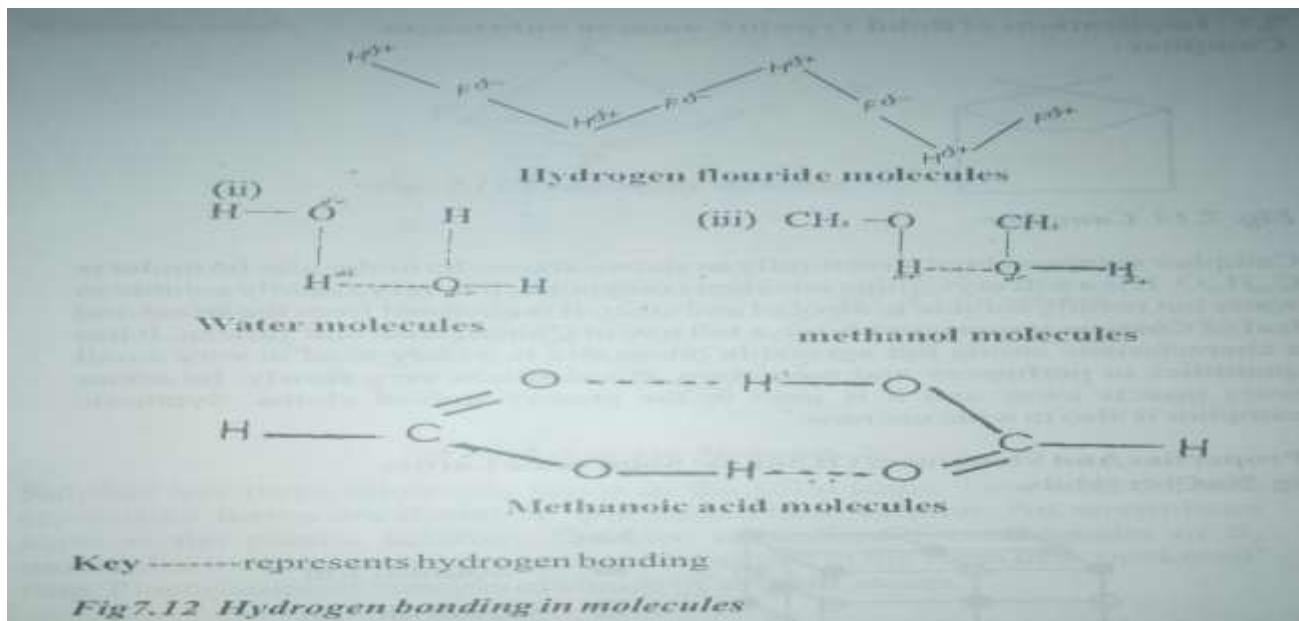


Fig. 7.11 Metallic bonding

7.5 Hydrogen Bonding:

It is an intermolecular attraction between two dipoles which occurs when the positive pole of one molecule attracts the negative pole of another. The intermolecular force arises when hydrogen atom is covalently bonded to highly electronegative atoms like oxygen, nitrogen or fluorine. Because of the electronegativity difference between hydrogen and highly electronegative elements, the latter tend to pull the shared pair of electrons in the covalent bonds towards themselves. This brings about the formation of dipole. Hydrogen becomes partially positive and nitrogen, oxygen or fluorine becomes partially negative. As a result of this, electrostatic

attraction arises between the partially positive pole of one molecule and the partially negative pole of another molecule.



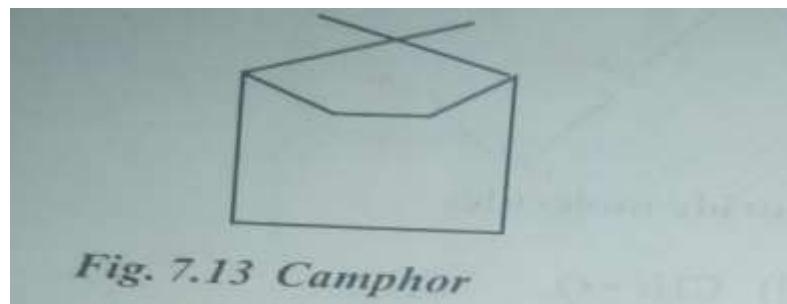
Hydrogen bonding causes the trans association of molecules eg Hydrogen fluoride (HF) can exist as H_2F_2 (a dimer) or H_3F_3 (a trimer).

7.6 Vander Waals Forces:

It is first studied and described by a Dutchman named J. D. Van der Waals. Hence the force is given the above name. It is a weak force that exists between discrete molecules. It is very important in the formation of molecular lattices in iodine and naphthalene crystals. It is also important in the liquefaction of gases.

7.7 Implications of Bond Type in Common Substances

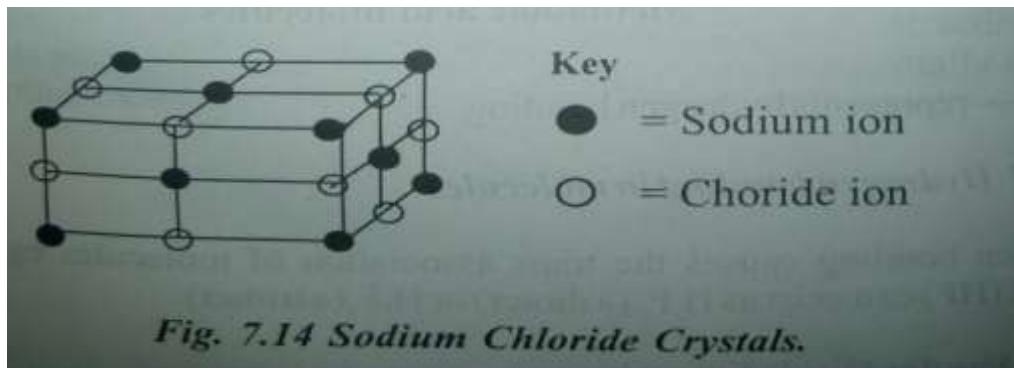
Camphor:



Camphor is represented structurally as shown above. Its molecular formula is $\text{C}_{10}\text{H}_{16}\text{O}$. It is a soft crystalline covalent compound. It is very slightly soluble in water but readily soluble in alcohol and ether. It is obtained from the wood and leaf of *Cinnamomum camphora*, a tall tree of

China, Japan and Taiwan. It has a characteristic strong but agreeable odour and is widely used in very small quantities in perfumery and medicines. It volatilizes very slowly. Its odour scars insects away and it is used in the preservation of cloths. Synthetic camphor is also in wide use now.

Properties and Structures Of Solids: Solid Ionic Lattice eg. NaCl crystals.



Properties of NaCl

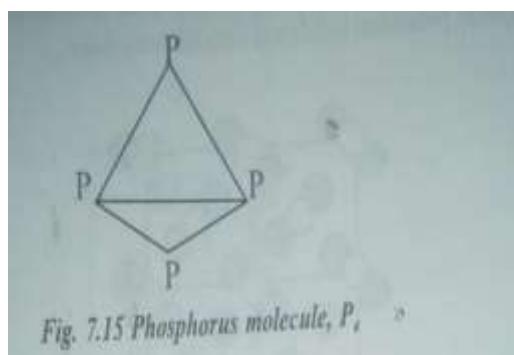
Sodium chloride exists as aggregate of ions.

It is soluble in inorganic solvents such as water.

It is a strong electrolyte either in an aqueous solution or in a fused or molten state.

It is a high melting point crystalline solid. Silver bromide, AgBr crystallises in the same way as NaCl.

Molecular solids



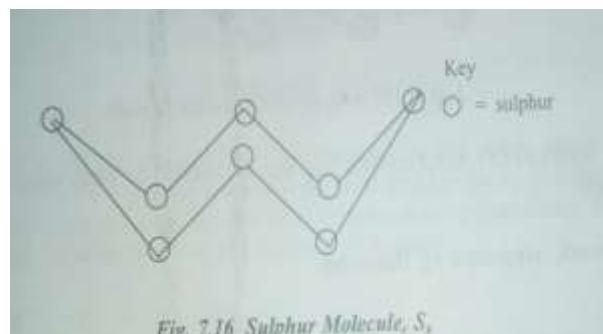


Fig. 7.16 Sulphur Molecule, S_8

Sulphur has three allotropic forms in the solid state. Two well-defined crystalline forms are rhombic and monoclinic sulphur. An amorphous form is the plastic sulphur. The two crystalline forms consist of S_8 molecules in which sulphur atoms are arranged in the form of a ‘puckered’ ring. Plastic sulphur consists of chains of sulphur atoms.

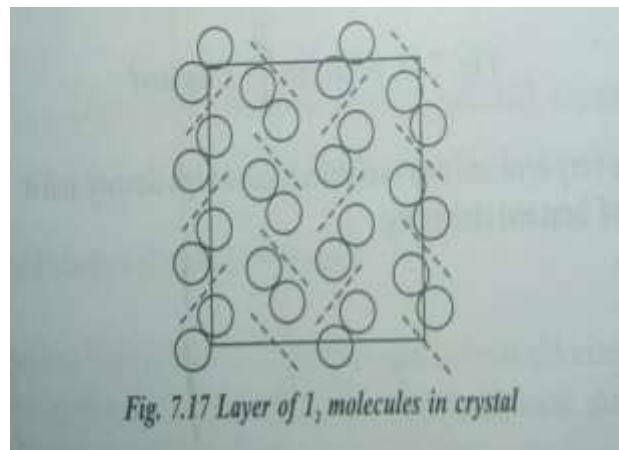


Fig. 7.17 Layer of I₂ molecules in crystal

Properties of iodine:

Iodine is a black crystalline solid with a slight metallic lustra. At atmospheric pressure, it sublimes producing violet vapour. It is sparingly soluble in water but readily soluble in alkanol, ether, carbon disulphide, potassium iodide and tetrachloromethane.

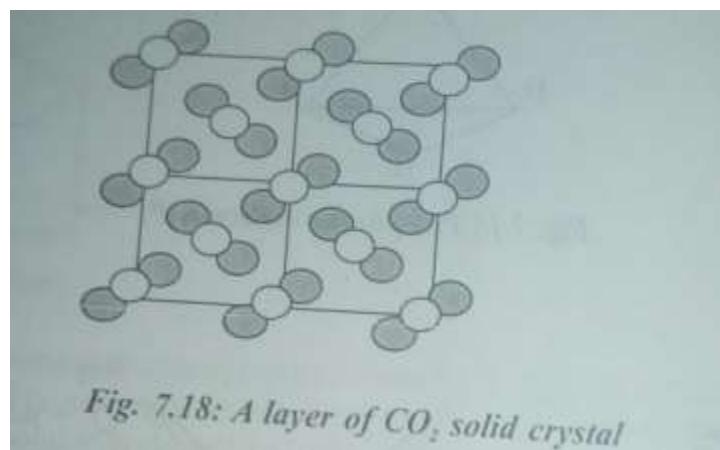
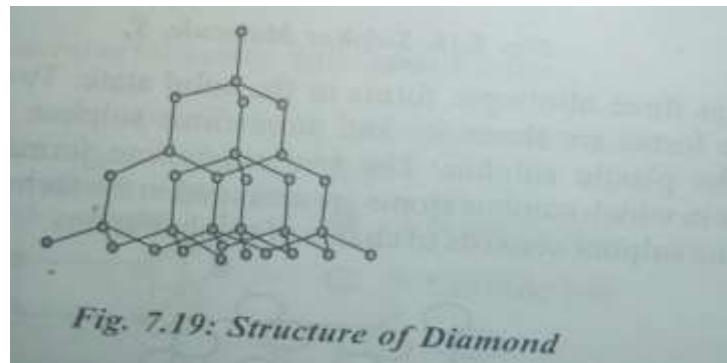


Fig. 7.18: A layer of CO₂ solid crystal

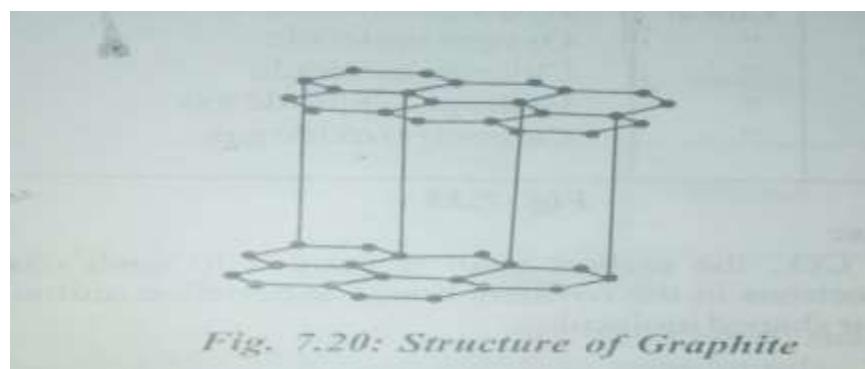
The layers of CO₂ solid crystals (dry ice) are bonded by weak van der Waal forces).

Network structure eg. Diamond

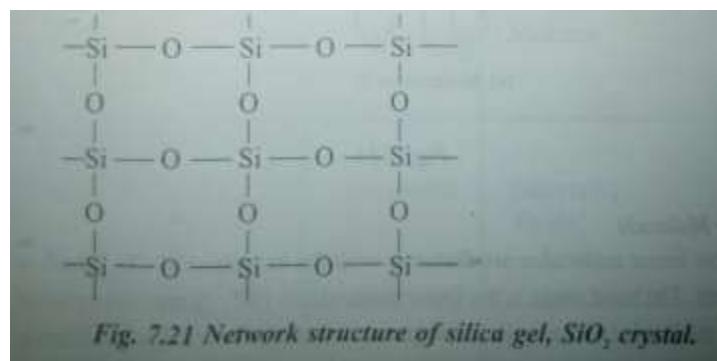


Diamond is very hard and non - volatile. It has a high density and it is a poor conductor of heat and electricity.

Layer lattice eg. Graphite



Graphite is very soft and resistant to chemical attack. It is shiny, black and opaque. Graphite is a good conductor of heat and electricity. It conducts electricity due to the presence of mobile electrons in the crystal lattice. Read graphite as an allotrope of Carbon in Chapter 21.4 of part I.



Bonding and Shapes of Simple Molecules

Some examples of linear molecules are hydrogen, oxygen, chlorine, hydrogen chloride and carbon (iv) oxide molecules. They have linear shapes. Their shapes can be represented as follows:

Structure	Formular	Shape	Name
H - H	H ₂	Linear	Hydrogen molecule
O - O	O ₂	"	Oxygen molecule
Cl - Cl	Cl ₂	"	Chlorine molecule
H - Cl	HCl	"	Hydrogen chloride gas
O=C=O	CO ₂	"	Carbon(iv) oxide gas

Fig. 7.22

Carbon (iv) oxide:

In this molecule, CO₂, the carbon atom is bonded to each oxygen atom covalently. The electrons in the covalent bonds experience mutual repulsion and produce a linear shaped molecule.

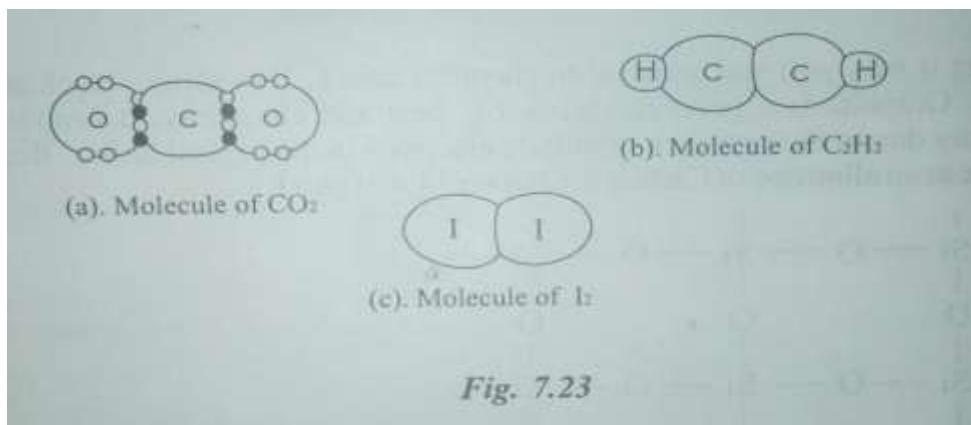


Fig. 7.23

Linear Molecule

All these linear molecules are diatomic with the exception of CO₂ which is triatomic. The bond angle in the linear molecules is 180°. Some examples of simple molecules with non-linear shapes are water and ozone. Each of them is triatomic. Their structures can be represented diagrammatically as follows:

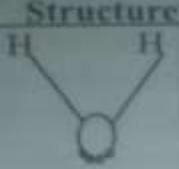
Structure	Formular	Shape	Name
	H ₂ O	Non-Linear	Water
	O ₃	Non-Linear	Ozone

Fig. 7.24

Some examples of molecules with tetrahedral shape are methane and silicon(iv)oxide. They are covalent compounds. Their structures are represented thus:

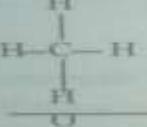
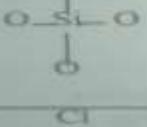
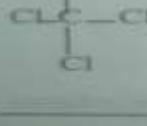
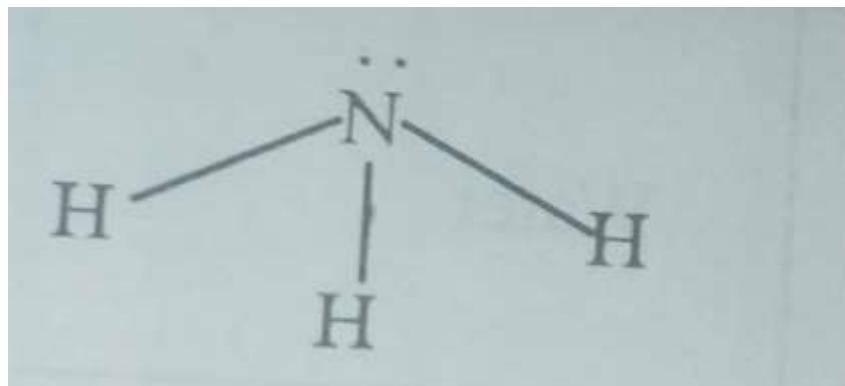
Structure	Formular	Shape	Name
	CH ₄	Tetrahedral	Methane
	SiO ₄	Tetrahedral	Silicon(iv) Oxide
	CCl ₄	Tetrahedral	Tetrachloro- methane

Fig. 7.25

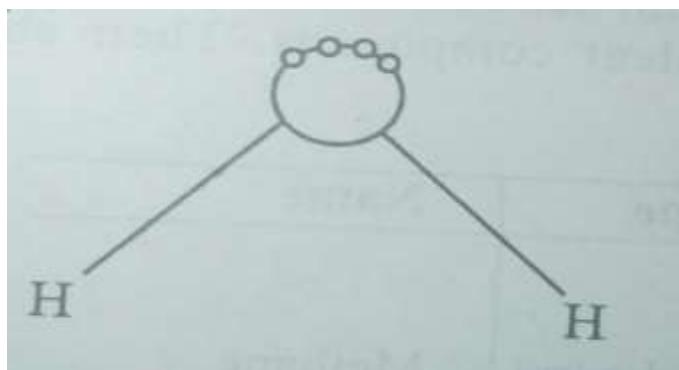
The inter bond angles in methane have the tetrahedral value of $109^\circ 28'$

Ammonia, NH₃:



An ammonia molecule consists of three hydrogen atoms bonded covalently to a nitrogen atom. A lone pair of electrons on nitrogen forces the bonding pairs closer together than in methane. The H - N - H angle is therefore reduced to 107°. The molecule of ammonia is trigonal pyramid.

Water, H_2O :



In the water molecule, the oxygen atom has two bonding pairs of electrons and two lone pairs. Due to mutual repulsion, the two bonding pairs are still brought more closely together. The H-O-H angle is only 105°. The molecule is V - shaped.'

Exercises

1. Name five types of bonding that held chemical substances together.
- 2.a. Differentiate between covalent bonding and electrovalent bonding.
- b. What are the types of bonding present in each of the following substances?
(i). NH_3
(ii). CO_2 (iii). CaO (iv). NaCl (v). NH_4^+ (v). $[\text{Cu}(\text{NH}_3)_4]^{2+}$
3. What are the properties of covalent compounds?
4. What is the difference between covalent bonding and co-ordinate covalent bonding?
- b. Write down two:
 - i. Covalent Compounds
 - ii. Co-ordinate Covalent substances or ions
5. Define (a) Metallic bonding (b) hydrogen bonding (c) Draw the Structure of two molecules of water showing how hydrogen bond held the two molecules together.
6. What are the allotropes of sulphur?
7. Write three properties of:
 - (a) Camphor (b) NaCl (c) Iodine

8. What are the differences between diamond and graphite?
9. Has bond type any relationship with the properties of.
 - (a) Camphor (b) NaCl? Explain.
10. Name five simple molecules with linear shapes.
11. Name two simple molecules with non-linear shapes.
12. Draw the structure of: (a) CCl_4 (b) SiO_2 (c) CH_4
13. Why is the H-N-H angle in NH_3 less than the tetrahedral value of $109^\circ 18'$?

CHAPTER 8

OXIDATION NUMBER (O.N.) AND IUPAC NOMENCLATURE.

The oxidation number of an element in any compound or ion is the electrical charge the element has in the compound or ion. Oxidation number of Na in NaCl is +1

8.1 Rules to be Born in mind while Determining Oxidation Numbers:

1. The oxidation number of all elements in the uncombined form or in the free state is zero. For instance the oxidation number of these substances are zero: K, Na, Cl₂, Al, P₄. This is usually represented as follows: K⁰, Na⁰, Cl₂⁰, Al⁰, P₄⁰.
2. The oxidation number of simple ion is the charge of the ion. Therefore, the oxidation numbers of K⁺, Na⁺, Al³⁺, Mg²⁺ and Cl⁻ are +1, +1, +3, +2, and -1 respectively. The oxidation of SO₄²⁻, NO₃⁻, NH₄⁺, OH⁻, PO₄³⁻ and MnO₄⁻ are -2, -1, +1, -1, -3 and -1.
1. The sum of the oxidation numbers of all the elements in a compound is zero. The oxidation number of CaCO₃ is zero.

8.2 Calculating the Oxidation Number of an Element in a Radical or a Compound.

Worked Examples

Stress the fact that oxidation number is either zero, a positive or negative integer, and is usually with respect to only one atom of an element.

Example I

Calculate the oxidation number of Copper in Cu₂O. Let x represents the O.N. of each copper atom, O is -2 In a neutral compound, sum of all O.N. is zero

∴ In Cu₂O: Cu₂O

$$\begin{aligned} 2x + (-2) &= 0 \\ 2x - 2 &= 0 \\ 2x &= 0 + 2 \\ 2x &= +2 \\ x &= \frac{+2}{2} = +1 \end{aligned}$$

∴ O.N. of a copper atom is Cu₂O = +1

IUPAC name: Copper (I)oxide

Conventional name: Cuprous oxide

Example 2

What is the oxidation number of phosphorus in phosphine, PH₃?

In PH₃, P is more electronegative than H

Hence, each H atom has O.N. of +1.

∴ In Ph₃ : PH₃

$$P + 3(+1) = 0$$

$$P + 3 = 0$$

$$P = 0 - 3$$

$$P = -3$$

∴ O.N. of phosphorus atom, P in PH₃ is -3.

IUPAC name: Phosphorus trihydrite

Conventional name: Phosphine

Example 3

Calculate the oxidation number of X in X₂O₇²⁻

O.N. of each O atom = -2

The net charge on the ion = -2

∴ In , X₂O₇²⁻ , X₂O₇²⁻

$$2x + 7(-2) = -2$$

$$2x - 14 = -2$$

$$2x = -2 + 14$$

$$2x = +12$$

$$x = \underline{+12}$$

$$2 = +6$$

∴ O.N. of each X atom in X₂O₇²⁻ is +6

Example 4

What is the oxidation number of nitrogen in HNO₃?

In HNO₃, the oxidation number of H atom is +1, while that of O is -2 per atom. Sum of all O.N. is zero.

Let the O.N. of nitrogen be N.

In HNO₃: HNO₃

$$+1 + n + 3(-2) = 0$$

$$+1 + n - 6 = 0$$

$$N = 0 - 1 + 6$$

$$N = +5$$

The O.N. of nitrogen atom, N in HNO_3 is +5.

IUPAC name: Trioxonitrate (v) acid

Conventional name: Nitric acid.

Example 5

What is the oxidation number of Al in Al_2S_3

In Al_2S_3 : Al_2S_3

$$\begin{array}{rcl} 2\text{Al} + 3(-2) & = & 0 \\ 2 \text{ Al} - 6 & = & 0 \\ 2 \text{ Al} & = & 0 + 6 \\ 2 \text{ Al} & = & +6 \\ \text{Al} & = & \underline{+6} \\ & 2 & = +3 \end{array}$$

The O.N. of Al in Al_2S_3 is +3

IUPAC name: Alulminum sulphide

Note: With adequate practice, it is possible to assign oxidation number to an atom in a formula, mentally, in one single step. For instance in NaNO_3 , Na atom is +1, while the three O atoms is -6; N atom will be +5, so that in $\text{NaNO}_3 = +1 + 5 - 6 = 0$

IUPAC name: Sodium trioxonitrate (v)

Conventional name: Sodium nitrate

8.3 Importance of Oxidation Number

The concept of oxidation number has been found useful in a number of ways:

In naming inorganic compounds by the IUPAC rules.

In determining the number of chemical bonds that can be formed by an atom (covalency)

In identifying oxidizing and reducing agents in redox reactions.

In balancing equations of redox reactions.

Exercises

1 (a) Define oxidation number

2 Determine the oxidation number of

(a) sulphur in SO_2 ; H_2SO_3 ; SO_4^{2-}

(b) Nitrogen in NH_3 ; N_2O ; NO_2^-

(c) Chlorine in Cl₂; HC10

CHAPTER 9

NATURE OF MATTER

Matter is anything that has mass and occupies space. It includes all living things such as plants and animals, and non-living things like stone, wood, water, air, sand, shoe, sodium chloride etc. it can be an atom, element, mixture or compound.

9.1 States of Matter: There are three states of matter. These are solid, liquid and gaseous states.

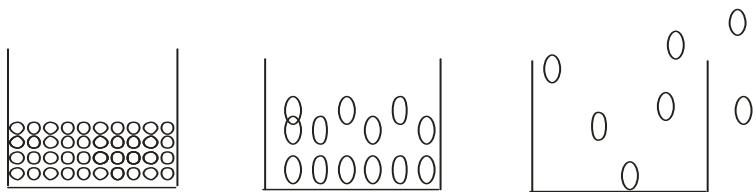


Fig. 9.1

The particles of matter in solid state are held together by some cohesive forces. The particles exhibit vibrational motion. Solid has a definite shape. The particles of matter in the liquid state are held together by cohesive forces less than that in the solid states. Liquids assume the shape of the containing vessels. They can flow. The particles of matter in the gaseous state are usually in rapid random motion. Gases occupy any available space because of these types of movement they exhibit. They can be compressed. Some substances can exist as solid, liquid and gas at various temperatures. For example, water can exist as solid at a very low temperature, it can exist as liquid at ordinary temperature and as a gas (vapour) at a temperature of 100°C or 373°K and pressure of one atmosphere. Some substances can exist in only one state. An example of this is stone. There are other substances that can exist in two states. For example iodine or ammonium chloride exists in solid state at ordinary temperature but when heated changes directly into the gaseous state. For such substance (sublimates), the vapour pressure of the solid becomes equal to an atmospheric pressure before its melting point is attained. Matter can change from one form to another by either evaporation or sublimation.

Evaporation:

This involves the heating of a liquid (e.g water) so that the liquid changes to gaseous state. The difference between evaporation and boiling are as follows.

Evaporation	Boiling
Energy is supplied to the evaporating particles by the liquid particles	Energy is supplied to the evaporating particles in form of heat from an external source (e.g Bunsen burner)
It can occur at any temperature	It occurs at a fixed temperature. For examples, water boils at a temperature of 100°C or 373°K
Evaporation can take place independent of the atmospheric pressure	Boiling occurs when the vapour pressure of the liquid is equal to atmospheric pressure. Lowering the atmospheric pressure decreases the boiling point of a liquid. Introduction of an impurity in a liquid increases its boiling point.

Sublimation:

This is a process by which substances change directly from solid state to gaseous state (when heated) without passing through the intermediate liquid state. The substances are said to sublime. Examples of substances that sublime are iodine, ammonium chloride, naphthalene, camphor, menthol etc.

9.2 Kinetic Theory of Matter

Kinetic theory is the theory which deals with the way in which the arrangement of the particles of a substance determine the properties of the substance and in particular the state in which the substance is likely to be found under given conditions.

Solid

In solid, the particles or atoms are closely packed. The force that binds the molecules of the solid together is very strong. The molecules or particles in solid have no freedom of movement. However, they show or exhibit vibrational motion about their mean position. Solids have definite shape and volume. Solid has high density. When a solid melts, there is usually little change in volume.

Liquid

In liquid, particles are relatively spaced out. The force that binds the particles or molecules together is not as strong as the force experienced by particles in the solid state. So the molecules, particles or ions have little freedom of movement. The molecules could move very short distance. The liquid have definite volume but no definite shape. The liquid boils at a temperature when the vapor pressure of the liquid is equal to the atmospheric pressure. On changing to gas, there is a large change in volume.

Gas

In gases, the particles, or molecules are widely spaced out. The force that matter binds the molecules of the matter together is very weak. The molecules have greater freedom of movement.

They could move about freely in space. Gases have no definite shape and volume. Gas molecules move about randomly, colliding with one another and the walls of the container. (Read chapter 2).

9.3 Assumption of Kinetic Theory about a Perfect Gas or An Ideal Gas:

1. The gas molecules move in straight line at great speed colliding with each other and with the walls of the container.
2. The actual volume occupied by the gas is negligible compared with the volume of its container.
3. The forces of attraction and repulsion between the gas molecules is negligible.
4. The average kinetic energy of the gas molecules measures (or is proportional to) the temperature of the gas.
5. The collision between the gas molecules is perfectly elastic.

9.4 Phenomena that Support the Kinetic Theory of Matter

The phenomena that support the kinetic theory of matter are:

1. Diffusion
 2. Osmosis
 3. Dialysis
 4. Brownian movement
1. **Diffusion:** is the movement of a fluid (gas or liquid) from a region of higher concentration to a region of lower concentration.
 2. **Osmosis:** is the passage of solvent molecules from area of lower concentration to that of higher concentration through a semi-permeable membrane.
 3. **Dialysis:** is the separation of crystalloids from colloidal particles using semi-permeable membrane. Crystalloid is anything that is soluble in water.

4. **Brownian:** the random motion of small particles suspended in a liquid or gas is called Brownian motion. The particles of matter are in constant motion.

This constant and zig-zag movement of particles can be observed when:

- I. A piece of sulphur suspended in water is viewed under a microscope or
- ii. A pollen grain suspended in a drop of water is observed under a microscope.

Brownian movement or the random movement of particles decreases from gases to liquids to solids.

The above listed four phenomena explains:

- a. The evidence of the movement of particles in liquid and gases.
- b. The evidence of the existence of particles.

9.5 Changes in Matter:

Two types of changes undergone by matter are physical and chemical changes.

Physical Changes:

This is the type of change that results in the formation of no new substance. Examples of physical change are:

1. All cases of the melting of a solid to a liquid (or the reverse).
2. All cases of vaporization of a liquid (or the reverse).
3. Magnetization of iron
4. The heating of a metal wire by electricity.

Chemical Change:

This is the type of change that results in the formation of new substance. Examples of chemical change are:

1. The burning of any substance in air.
2. The rusting of iron
3. The addition of water to calcium oxide
4. Explosion of natural gas or hydrogen in air.

Differences Between Physical and Chemical Changes:	
Physical Change	Chemical Change
1. Produces no new substance	Always produces a new substance
2. Is generally reversible	Is generally not reversible
3. Is not accompanied by great heat change (except latent heat effects accompanying changes of state)	Is usually accompanied by considerable heat change.
4. Produce no change of mass	Produces individual substances whose masses are different from those of the original individual substances. Thus, if two substances, A and B, react chemically to form substances Y and Z, the mass of Y will be different from the mass of A and B, and the mass of Z will be different from the mass of A or B

9.6 Classification of Matter:

Matter can be classified into element, compound and mixture.

Element:

An element is a substance which cannot be split up into two or more simpler substances by any known chemical process. Examples of element are: oxygen, hydrogen, carbon, sodium, potassium, calcium, nitrogen, chlorine, silicon, aluminum etc.

Compound:

A compound is a substance which contains two or more elements chemically combined together. Examples of compounds and the elements that make them up are:

Compound	Elements Contained
Water	Hydrogen; Oxygen,
Sugar	Carbon; Hydrogen; Oxygen,
Sodium Chloride (common salt)	Sodium; Chlorine
Potassium trioxonitrate (v)	Potassium; Nitrogen; Oxygen
Marble	Calcium; Carbon; Oxygen
Chalk	Calcium; Carbon; Oxygen
Sand	Silicon; Oxygen
Tetraoxosulphate (vi) acid	Hydrogen; Sulphur; Oxygen

Mixture: A mixture is anything which contains two or more substances physically combined together. Or it can be defined as a physical combination of two or more substances. Examples of mixture are; a mixture of sand and stone, a solution of common salt in water (brine). Other examples of mixture and their constituents are;

Mixture	Constituents
Palm wine	Water, sugar, yeast, vitamins, fat, mineral salts and alcohol
Milk	Water, fat, protein, sugar, hormones and vitamins
Crude oil or petroleum	Petrol, engine oil, kerosene, diesel oil etc
Sea water	Water, bacteria, mineral salts and remains of organic matter.
Urine	Urea, mineral salts and water
Air	Oxygen, carbon (iv)oxide, nitrogen, water vapour and noble gases.

Differences Between Mixtures and Compounds are Summarized in The Table Below:

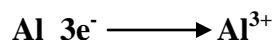
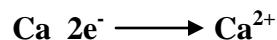
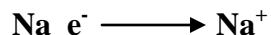
Mixture	Compound
The properties of a mixture are the sum of the properties of the constituent substances that make up the mixture	The properties of a compound are different from those of its constituent elements.
The constituents can be separated from one another by physical methods.	The constituents can be separated by chemical methods
Mixture may vary in composition. This means that the composition of a mixture is variable.	The composition of compound is not variable. The constituents are present in a fixed ratio by mass
Mixing is not usually accompanied by evolution of heat or volume change (for gases)	In the formation of some chemical compounds, heat is usually evolved or absorbed.

9.7 Classification of Element:

Element has already been defined above. It can be classified or divided into metal and non-metal.

Metal:

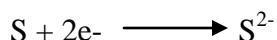
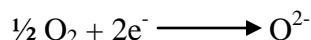
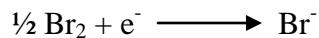
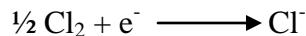
A metal is an element which can ionize by electron loss. The number of electrons lost per atom is equal to the valency of the metal and the ion formed carries an equal number of positive charges as shown below;



Some examples of metals are sodium, magnesium, iron, gold, silver, mercury. Mercury is a liquid metal.

Non-metal:

A non-metal is an element, which can ionize by electron gain. This means that non-metals are electronegative. The number of electrons gained per atom is the valency of the non-metal. The ion formed carries an equal number of negative charges, as shown below.



9.8 Comparison Of Physical Properties of Metals and Non-metals.

Metals	Non metals
Metals have relatively high densities	Non-metals have relatively low densities
They have characteristic lustre and can be polished	Lack luster and cannot be polished
With the exception of mercury which is liquid at room temperature all metals are solid at room temperature with high melting and boiling points.	They are mostly gases with low boiling and melting points.
They have high tensile strength and as a result can withstand stress or strain.	Are usually brittle
Metals are malleable and ductile and as a result can be hammered into sheets or drawn into wire.	Non-metals are neither malleable nor ductile.
Metals are good conductors of heat and electricity	They are poor conductors of heat and electricity.
Metals give off sound when hit, which differs from metal to metal. This means that they are sonorous	They are not sonorous.

9.9 Comparison of Chemical Properties of Metal and Non-metals.

Metals	Non metals
Metals are electropositive	Non-metals are electronegative
They ionize by electron loss in solution to form cations	They ionize in solution by electron gain to form anion
They form ionic or electrovalent compounds	They form covalent compounds
Metals are reducing agent or electron donor	Non-metal are oxidizing agent or electron acceptor
Metals are deposited at the cathode during electrolysis	Non-metals are deposited at the anode during electrolysis.
Metals react with oxygen to form basic oxides eg. $4K + O_2 \longrightarrow 2K_2O$	Non-metals form acidic oxides with oxygen eg. $C + O_2 \longrightarrow CO_2$
Metals higher than hydrogen in the electro-chemical series can displace hydrogen from acids eg. $Zn_{(s)} + 2HCl_{(aq)} \longrightarrow ZnCl_{2(aq)} + H_2$	They cannot displace hydrogen from acids
In metals, atoms are held together by metallic bonds	Atoms in non-metals are usually held together by covalent or ionic bond

Exercises

1. a. What is matter?
 - b. Write three examples of matter.
 - c. What are the three states of matter?
2. What do you understand by these terms?
 - a. Evaporation
 - b. Sublimation
 - c. Name three substances that can sublime.
3. What do you understand by:
 - (a) Kinetic theory (b) Brownian motion?
4. Write two differences between:
 - (a) Solid and liquid (b) Liquid and Gas
5. What are the five assumptions of kinetic theory about an ideal gas?
6. Write down the four phenomena that support the kinetic theory of matter.

7. Define these terms and give an example of each
 - a. Physical change
 - b. chemical change
 - c. What are the differences between physical and chemical changes?
8. Define and give two examples of each of these terms.
 - a. element
 - b. compound
 - c. Mixture
9. Write three differences in chemical properties between metal and non-metal.

CHAPTER 10

THE GAS LAWS

10.1 Boyle's Law:

Boyle's law state that the volume of a given mass of gas is inversely proportional to its pressure at constant temperature. This can be expressed mathematically as:

$$V \propto \frac{1}{P}$$

$$V = \frac{k}{P}$$

$$PV = K$$

$$(P_1 V_1 = P_2 V_2) \text{ Boyle's law}$$

If P_1, V_1 are the initial pressure and volume, and P_2, V_2 are the final pressure and volume, then it can be shown that:

$$P_1 = P_2$$

We measure volume in cm^3 . We measure pressure in mmHg. In Boyle's law, each time the pressure on the same amount of gas is increased, the gas decreases proportionally in volume, and vice versa. This general behaviour was observed by Robert Boyle's, an English scientist. Hence, the formulated law was named after him.

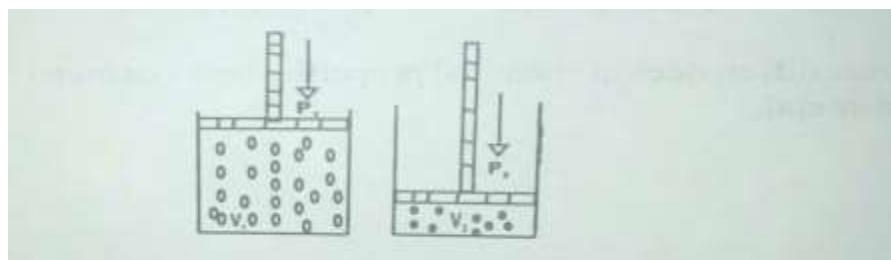


Fig. 10.1 Boyle's law can be represented graphically as:

Boyle's can be represented graphically as:

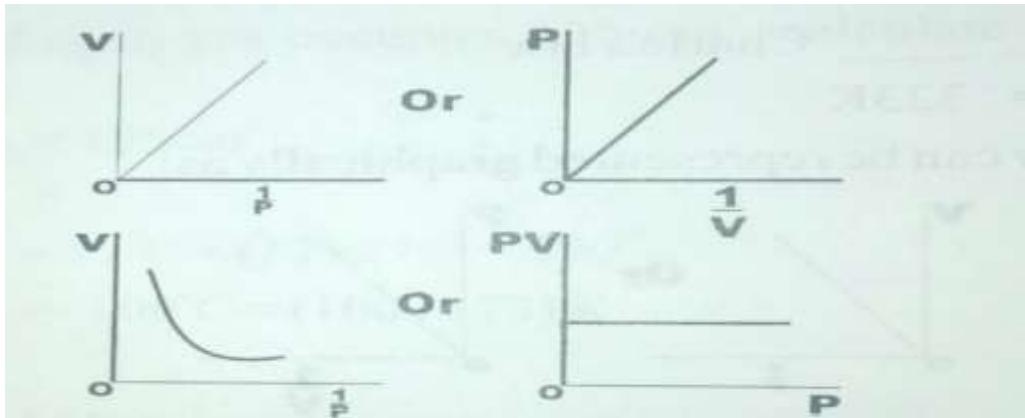


Fig 10.2

Question

1000cm³ of a gas have a pressure of 350 mm Hg. Calculate the volume at 700 mm Hg

Solution

$$P_1 V_1 = P_2 V_2$$

$$P_1 V_1 = V_2$$

$$350 \times 1000 = V_2$$

$$700 = V_2 = 500\text{cm}^3$$

10.2 Charle's Law:

Charle's law states that the volume of a give mass of gas is directly proportional to its absolute temperature (Kelvin temperature) at constant pressure. It means that when temperature is increasing, the volume is increasing.

Mathematically, it can be expressed as:

$$V \propto T$$

$$V = KT$$

$$\frac{V}{T} = K (\text{constant})$$

$$T$$

If V_1, T_1 are the initial volume and temperature of a given mass of gas and V_2, T_2 are the final volume and temperature of the same mass of gas, then

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{Charle's law}$$

$$\frac{V_1}{T_1} = 323\text{K}$$

Charle's law can be represented graphically as:

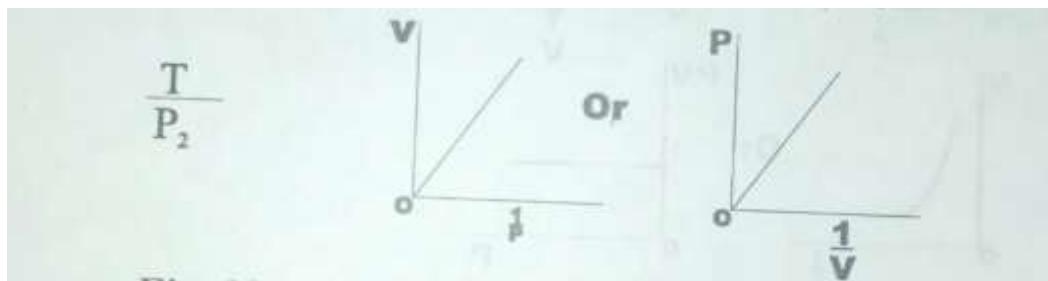


Fig. 10.3 Graphical representation of Charle's law

10.3 Conversion of Celsius Temperature 0°C to Kelvin (absolute) Temperature

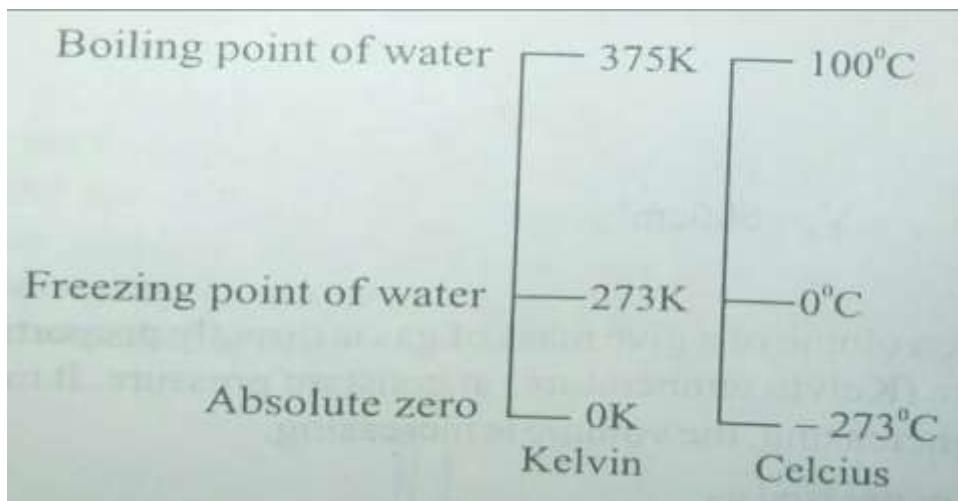


Fig. 10.4

To convert Celsius scale to absolute temperature, we add 273 to it. To convert Kelvin temperature to Celsius scale, we subtract 273.

Example 1

Convert 50°C to absolute temperature.

$$\text{TK} = 273 + 50^{\circ}\text{C}$$

Example 2

Convert 292K to Celsius temperature.

$$^{\circ}\text{C} = 292 - 273$$

Example 3

At 17°C , a sample of hydrogen gas occupies 125cm^3 , calculate its volume at 100°C .

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$
$$V_1 = 125\text{cm}^3$$
$$T_1 = 17^{\circ}\text{C} = (17+273) = 290$$
$$V_2 = \frac{V_1 T_2}{T_1}$$
$$T_2 = 100^{\circ}\text{C} = (100+273) \text{ K}$$

$$V_2 = \frac{125 \times 373}{290} = \frac{46625}{290} = 161\text{cm}^3$$

10.4 Pressure's Law

Pressure's law states that the pressure of a given mass of gas is directly proportional to its absolute temperature at constant volume.

This law can be represented mathematically as:

$$P \propto T \text{ at constant volume}$$

$$P = KT$$

$$\frac{P}{T} = K \text{ (Constant)}$$

For a gas with P_1 and T_1 as its initial pressure and temperature and P_2 and T_2 as its final pressure and temperature, we have that at constant volume.

$$\frac{P_1}{T_1} \propto \frac{P_2}{T_2}$$

10.5 General Gas Law

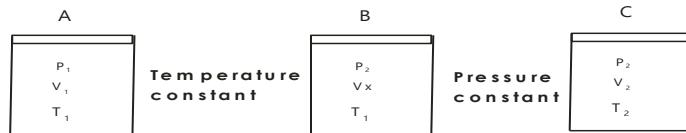


Fig 10.5

The relationship between pressure, volume, absolute temperature and the number of moles of gas constitute the general gas law. It is expressed mathematically as:

$$P \times V = n \times R \times T$$

$$\text{Or } PV = nRT$$

Where P is the pressure, V = Volume, n = number of moles of gas, R = universal gas constant and T = absolute temperature. From the above equation, nR is constant

Therefore, $\frac{PV}{T} = \text{a constant}$

The combination of Boyle's and Charle's law gives a single mathematical expression called an ideal gas equation i.e.

T_1 = Initial temperature

P_2 = Final pressure

V_2 = Final volume

T_2 = Final temperature

$P \propto \frac{1}{V}$ (Temperature constant)

$PV = \text{a constant}$

Therefore $P_1V_1 = P_2V_2$

Charle's law say that: $V \propto T$ (pressure constant)

$\frac{V}{T} = \text{a constant}$

Therefore $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

For a given mass of gas to change from $P_1V_1T_1$ to $P_2V_2T_2$, it has to undergo an intermediate stage $P_2V_XT_1$ (See Fig 16.1). This means that the gas can obey both Boyle's and Charle's law to enable us formulate an ideal gas equation.

For the given mass of gas to change from A to B (Fig 16.1), Boyle's law is obeyed. The temperature is constant.

Therefore: $P_1V_1 = P_2V_X$

Hence $V_X = P_1V_1$ ----- (1)

For the same gas to change from B to C, Charle's law is obeyed. The pressure is constant.

Therefore: $\frac{V_X}{T_2} = \frac{V_2}{T_2}$

$$\text{Hence, } V_x = \frac{T_1 V_2}{T_2} \quad \dots \quad (2)$$

$$\text{Since, } V_x = \frac{P_1 V_1}{P_2} = \frac{T_1 V_2}{T_2}$$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad \dots \quad (3)$$

Question 1

A gas occupied a volume of 500cm^3 at 17°C and 870 mm Hg. Calculate the volume of the gas at 10°C and 750 mm Hg.

Solution

$$P_1 = 870 \text{ mm Hg}$$

$$V_1 = 500\text{cm}^3$$

$$T_1 = 17^\circ\text{C} (17+273) = 290\text{K}$$

$$P_2 = 750 \text{ mm Hg}$$

$$V_2 = ?$$

$$T_2 = 10^\circ\text{C} (10+273) = 283\text{K}$$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$P_1 V_1 T_2 = P_2 V_2 T_1$$

$$V_2 = P_1 V_1 T_2$$

$$V_2 = \frac{870 \times 500 \times 283}{750 \times 290} = \frac{246210}{435}$$

$$= 566\text{cm}^3$$

S.T.P means standard temperature and pressure.

Standard pressure = 760mm Hg or 1 atmosphere or Nm (Newton metre)

Standard temperature = 0°C or 273K

10.6 Dalton's Law of Partial Pressure

The law states that in a mixture of gases which do not react chemically, the total pressure exerted by the mixture is the sum of partial pressures of the individual gasses that make up the mixture.

$$PT = PA + PB + PC$$

PT = Total pressure of the mixture

PA, PB, and PC = Pressure of gases A, B and C that make up the mixture.

A gas is likely to be saturated with water vapour if it is collected over water and the total pressure becomes.

$$PT - P \text{ gas} + P \text{ water vapour}$$

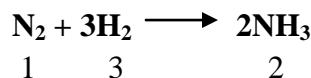
The correct pressure of the dry gas is obtained by subtracting the vapour pressure from the total pressure.

10.7 Gaylussac's Law of Combining Volume

The law state that when gases react, they do so in volumes which bear simple ratio relationship to one another and to the product if gaseous at constant temperature and pressure.

Gay-lussac's law was explained by Avagadro's hypothesis.

Example



So the volume ratio is 1:3:2

10.8 Avagadro's Law

The law states that equal volumes of all gases at the same temperature and pressure contain the same number of molecules.

It explains then 10cm^3 of N_2

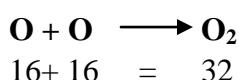
10cm^3 of O_2

and 10cm^3 of H_2

Must all contain the same number of molecules at the same temperature and pressure.

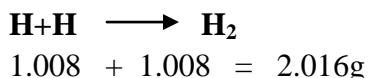
10.9 Mole and Molar Volume

1 mole is the atomic mass or molecular mass of a substance expressed in grammes. The unit for the mole is the gram. The molar mass of oxygen molecules is obtained thus:



Therefore one mole of oxygen gas weighs 32g.

One mole of Hydrogen, H₂ weighs



One mole of sodium atom, Na weighs 23g

One mole of oxygen atom weighs 16g

The number of particles in one mole of a substance is 6.02×10^{23}

This number is called Avogadro's number or constant.

Question 2:

How many moles are there in 54g of Aluminum (Al = 27)³

Solution

$$27\text{g} = 1 \text{ mole}$$

$$54\text{g} = ?$$

$$1 \times \frac{54}{27} = 2 \text{ moles}$$

To convert gram to mole,

We divide the gram by the atomic mass or the molecular mass.

Question 4

Convert 54g of Al to mole (Al = 27)

Solution

$$\frac{54^2}{27_1} = \underline{\underline{2}} = 2 \text{ moles}$$

Question 5

Convert 250 grammes of CaCO₃ to mole (Ca = 4O, C = 12, O = 16)

Solution



$$40+12+(16\times 3)$$

$$40+12+48 = 100$$

$$\frac{250}{100} = 2.5 \text{ moles}$$

$$100$$

To convert mole to gram

Multiply the mole by the atomic or the molecular mass.

Question 6

Convert 0.5 mole of Al to gram (Al = 27)

Solution

$$0.5 \times 27 = 13.5 = 13.5\text{g}$$

Molar Volume is the volume occupied by one mole of a gas at S.T.P. and numerical equal to 22.4 dm^3

One mole of these gases (O_2 , H_2 , N_2 , CO_2 , CO etc) occupy 22.4 dm^3 at S.T.P

Question 7

Calculate the volume of carbon (iv) oxide produced at S.T.P when 50g of calcimtrioxocarbonate (iv) is completely decomposed by heat.



$$40+12+(16\times 3)$$

$$40+12+48$$

$$\begin{array}{rcl} 100\text{g} & \longrightarrow & 22.4 \text{ dm}^3 \\ 50\text{g} & \longrightarrow & 22.4 \times \frac{50}{100} \\ & & = 11.2\text{dm}^3 \end{array}$$

Molar volume is the volume occupied by one mole of a gas at S.T.P. and is numerically equal to 22.4 dm^3

One mole of these gases (O_2 , H_2 , N_2 , CO_2 , CO etc) occupy 22.4 dm^3 at S.T.P

Question 7

Calculate the volume of carbon (iv) oxide produced at S.T.P when 50g of calciumtrioxocarbonate (iv) is completely decomposed by heat.



$$40+12+(16\times 3)$$

$$40+12+48$$

$$100\text{g} \quad 22.4 \text{ dm}^3$$

$$\frac{50\text{g}}{100} = \frac{22.4 \times 50}{11.2\text{dm}^3}$$

10.10 Graham's Law of Diffusion

State that at constant temperature and pressure, the rate of diffusion of a gas R is inversely proportional to the square root of its density, d.

$$R = \frac{1}{\sqrt{d}}$$

The relative rates at which two gases diffuse are inversely proportion to the square roots of their densities or molar masses.

Hence;

$$\frac{R_1}{R_2} = \sqrt{\frac{d_2}{d_1}} \quad \text{or} \quad \frac{R_1}{R_2} = \sqrt{\frac{m_2}{m_1}}$$

Where R_1 = rate of diffusion of gas 1

R_2 = rate of diffusion of gas 2

d_1 = density of gas 1

d_2 = density of gas 2

m_1 = molar mass of gas 1

m_2 molar mass of gas 2

10.11 Relative Vapour Density:

The vapour density of a gas or vapour is the number of times a given volume of the gas or vapour is as heavy as the same volume of hydrogen at a particular temperature and pressure.

The vapour density (V.D) of a gas is expressed as

$$\begin{aligned} \text{V.D.} &= \frac{\text{Mass of a given volume of the gas (vapour)}}{\text{Mass of an equal volume of hydrogen}} \\ &= \frac{\text{Mass of 1 molecule of gas (vapour)}}{\text{Mass of 1 molecule of hydrogen}} \end{aligned}$$

But 1 molecule of hydrogen contains 2 atoms of hydrogen

$$\begin{aligned} \text{V.D.} &= \frac{\text{Mass of 1 molecule of gas (vapour)}}{\text{Mass of 2 atoms of hydrogen}} \end{aligned}$$

$$2 \times V.D = \frac{\text{Mass of 1 molecule of gas (vapour)}}{\text{Mass of 1 atom of hydrogen}}$$

$$\text{But } \frac{\text{Mass of 1 molecule of gas}}{\text{Mass of 1 atom of hydrogen}} = \text{Relative molecular mass of gas}$$

Therefore

$$2 \times V.D = \text{Relative molecular mass of gas}$$

So the relationship between the vapour density and the relative molecular mass of gas can be expressed as follows:

$$V.D = \frac{1}{2} \text{ Relative Molecular Mass of Gas}$$

$$\text{i.e. } V.D. = \frac{1}{2} \text{ RMM}$$

$$T = \frac{7 \times 60 \times 6}{30 \times 4}$$

$$t = 21 \text{ secs}$$

Exercise

1. State these gas laws and write their mathematical expression
 - a. Boyle's law
 - b. Charle's law
 - c. Pressure's law
 - d. Pressure's law
2. Draw graphical representation of:
 - a. Boyle's law
 - b. Charle's law
3. Convert these Celsius temperature ($^{\circ}\text{C}$) to Kelvin temperature (K).
 - a. 150°C
 - (b) -15°C
- At 23°C , a sample of hydrogen gas occupies 125cm^3 , what will the volume be at 106°C if the pressure remains constant.
4. Briefly explain and show how an ideal gas equation is derived from the combination of Boyle's and Charle's laws.
5. State these laws:
 - i. Gay-lussac's Law

- ii. Avogadro's law
 - iii. Graham's law of diffusion
6. Define these
- i. Mole (ii) Molar Volume and (iii) Vapour density
7. What is the relationship between the vapour density and the relative molecular mass of a gas.

CHAPTER 11

SEPARATION TECHNIQUES FOR MIXTURES

It was pointed out what a mixture is in chapter 9.6. It was also pointed out that the constituents of a mixture can be separated from one another by physical method. Therefore, the techniques for separating mixtures are essentially physical. A mixture can be in solid state, liquid state or gaseous state. The physical methods of separating mixtures are: Filtration, Decantation, Evaporation, crystallization, fractional crystallization, use of separation funnel, Distillation, Fractional Distillation, Sublimation, Precipitation, Magnetization, Chromatography, Sieving and use of centrifuge machine.

11.1 Filtration:

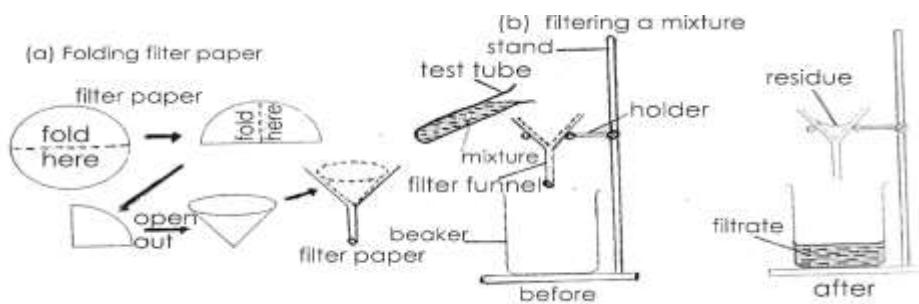


Fig 11.1

Filtration is used to separate insoluble solids from liquids or solutions. Suspensions can also be filtered to remove solids. Drinking water is filtered to remove insoluble solids. In this method filter paper is used. Filtration is used to separate chalk and water or sand and water. The insoluble sand will remain on the filter paper as residue while water will pass through the filter paper and is collected in the beaker as filtrate.

Industrially, filtration is employed in the purification of town water supply and in breweries. It is also employed in a chemistry laboratory to separate a mixture of soluble and insoluble substances. For example, a mixture of insoluble lead tetraoxosulphate (vi), (PbSO_4) and soluble zinc tetraoxosulphate (vi), (ZnSO_4) can be separated by the process of filtration.

11.2 Decantation:

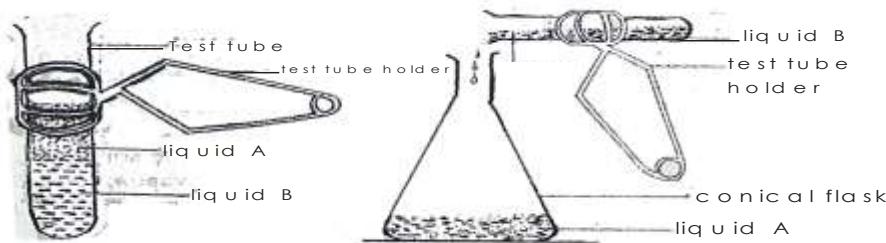


Fig11.2: Separation of two immiscible liquids by decantation

This method is used to separate a mixture of solid and liquid where the solid settles at the lower part of the test-tube and the liquid at the upper part. The upper liquid can then be separated from the solid by carefully running out the upper liquid into another container. This method can also be used in the separation of two immiscible liquid, where the lighter liquid A floats on the denser liquid B in the same test tube. The disadvantage of this method is that while pouring out the lighter upper liquid, the denser lower liquid may run off.

11.3 Evaporation:



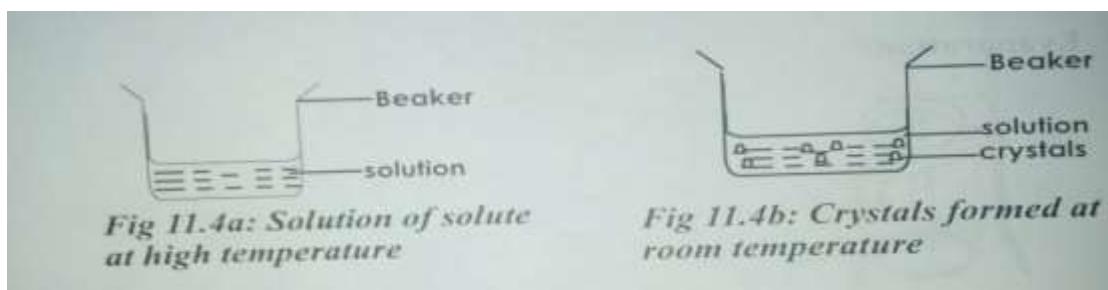
Fig 11.3: Apparatus for evaporation

This is a method used to separate a solid (solute) from a solvent. The solution is heated to dryness. Evaporation involves the heating of a solution to remove the solvent from the solid substance. A solution of common salt in water is called brine. Brine can be heated to evaporate all the water molecules leaving only solid salt after evaporation. Evaporation method is mainly used where the liquid component is not required. Evaporation is one of the fast methods of separating mixtures but the product obtained using this procedure may not be very pure. It is used where the components cannot be decomposed by heat.

11.4 Crystallization:

This method is also used in the separation of solids from solutions. The knowledge of the solubility of solids at various temperatures is very vital here. The use of this method is dependent

on the solubility of the component of the mixture at various temperatures. The solid to be separated is first of all dissolved in a solvent in which the component to be separated is very soluble at high temperatures and less soluble at lower temperatures. The solid is then dissolved in minimum amount of the solvent at high temperature, and the solution is allowed to cool. If the solid (solute) is in excess solvent, crystalline substance (solid) is separated by concentrating the solution by evaporation. As evaporation is going on, after a time, crystals will start to appear at the edge of the solution. When the crystalline substance is observed, the solution is poured into a crystallizing dish and allowed to cool. As the solution cools, solid particles called crystals will begin to appear in the solution. At room temperature, the crystals are obtained in the solution. The crystals can then be separated from the solution by filtration. It is washed with distilled water and allowed to dry on the filter paper.



11.5 Fractional Crystallization:

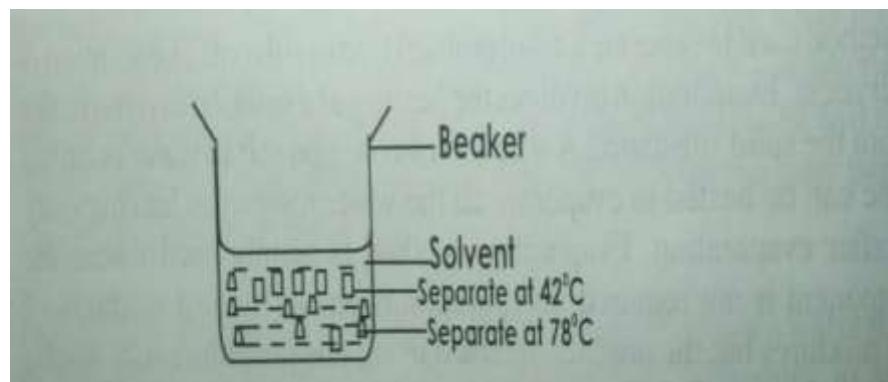


Fig 11.5: separating the various components of a mixture by varying the temperature

This method is used where there are more than one solute in a given solution. There must be difference in the solubility of the solutes at different temperatures. As the boiled solution is allowed to cool, the different crystals can be obtained at different temperatures. For example, at a temperature of 78°C the less soluble component will separate out of the solution (or it crystallizes) and at a temperature of about 42°C the crystals of the more soluble component will crystallize. Thus by varying the temperature, various fractions of the solute components can be separated.

11.6 Use of Separating Funnel:

Separating funnel can be used in separating two immiscible liquids. It can also be used in separating a mixture of two different solids in which the components are soluble in different immiscible solvents. For example, if solid A is soluble in solvent B, and solid C is soluble in D but not in B, a mixture of solids A and C can be separated by dissolving the solid mixture in the mixture of the two solvents in a separating funnel. The content of the separating funnel is shaken very well and allowed to settle. The less dense solution will then float on top of the denser solution. The control tap is then opened and the denser solution enters into a conical flask. To avoid being contaminated by the remains of the denser solution, the less dense solution is removed through the upper opening of the funnel.

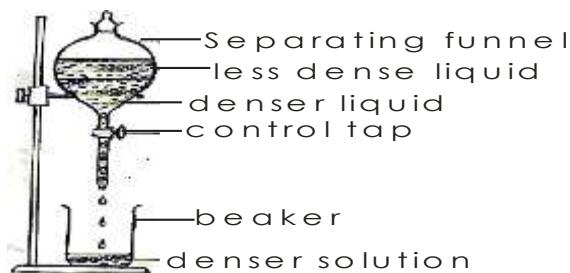


Fig 11.6: Using a separating funnel

11.7 Distillation:

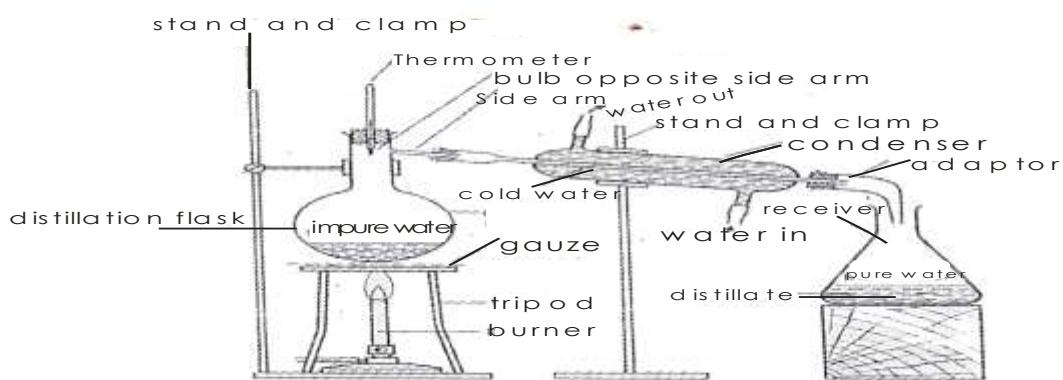


Fig 11.7

Distillation consists of the two basic processes of boiling and condensation. For example impure water can be purified by distillation because the impurities in water are solids which are non-volatile. The impure water is heated to make it boil. Its vapour is collected and cooled. As it

cools, the vapour condenses into pure distilled water. The non-volatile impurities remain behind. This method is commonly used in separating solvents from solutions especially when the solvent is needed. It can be used in separating the more volatile (lower boiling point) from less volatile substances. Pure water can be obtained from brine (common salt solution) by distillation. Distillation can also be employed in separating mixture of liquids with widely differing boiling points. The difference in the boiling points should be large so that only the more volatile constituent will be present in the vapour above the boiling point of the constituent, in the liquid mixture.

Liebig condenser is used to convert vapour to liquid while distilling. To enhance the condensation of the vapour, cold water is continuously being introduced through an inlet. The thermometer helps one to know when the component of the mixture has distilled off. For example, if the boiling point of the component is 60°C at a temperature of about 62°C that component must have distilled off. The condensed steam is called the distillate or condensate.

11.8 Fractional Distillation:

Fractional distillation technique is used in separating mixture of liquids with boiling points that are close together.

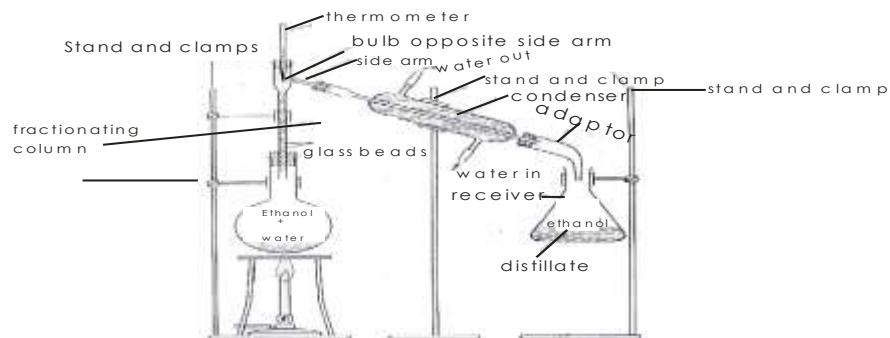


Fig 11.8: Fractional distillation

Fractional distillation method is similar to the ordinary distillation described above except that a fractionating column is fitted between the distilling flask and the condenser. This process is used to separate two or more mixture of volatile liquids with very close boiling points. A fractionating column is a tube with irregular interior. It is a glass column packed with glass beads, which exposes a large cooling surface to ascending vapours. On heating a mixture of, for example ethanol (boiling point 78°C) and water (boiling point 100°C), the liquid mixture boils and its vapour rises into the fractionating column. This vapour mixture contains a higher proportion of the more-volatile ethanol than water. The vapour begins to condense in the fractionating column. Water vapour condenses more easily than the ethanol vapour. The column is cooler at the top than at the bottom. Thus the higher the vapour mixture rises in the column, the richer it becomes in ethanol. Finally, at the top of the column, pure ethanol can be collected.

Any liquid in the above experiment boiling at a higher temperature than 78°C, will condense and run back into the flask. In this way, only ethanol reaches the top of the fractionating column as a vapour, and distills over into the receiver. At any particular time, the thermometer indicates the boiling point of the distilling liquid. A rapid rise in temperature, shows that the next constituent is beginning to distill off and should be collected in another receiver flask.

11.9 Industrial Application of Distillation:

- (a). Crude oil as obtained from the ground is not suitable for use. It is a mixture of many liquids, which must be separated using fractional distillation method.
- (b). Distillation is the major process used in the manufacture of spirits such as whisky, gin, rum and brandy. It serves both to separate the spirit from the raw materials and also to strengthen it (the distillate). The distillate or the condensed vapour will then contain a higher percentage of alcohol than the original alcohol/water mixture.

11.10 Sublimation:

There are substances which when heated changes directly from solid state to gaseous state without passing through the liquid state. Such substances are said to sublime and the process or the phenomenon is known as sublimation. Examples of the substances which sublime when heated are anhydrous aluminum chloride (AlCl_3), anhydrous iron (III) chloride (FeCl_3), benzoic acid and iodine. A substance which sublimes can be separated from other components of a mixture by heating the mixture.

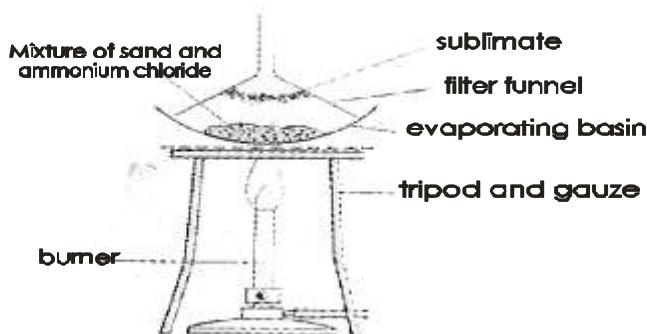


Fig 11.9 Sublimation

A mixture of sand and ammonium chloride can be separated as follows:

The mixture is placed in an evaporating dish or basin. An inverted filter funnel is placed on the dish as shown in Fig 11.9 above. The basin is put on a tripod and gauze and is heated gently. If the funnel becomes hot, strips of filter paper, sparked in water is placed on it to cool it. On

heating, ammonium chloride vaporizes wall of the funnel. The sublimate can be scraped off the funnel. Sublimate is a substance formed by sublimation.

11.11 Precipitation:

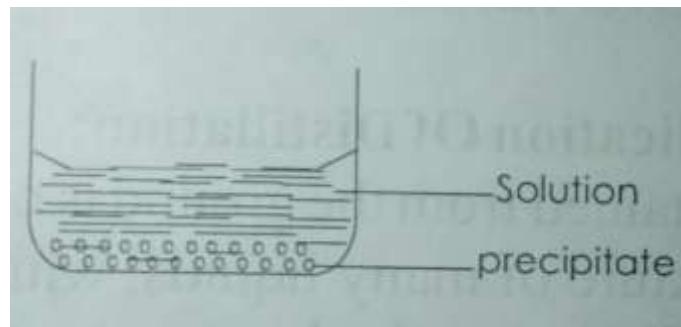


Fig. 11.10

Precipitation is the formation of an insoluble solid by a reaction, which occurs in solution. It is one of the separation techniques where chemical means is applied to facilitate the ease of separation by physical means (filtration). For example, if silver ion is to be removed from an aqueous solution of its salt (Silver trioxonitrate v), an aqueous solution of another compound such as Zinc Chloride is added. This will help in precipitating on insoluble silver compound such as silver chloride. The silver compound can then be separated from the soluble compound in solution by filtration..



This method uses the differences of solubilities of a solid in different miscible solvents. For instance, iron (ii) tetraoxosulphate (vi) is soluble in water and insoluble in ethanol. Addition of ethanol to an aqueous solution of iron(ii) tetraoxosulphate (vi) will precipitate the pure solid. To obtain a high degree of purity, the solid can then be recrystallised.

1.12 Magnetization:

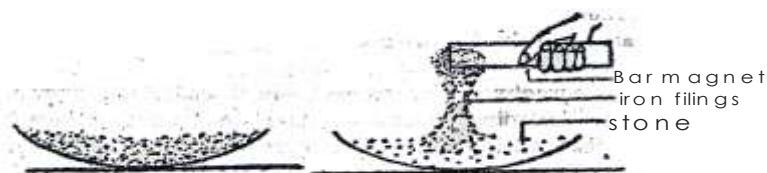


Fig 11.11 Separations of Iron filings from common salt using a bar magnet

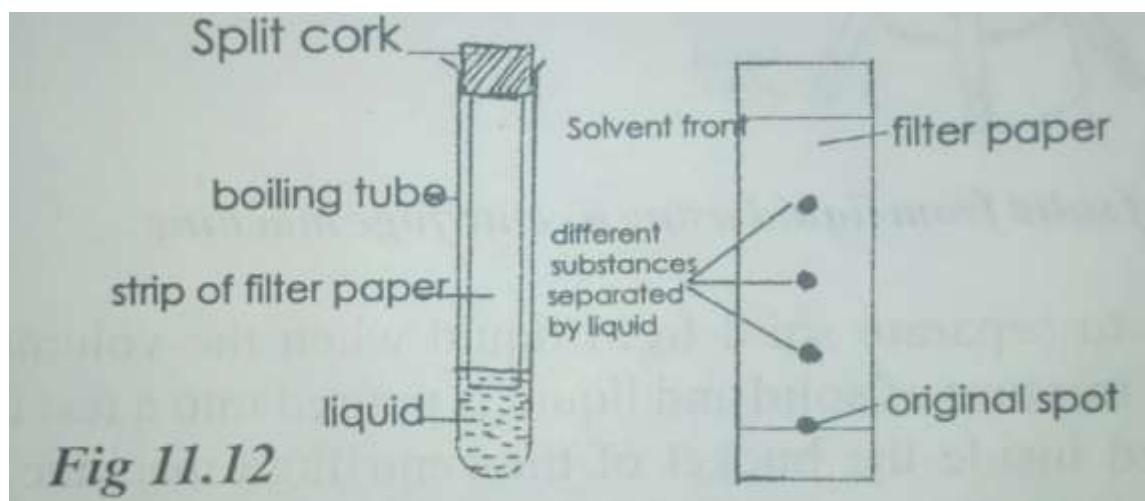
This is a method used in the separation of mixture where one of the components of the mixture is a magnetic material. In magnetization, any magnetic material in a mixture can be separated by bringing a bar magnet very close or in contact with the mixture. The magnetic material is attracted to the magnet. It attaches itself to the bar magnet. A mixture of common salt and iron fillings can be separated by magnetization method. (see Fig3.11 above.) A mixture of sulphur and iron filling can also be separated using this method.

11.13 Chromatography:

This method is used in separation of complex mixture. The complex mixture includes coloured substances. In this method there is mobile phase, which can be a liquid or gas while the fixed phase is a solid. Types of chromatographic techniques are:

- i. Paper Chromatography
- ii. Column chromatography
- iii. Thin-layer chromatography
- iv. Gas chromatography

Only the paper chromatography will be treated in this book.



A complex mixture of ink in a black bic can be separated using paper chromatographic method. A spot is made on an oblong strip of filter paper. The filter paper is mounted in a split cork in a boiling tube dipping the bottom end of the filter paper in a liquid (propanone or ethanol). The liquid moves or rises up the filter paper and dissolves the substances in the mixture (black ink). As the liquid rises up the paper, it carries the dissolved substances with it. Each substance is carried at a different speed so that the liquid separates out the components in the mixture. Chromatography is used mainly in the analysis of mixtures. a variant of paper chromatographic method is this:

If the solution containing the mixture to be separated is spotted on the filter paper, the components of the mixture move at different speeds and as a result become separated. The mixture may be placed on the filter paper, an appropriate solvent (ethanol) is dropped on the mixture. The components of the mixture then move with the solvent at various speeds and are thus separated.

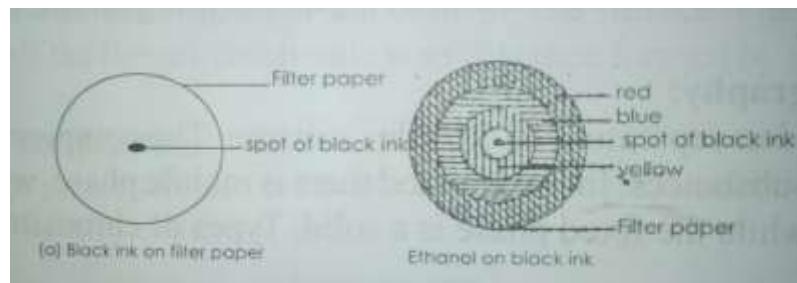


Fig 11.13 Separating the colours of black ink by chromatography

11.14 Use Of Centrifuge Machine:



Fig 11.14 Separation of solid from liquid using a centrifuge machine

This method is used to separate solid from liquid when the volume of the mixture is small. The mixture of solid and liquid is poured into a test tube and the test tube is placed inside the bucket of the centrifuge machine. As the centrifuge machine is put on, it moves the bucket and its content at a very great speed in a circular form. After sometime, the solid particles are thrown to the bottom of the test tubes while the liquid is on top. The top liquid can now be decanted from the solid.

11.15 Sieving:

This method is used in separating mixture in solid state. Separation using this method is dependent on the particle size of the solids and the size of the openings in the sieve. The following mixtures can be separated by sieving:

- i. Mixture of sand and beans
- ii. Mixture of sand and stone
- iii. Mixture of soya-beans and stone.

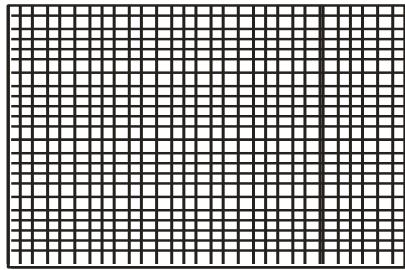


Fig 11.15 A sieve

11.16 Criteria For Purity:

After the separation of mixture, one may wish to know if the separated substances are actually pure. To find this out, one has to determine some physical constants of the separated substances like

- i. Melting points of solids
- ii. Boiling points of liquids
- iii. Refractive index or the relative density of both solids and liquids.

After finding these out, he has to compare the results with scientifically determined values. If any of the constants determined is the same with scientifically determined value, the substance is pure but if any of the constants determined differs from the established value, then the substance is not pure.

Exercises:

1. Write out ten physical methods of separating mixture.
2. What are the industrial application of filtration as a separation technique?
3. What are the main difference between the following Separation techniques?
 - i. Crystallization and fractional crystallization.
 - ii. Distillation and fractional distillation
4. Define these terms:
 - i. Sublimation
 - ii. Evaporation
 - iii. Precipitation

5. Briefly describe how a complex mixture of black ink can be separated in your school chemistry laboratory.

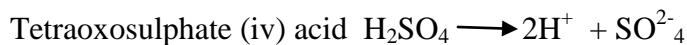
CHAPTER 12

ACIDS

12.1 Acids

Acid is a substance which when dissolved in water produces hydrogen ions (or protons) as the only positive ion.

Examples of acid are:



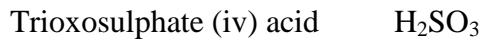
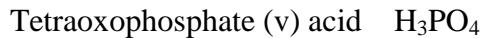
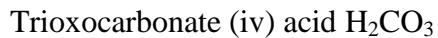
The process by which acids dissolve in water to produce hydrogen ions, H^+ , with the corresponding negative ions is known as ionization.

Strong and Weak Acids

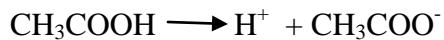
Strong acids ionize completely in water to give hydrogen ions and anions. There is complete ionization of hydrogen ion in these acids. The concentration of hydrogen ions is very high. Examples of strong acids are:



Weak acids are only partially ionized in water. There is incomplete ionization of hydrogen ions in these acids. The concentration of hydrogen ion is very low. Examples of weak acids are: -



Ethanoic acid has only 0.4% ionization in water.



12.2 Classes of Acids

The two classes of acids are: -

Organic acids and Inorganic acid or mineral acids. Organic acids occur in organic matter (plants and animals) while inorganic acids can be prepared from mineral elements or inorganic matter. The following are the organic acids:

Organic acid	Source
Ethanoic acid	Vinegar
Lactic acid	Milk
Fatty acids	Fats and oil
Amino acids	Proteins
Citric acid	Lime, lemon

The following are the inorganic acids:

Inorganic acid	Formula	Constituents
Hydrochloric acid	HCl	Hydrogen, chlorine
Tetraoxosulphate (vi) acid	H ₂ S0 ₄	Hydrogen, Sulphur, Oxygen

Basicity of acids:

Basicity of an acid is the number of replaceable hydrogen ions (H⁺) in one molecule of the acid.

12.3 Basicity of some Acids

Acid	Ions produced	Basicity
HCl	H ⁺ , Cl ⁻	1 or monobasic
HNO ₃	H ⁺ , NO ₃ ⁻	1 or monobasic
H ₂ SO ₄	2H ⁺ , SO ₄ ²⁻	2 or dibasic
H ₂ CO ₃	2H ⁺ , CO ₃ ²⁻	2 or dibasic
H ₃ PO ₄	3H ⁺ , PO ₄ ³⁻	3 or tribasic

Ethanoic acid react with sodium metal to form sodium ethanoate.



Although ethanoic acid has four hydrogen atoms per molecule of acid, only one of them is replaceable by a metal. Ethanoic acid is, therefore said to be monobasic.

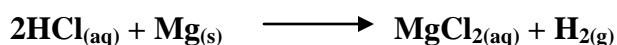
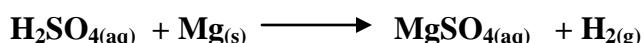
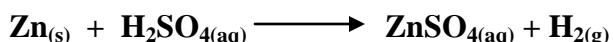
12.4 Physical Properties of Acids

1. Acids turn wet blue litmus red,
2. Concentrated forms of strong acids are corrosive
3. Dilute acids have sour taste. Vinegar and unripe fruit have sour taste due to the presence of acid in them.

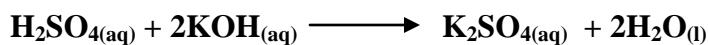
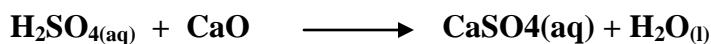
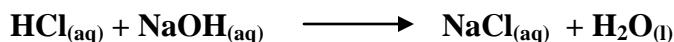
N.B: When preparing a dilute solution of acid, always add concentrated acid to water. Never add water to concentrated H_2SO_4 acid because, the hot acid solution may splatter unto your face and body and cause serious burns.

12.5 Chemical Properties of Acids

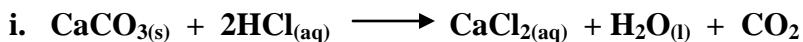
1. With the exception of trioxonitrate (v) acid, acids react with some metals like zinc, magnesium and iron to liberate hydrogen gas.



2. Acids react with bases to form salt and water only. This reaction is known as a neutralization reaction.

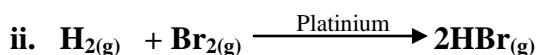
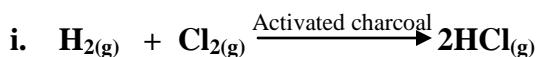


3. Acids reacts with trioxocarbonate(iv) salts to liberate carbon(iv)oxide



12.6 Preparation of Acids:

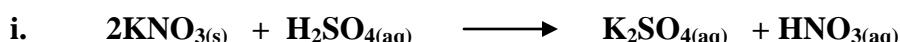
Acid is prepared by direct combination of the constituent elements in the presence of a catalyst. The gas formed is then dissolved in water.



2. Acid is prepared by the dissolution of an acid anhydride in water. An acid anhydride is a non-metallic oxide which dissolves in water to form acid. Some examples of acid anhydride are: Carbon(iv)oxide, Sulphur(iv)oxide, Sulphur(vi)oxide etc.



3. Acid is prepared by the displacement of a weak acid or a volatile acid (from its salt) by a strong acid.



Uses of Acids

1. They are used in the manufacture of drugs, fertilizers, soaps and detergents.

2. Industrially, some acids are used as:

i. Drying agents

ii. Oxidizing agents

iii. Catalysts

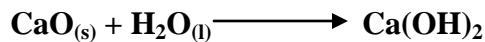
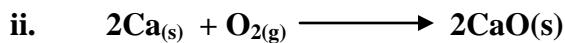
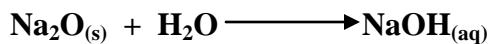
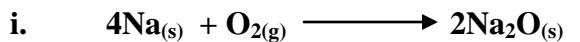
12.7 Bases

A base is any oxide or hydroxide that dissolves in water to yield hydroxide ion, OH^- , as the only negative ion.

Some common metallic oxides and hydroxides which are bases are:

Base	Formula	Solubility
Sodium oxide	Na ₂ O	Soluble
Calcium oxide	CaO	Slightly soluble
Potassium oxide	K ₂ O	Soluble
Sodium hydroxide	NaOH	Soluble
Calcium hydroxide	Ca(OH) ₂	Slightly soluble
Potassium hydroxide	KOH	Soluble
Zinc hydroxide	Zn(OH) ₂	Insoluble

The metallic oxides are formed when the metals burn in air and they react with water to form their corresponding hydroxides:



12.8 Alkali

Alkali is a base which is soluble in water.

Examples of alkali



Ionization of bases

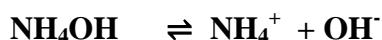


12.9 Strong And Weak Bases

A strong base ionizes completely in aqueous solution to produce negatively charged hydroxide ions, OH^- , and the positively charged metallic ion. Examples of strong alkali are:



A weak base is a metallic oxide or hydroxide which is incompletely ionized in aqueous solution to produce relatively few ions. Examples of a weak base are:

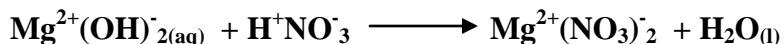
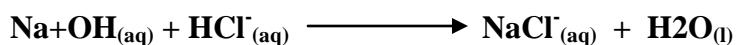


12.10 Physical Properties of Bases/Aalkalis

- i. Concentrated form of the alkalis sodium hydroxide and potassium hydroxide, are corrosive.
- ii. Bases turn wet red litmus paper blue
- iii. Bases have a slippery or soapy feel.
- iv. Bases have a bitter taste.

12.11 Chemical Properties of Bases/Aalkalis

Neutralization reaction: Bases react with acids to form salt and water only.



Hydrogen ion, H^+ , from the acid neutralized by the hydroxide ion, OH^- , from the base to base to form the water molecules

Neutralization reaction can be defined as the process in which acid react with a base or an alkali to form salt and water only.

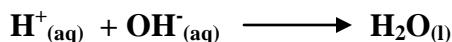
Or

Hydrogen ion, H^+ , from the acid is neutralized by the hydroxide ion, OH^- , from the base to form the water molecules

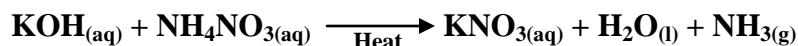
Neutralization reaction can be defined as the process in which acid react with a base or an alkali to form salt and water only.

Or

Neutralization is the combination of hydrogen ions, H^+ , and hydroxide ions, OH^- , to form water molecules. A salt is also formed.



2. Bases react with ammonium salt to liberate ammonia gas.

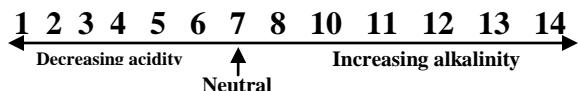


12.12 Uses Of Alkalies

They are used in the manufacture of soap, glass, paper. Some alkalis are used in the softening of hard water.

12.13 Measurement Of Acidity And Alkalinity (the pH Scale)

pH is a negative logarithm of hydrogen ion concentration, or pH is $-\log [\text{H}^+]$. pH is a measure of hydrogen ion, H^+ , concentration. The number 7 on the pH scale represents the condition of exact neutrality. Numbers less than 7, that is, pH 6,5,4 etc indicate increasing acidity as the number decreases. Numbers greater than 7, that is, pH 8,9,10 etc indicate increasing alkalinity as the number increases.



The pH value of a colourless solution can be obtained by adding a universal indicator or by spotting the liquid on a universal indicator paper. In both cases, a colour will appear from which the pH of the liquid can be decided.

Colour	pH Number	Acid/Base
Red	1-3	Very acidic
Orange	4-5	Weak acid
Yellow	6	Very weak acid
Green	7	Neutral
Blue	8	Very weak base
Indigo	9-10	Weak base
Violet	11-14	Very basic

Table 12.1 pH range and colour changes of universal indicator

Calculation of pH

Example 1

Find the pH of 0.1 mole NaOH



$$[\text{H}^+] [\text{OH}^-] \longrightarrow K_w \text{ (ionic product of water)}$$

$$[\text{H}^+] [\text{OH}^-] = 10^{-14}$$

$$[\text{H}^+] [10^{-1}] = 10^{-14}$$

$$[\text{H}^+] = \frac{10^{-14}}{10^{-1}} = 10^{-14} \times 10^{-1} = 10^{-13}$$

$$\text{pH} = 13$$

Example 2

Calculate the pH of 0.01M of KOH

$$[\text{H}^+] [\text{OH}^-] \longrightarrow K_w \text{ (ionic product of water)}$$

$$[\text{H}^+] [\text{OH}^-] = 10^{-14}$$

$$[\text{H}^+] [10^{-2}] = 10^{-14}$$

$$[\text{H}^+] = \frac{10^{-14}}{10^{-2}} = 10^{-14+2} = 10^{-12}$$

pH of the alkaline is 12

Example 3

Calculate the Ph of 0.1M HCl

$$[\text{H}^+] [\text{OH}^-] = 10^{-14}$$

$$[10^{-1}] [\text{OH}^-] = 10^{-14}$$

$$[\text{OH}^-] = \frac{10^{-14}}{10^{-1}} = 10^{-14} \times 10^{-1} = 10^{-13}$$

pH of the acid is $14 - 13 = 1$

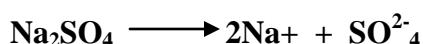
12.14 What is a Salt?

A salt is a compound consisting of positive metallic ion and a negative ion derived from an acid.

e.g. NaCl; Sodium chloride

Na₂SO₄; Sodium tetraoxosulphate(vi)

ZnCl₂; Zinc chloride

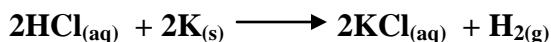


ii. A salt is the compound formed when all or part of ionizable hydrogen of an acid is replaced by metallic or ammonium ions.

Eg. NH₄Cl: Ammonium Chloride

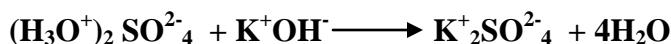
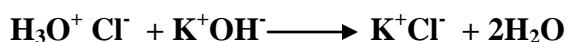
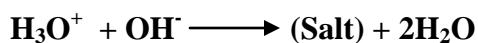
NaHSO₄; Sodium hydrogen tetraoxosulphate(vi)

Ca(NO₃)₂; Calcium trioxonitrate(v)



iii. Most salts, are products of neutralization reaction.

E.g Acid + Base \longrightarrow Salt + Water



12.15 Types Of Salts

Five main types of salts are:

Normal salt, Acid salt, Basic salt, Double salt, and Complex salt.

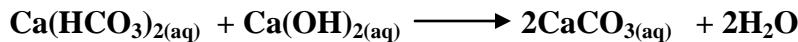
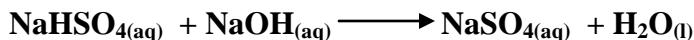
Normal salt: Normal salt is formed when all the replaceable hydrogen ions in an acid have been completely replaced by metallic ions of a base.

Metallic ions of a base.

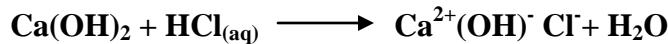
E.g NaCl, CaCl₂, MgSO₄, Na₂SO₄. A normal salt contains no replaceable hydrogen, H⁺ or hydroxyl, OH⁻, ion. Aqueous solution of normal salt is neutral to litmus and it has a pH of 7.

Acid salt: An acid salt is formed when the replaceable hydrogen ions in the acid are partially replaced by a metal. An acid salt contains replaceable hydrogen ion E.g NaHCO₃; sodium hydrogen trioxocarbonate(iv). NaHSO₄; Sodium hydrogen tetraoxosulphate(vi) Ca(HCO₃)₂; Calcium hydrogen trioxocarbonate(iv).

A solution of an acidic salt turns blue litmus paper red, and it can react with excess alkali to form a normal salt and water. Eg.



Basic salt: A basic salt is formed when a base reacts with a limited quantity of an acid. That is, the oxonium ion, H₃O⁺, is insufficient for the complete neutralization of the base. A basic salt therefore contains replaceable hydroxyl ions, OH⁻.



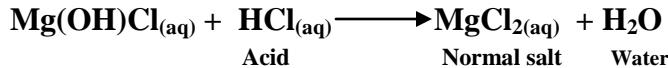
Examples of basic salt are:

Ca(OH)Cl; Calcium hydroxide chloride ^{Basic salt}

Zn(OH)Cl; Zinc ii hydroxide chloride

Mg(OH)Cl; Magnesium (ii) hydroxide chloride

A basic salt turns red litmus blue and reacts with excess acid to form a normal salt and form a normal salt and water.



Double salt: Double salt is a salt which ionizes to yield three different types of ions in solution. Two of them are positively charged (metallic ions) while the other is negatively charged. Eg

(NH₄)SO₄ FeSO₄ .6H₂O Ammonium iron(ii) tetraoxosulphate(vi) hexahydrate

KAl(SO₄)₂ .12H₂O Ammonium potassium tetroxosulphate (vi) dodecahydrate

Complex salt: This is a salt that contains complex ions. A complex ion is the ion consisting of a charged group of atoms. Eg

$\text{Na}_2\text{Zn}(\text{OH})_4$: Sodium tetrahydroxozinate(ii)

$\text{K}_4\text{Fe}(\text{CN})_6$: Potassium hexacyanoferrate(ii)

The complex salts ionize as follows to produce the complex ions:

$(\text{Zn}(\text{OH})_4)^{2-}$ Tetrahydroxozincate (ii) ion

and

$(\text{Fe}(\text{CN})_6)^{4-}$ Hexacyanoferrate (ii) ion



12.16 Preparation of Soluble Salts

Note: A method adopted for the preparation of a particular salt is dependent on two factors. These are:

1. The solubility of the salt in water
2. The stability of the salt to heat.

The methods used in the preparation of soluble salts are:

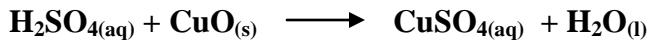
- i. Reaction of an acid with a metal which is more reactive than hydrogen e.g Na, Ca, Zn, Mg and Fe.



- ii. Neutralization reaction: Reaction between an alkali and acid



- iii. Neutralization reaction: reaction between a hot dilute acid and an insoluble base.



- iv. Reaction between a dilute acid and trioxcarbonate (iv)

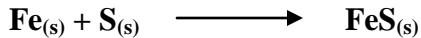
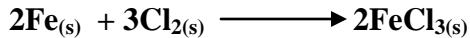


The soluble salts can then be recovered from the solution by either evaporation to dryness or crystallization.

Preparation of Insoluble Salts

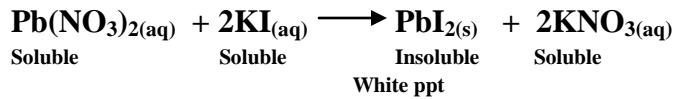
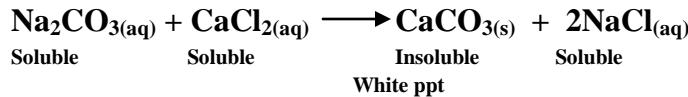
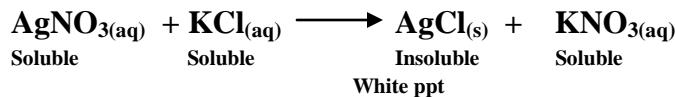
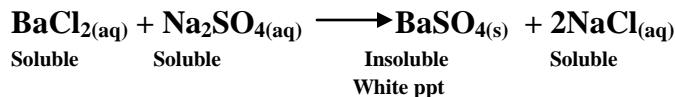
Insoluble salt can be prepared by:

- Direct combination of the constituent elements.

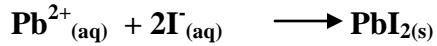
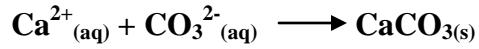
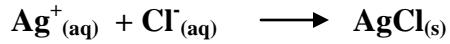
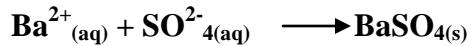


- Double decomposition reaction.

Two appropriate insoluble salts react to precipitate the required insoluble salt. The latter is then filtered, washed and dried. These insoluble salts that can be prepared by this method are; PbCl_2 , PbI_2 , CaCO_3 , AgCl and BaSO_4



Above reactions can be represented ironically as follows:



Soluble salts	Insoluble salts
1. All salts of potassium, sodium and ammonium	
2. All trioxonitrate(v) of metals	
3. Most common chlorides	Silver chloride, mercury(ii)chloride, and lead(ii)chloride
4. Potassium trioxocarbonate(iv), sodium trioxocarbonate(iv) and ammonium trioxocarbonate(iv)	Most common trioxocarbonate(v) of metals
5. Most common tetraoxosulphate(vi) of metals	Barium tetraoxosulphate(vi) and lead(ii) tetraoxosulphate(vi) (calcium tetraoxosulphate(vi) is sparingly soluble)

12.17 Water of Crystallization

There are two kinds of salts.

They are hydrated salts and anhydrous salts.

Hydrated salts: are salts that contain water of crystallization. Anhydrous salts: are salts that contain no water of crystallization.

Water of crystallization: is a definite or fixed amount of water associated with salts as they come out of solution.

Although all hydrated salts (hydrates) are crystalline, not all crystalline salts are hydrates, eg NaCl, KMnO₃.

Examples of hydrated salts are CuSO₄. 5H₂O Copper(ii)tetraoxosulphate(vi) pentahydrate.

FeSO₄. 7H₂O Iron(ii)tetraoxosulphate(vi) heptahydrate

Na₂CO₃. 10H₂O Sodium trioxocarbonate(iv)decahydrate

Other hydrates are; Zn(NO₃)₂. 6H₂O; MgSO₄. 7H₂O; ZnSO₄. 7H₂O; Cu(NO₃)₂. 3H₂O

When the hydrated salts loose their water of crystallization as a result of heating, they become anhydrous or amorphous (shapeless).

Some crystalline salts which are not hydrates are; NaCl, KNO₃, (NH₄)₂SO₄, KMnO₄ etc.

12.18 Efflorescent; Deliquescent and Hygroscopic Substances or Salts;

Certain crystals (salts) containing water of crystallization gradually loose all or part of it, when they are exposed to the air and become anhydrous salts. Such salts are said to be efflorescent and the phenomenon is known as efflorescence. Examples of efflorescent salts are:

$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ Sodium troxocarbonate (iv) decahydate.

$\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ Sodium tetraoxosulphate (vi) decahydate.

Some substances or compounds, when exposed to the air absorb enough water from the air to form solution. Such substances are called the deliquescent substances and the phenomenon is known as deliquescence. Examples of deliquescent compounds are:

NaOH Sodium hydroxide

KOH Potassium hydroxide

CaCl_2 Calcium chloride

FeCl_3 Iron(iii) chloride

MgCl_2 Magnesium chloride

P_2O_5 Phosphorus(v)oxide

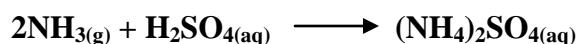
Some substances when exposed to the air can absorb water but without dissolving in it. Such substances are said to be hydroscopic. Examples of hydroscopic substances are: CaO , CuO , NaNO_3 and Conc. H_2SO_4 . Concentrated H_2SO_4 absorbs water and becomes diluted.

12.19 Uses Of Hydroscopic and Deliquescent Substances

They are used as drying agents in the desiccators and in the laboratory for drying gases.

Drying agents	Gases
Fused CaCl_2	All gases except NH_3
CaO	NH_3
P_2O_5	All gases except NH_3
Conc. H_2SO_4	All gases except NH_3 and H_2S
Silica gel	All gases

Conc. H_2SO_4 cannot be used in drying NH_3 gas because the gas can react with it.



Exercises

1. What do you understand by these terms:
 - a. Acid (b) Strong acid (c) Weak acid (d) Organic acid (e) Inorganic acid. Write two examples of each term
2. Why is it not proper to add water directly to concentrated tetraoxosulphate(vi)acid when diluting it?
3. What are the two physical properties of acids?
4. What are the two chemical properties of acids?
5. Briefly explain two methods of preparing acids
6. What are the industrial importance of acids?
7. What do you understand these terms to mean?
 - (a) Base (b) Alkali (c) Strong base (d) Weak base

Write an example of each

2. Write the ionization products of these bases
 - a. Mg(OH)_2
 - b. Cu(OH)_2
 - c. Na_2O
8. Write two physical properties of bases.
9. Write one chemical properties of a base
10. (a) What is pH?
 - (a) Calculate the pH value of 0.01M potassium hydroxide.
11. What is a neutralization reaction?

Illustrate with an example
12. Define these terms and give an example of each.
 - a. Salt (b) Normal salt (c) Acid salt (d) Basic salt (e) Double salt and (f) Complex salt

13. Use chemical equations only to represent two methods of preparing soluble salt.
14. Use chemical equations to show how insoluble salts can be prepared in the science laboratory.
15. What do you understand these terms to mean? Write an example of each.
 - (a) Hydrated salt (b) Anhydrous salt (c) Efflorescent salt (d) Deliquescent compound
 - (d) Hydroscopic substance.
16. Can concentrated H_2SO_4 be used in drying ammonia gas? Why?

CHAPTER 13

WATER, HARDNESS OF WATER

13.1 Sources of Water

Water is referred to as universal solvent. It is a solvent that can dissolve many substances. Natural water is the water that is found in its natural state, while treated water is the one that has received some form of treatment. Examples of natural water are; rainwater, well water, spring water, river water, sea-water and lake water.

Sources of water include: Rain, spring, well, river, lake and sea.

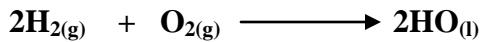
Rainwater is the purest form of natural water because it is formed as a result of the condensation of water vapour in the atmosphere. It is free from mineral salts. It is soft and lathers easily with soap. Rainwater sometimes dissolve some impurities from the polluted environment (atmosphere) and this makes it to be hard. Spring water is the rainwater that sank through the porous soil layers and collected above an impervious layer. Sometimes some of this water emerge again on the soil surface as spring water. Spring water is not a good source of drinking water because it contains some dissolved mineral salts, dust and disease causing germs such as bacteria and fungi.

Well-water is a stagnant water that collects on the top of the impervious layers of the soil after raining. It is reached by digging deep holes known as ‘wells’. It is impure for drinking purposes because it contains dissolved clay, mineral salts and remains of dead organic matters. River water is formed by coming together of springs and other running waters. It contains a lot of dissolved mineral salts, air and remains of dead organic matters (plants and animals).

Lakes and seas form a reservoirs for rivers and other running waters. Water from these two sources is not good for drinking because they contain a lot of impurities due to the accumulation of impurities from the various sources of water.

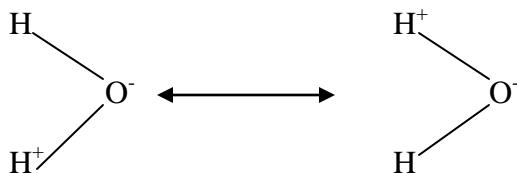
13.2 Composition of Water:

It is made up of hydrogen and oxygen in the ratio of 2:1 by volume. It is formed when hydrogen is burnt in oxygen.



Water is a binary compound and it is V shaped in structure. Some sources of water supply are ; spring, rain, river, well, lake, sea, ocean. Water is a universal solvent because it can dissolve as many substances as possible. It dissolve more substances than any other solvent. The presence of dissolved substances (solids) in water can easily be shown by evaporating tap water on a watch

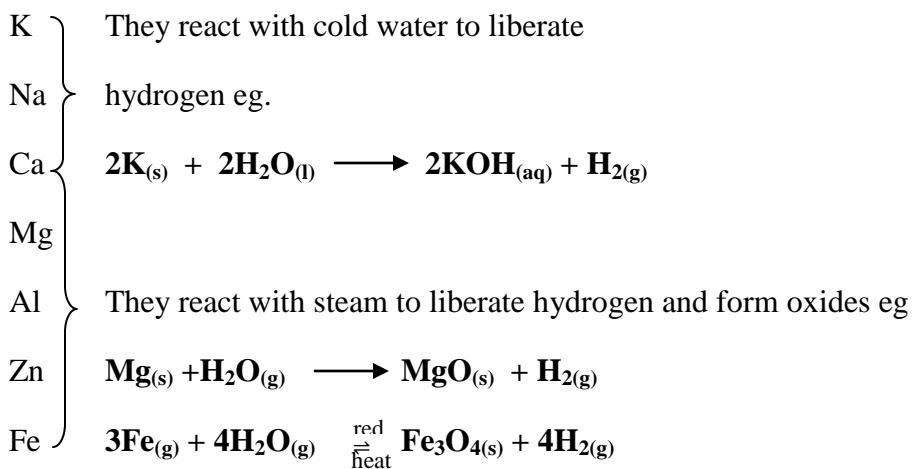
glass. Water is a polar solvent. Its molecular structure is made up of both positive and negative ends. It can therefore ionize into positive and negative ions.



Water is a solvent for electrovalent solutes such as mineral salts, mineral acids and bases. It is also an excellent solvent for covalent substances which posses the hydroxyl group, - OH. Examples of such substances are propan - 1, 2- diol, propan 1, 2, 3 - triol, alcohol, sugar. All gases can dissolve in water to some extent. Examples are ammonia, nitrogen (iv) oxide, hydrogen chloride, sulphur (iv) oxide, carbon (iv) oxide, chlorine, oxygen, nitrogen etc.

13.3 Properties of Water:

- Pure water is a colourless, odourless and tasteless liquid. It is neutral to litmus. Its boiling and freezing points are 100°C and 0°C respectively.
- Water reacts with metals in various ways depending on their position in the reactivity series.

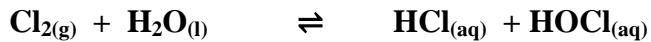


Pb
 Cu
 Hg
 Ag
 Au

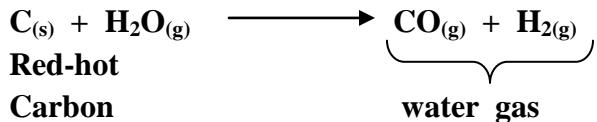
They neither react with cold water nor steam

3. Water reacts with some non-metals

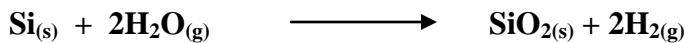
i. Chlorine reacts with water to form a mixture of two acids - ie hydrochloric acid and oxochlorate (i) acid respectively.



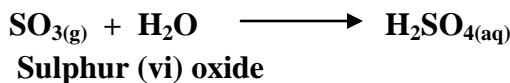
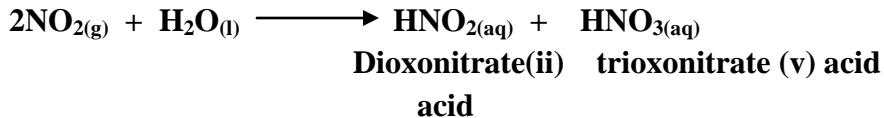
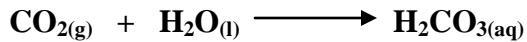
ii. Red-hot carbon, reacts with steam to form water gas (ie a mixture of hydrogen and carbon monoxide).



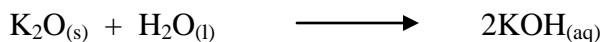
iii. Silicon at white heat react with steam to form silicon (iv) oxide and hydrogen



4. Some non-metallic oxides react with water to form acids.



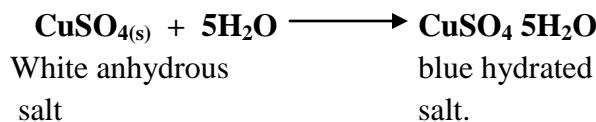
5. Some metallic oxides (eg. K₂O, Na₂O, CaO) react with water to form their hydroxides.



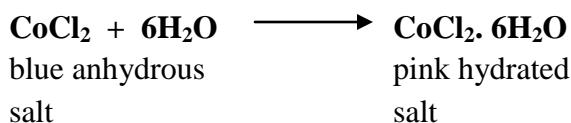
Test for water:

We test for water using anhydrous copper (ii) tetraoxosulphate (vi) ie CuSO_4 and Cobalt (ii) Chloride paper

1. Water turns white anhydrous copper (ii) tetraoxosulphate (vi) to blue.



2. Water turns blue cobalt (ii) chloride paper pink.



13.4 Hardness Of Water and Its Causes:

A hard water is a water that cannot lather easily with soap. Hard water is caused by the presence of soluble metallic salts. The most common of these are hydrogen trioxocarbonate (iv), tetraoxosulphate (vi) and chlorides of calcium and magnesium. Hard water cannot produce lather easily because the salt in water causing the hardness react chemically with soap. The cleansing dirty scum is produced. Lather will not form until all the salts are removed in form of scum.

Types of hardness of water: Two types of hardness of water are: Temporary hardness and Permanent hardness.

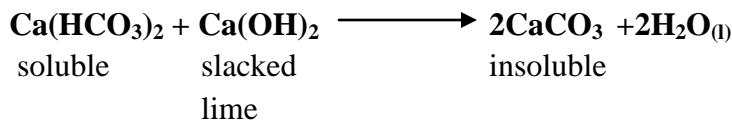
Temporary hardness: Temporary hardness is that hardness of water which can be removed or softened by boiling. Temporary hardness is caused by hydrogen trioxocarbonate (iv) of calcium or magnesium ie $(\text{Ca}(\text{HCO}_3)_2$ or $\text{Mg}(\text{HCO}_3)_2$.

When the water containing calcium or magnesium hydrogen trioxocarbonate (iv) is heated, the acidic salt decomposes into insoluble calcium or magnesium trioxocarbonate (iv) and carbon (iv) oxide.



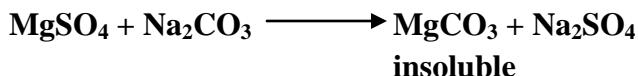
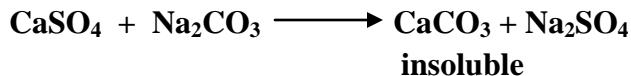
CaCO_3 is insoluble and cannot cause hardness of water.

Temporary hard water can also be removed by adding slacked lime ie $\text{Ca}(\text{OH})_2$.

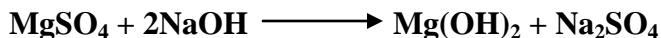


Permanent hardness: Permanent hardness of water is that hardness which cannot be removed by boiling. It is caused by dissolved calcium tetraoxosulphate (vi) or magnesium tetraoxosulphate (vi) ie CaSO_4 or MgSO_4 . These salts causing the hardness are not decomposed by boiling. The permanent hardness of water can therefore be removed by the addition of chemicals eg.

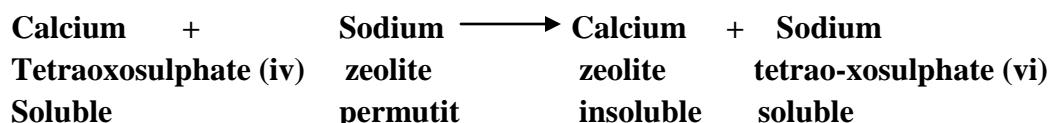
- i. Addition of washing soda: $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$



- ii. Addition of caustic soda: NaOH



- iii. Addition of permutit or zeolite: This is also known as ionexchange process or permitit method. The ion exchanger contains sodium aluminium trioxosilicate (iv) ie NaAlSiO_3 . When the hard water passes through a resin container, the sodium ions enter the solution while the unwanted calcium and magnesium ions take their place in a double decomposition reaction.



The common name for NaAlSiO_3 is sodium zeolite or sodium permutit.

Advantages of hard water:

It contains calcium compound which is essential for our teeth and bones. Hard water has pleasant taste unlike pure water which is tasteless.

Disadvantages of hard water:

Hard water makes the dyeing of cloths difficult.

It forms scum with soap and wastes soap.

It forms fur or scale deposit in kettles and boiler pipes.

Water cycle:

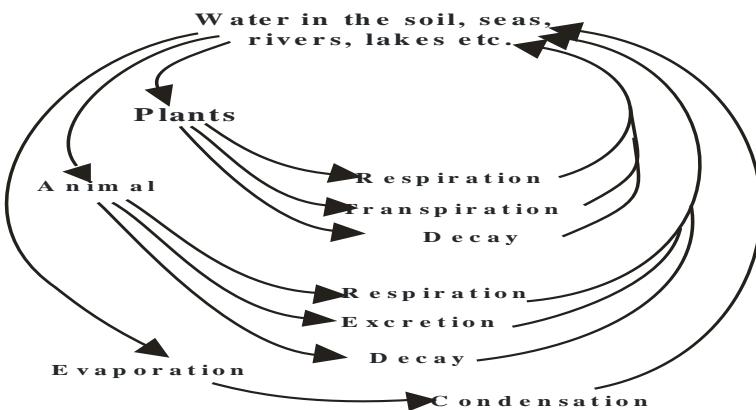


Fig. 13.1: Water cycle

Water cycle is the way water circulates in nature.

13.5 Treatment of Town Water Supply

Some examples of treated water are: distilled water, pipe-borne water for township and chlorinated water for use in swimming pools. Treated water is prepared for special purposes.

Pipe-borne water can be purified using the following methods:

Coagulation, sedimentation, filtration and disinfection.

Coagulation: This is a method whereby chemical known as coagulating substances (such as silica gel) is added to the water which has been already tapped from the sources, to remove any colloidal particles.

Filtration: In this method, organic matter is separated from the water.

Sedimentation: Sediments which are formed during the coagulation are separated from the bore hole water.

Disinfection: After all these, chemicals are added to the water so that the water will be free from germs. The chemicals used include: silica of soda, hydrated lime, liquid chlorine, alum etc.

Exercises:

1. List the components of water.
2. Write five sources of water.
- 3a. List six examples of natural water.
 - b. Which of them is considered to be purest form? Why?

- c. What makes the rainwater to be hard?
- 4a. Why is water called a universal solvent?
 - b. Draw the molecular structure of water.
- 5. Write two chemical properties of water.
- 6a. What is hard water?
 - b. Write one method of softening:
 - (i) Temporary hard water. (ii) Permanent hard water.
- 7. Write (a) two advantages of hard water.
 - (b) two disadvantages of hard water.
- 8. Briefly describe one method of purifying water.

CHAPTER 14

CARBON AND ITS ALLOTROPES

14.1 Carbon

Carbon is generally a non-metallic black solid. It occurs as diamond, graphite, coal. It occurs as wood, petroleum, natural gases, metallic trioxocarbonates (iv) and carbon (iv) oxide in the combined form. Carbon is the first four (4) element in the periodic table. It has the atomic number of six (6). Its electronic configuration is $1s^2 2s^2, 2p^2$. It has the atomic mass of twelve (12). Carbon is contained in all living cells. It is an important source of heat and light energy.

14.2 Allotropes of Carbon

Allotropy is the existence of an element in two or more structural form in the same physical state. Carbon, tin, sulphur and phosphorous exist in different structural forms in their free states. Carbon exists in two main crystalline allotropic forms called diamond and graphite. It also exist in a non crystalline allotropic form known as amorphous carbon.

14.3 Diamond

Diamonds is found in Africa, India, Brazil, Republic of Guyana etc.

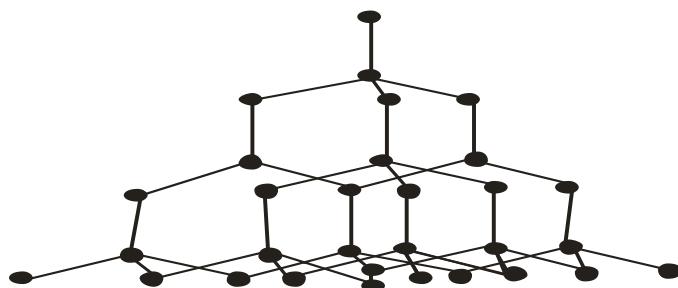


Fig 14.1 structure of diamond

It has a network structure. Diamond is octahedral in shape. When cut and polished, it has an amazing luster which makes it valuable as jewelry. It is the hardest substances known. It is therefore used in making glass cutters and rock borers. It is also used as dies for drawing metal wires.

In diamond each carbon is covalently bonded to four other carbon atoms at tetrahedral positions about it. Thus, since each of the four valence electrons is involved in a strong covalent bond, there is no free electron to conduct electricity. Diamond is therefore a poor conductor of electricity. It has a high melting point. It is resistant to high temperature and chemical attack. It sublimes at 3500°C .

14.4 Graphite:

Graphite is found in West Germany, Madagascar, Mexico, China, Australia etc.

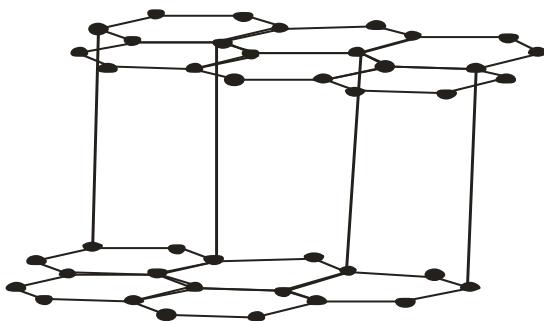


Fig 14.2 Structure of graphite

Graphite occurs naturally as an opaque black solid called plumbago. It has a layer structure. One layer solid slides over another. Because it is soft and slippery, it is used a lubricant. In graphite each carbon atom is bonded to three other carbon atoms and the fourth electron is free or delocalized. It is free to move round the whole structure. Thus, because one electron from each atom in graphite structure is mobile, graphite is a good conductor of electricity. It is inert, and it is often used as electrodes in electroplating and in dry cells. A mixture of graphite and clay is used as 'lead' in pencils. Graphite has a high melting point.

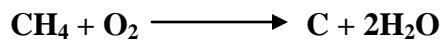
14.5. Amorphous Carbon

Non crystalline allotropic forms of carbon are called amorphous carbon. Amorphous forms of carbon include the following:

Coal, coke, carbon black, lamp black or soot, charcoal and carbon fibres.

Carbon black, lamp black or soot:

Carbon black is formed when natural gas (methane) and petroleum products is burnt in a limited supply of air. Lamp black is formed from an incomplete combustion of fuel (kerosene). It is found in the chimney of a kerosene lamp. Soot as formed in the kitchen chimney or above fire place where wood is used as a source of fuel. Carbon black, lamp black and soot are black and powdery in nature. They are used in making printers' ink, shoe polish and carbon papers.



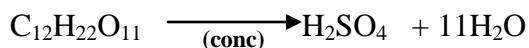
Large quantity of carbon black can be prepared by heating natural gas to a very high temperature in the absence of air.



Carbon fibres:

Carbon fibres are formed or produced when fibres are heated to a very high temperature in the absence of air. The fibres char to form carbon. Some examples of fibre materials are: Nylon 6,6, Terylene, Polyester, Rayon etc.

Charcoal: different types of charcoal are wood charcoal, animal charcoal and sugar charcoal. Wood charcoal is prepared or formed when wood is heated in the absence or limited supply of air. Animal charcoal is formed when bones and animal refuse are heated in the absence or limited supply of air. Sugar charcoal is made by dehydrating sugar with concentrated H₂SO₄ or by heating the sugar in a limited supply of air.

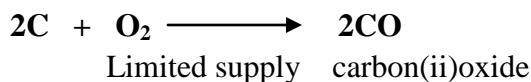


Wood charcoal contains sulphur as an impurity while animal charcoal contains calcium tetraoxophophates(v) as an impurity. Charcoal has a high adsorptive capacity. Not all gases are however adsorbed by charcoal. Carbon(iv) oxide is one of the gases that cannot be adsorbed by charcoal or CO₂ cannot be adhered to the internal surfaces of charcoal.

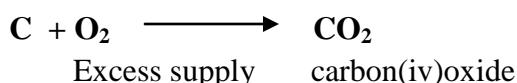
14.6 Chemical Properties of Carbon

1. Carbon reacts with oxygen to form its oxide.

When carbon burns in a limited supply of air, a toxic gas called carbon(ii)oxide is formed.



If oxygen is in excess supply, carbon(iv)oxide is formed.

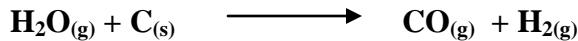
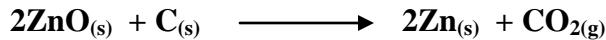
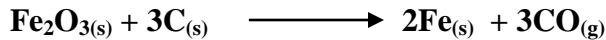


If diamond, graphite and amorphous carbon are heated separately in an excess, supply of air, the same carbon(iv)oxide is formed.

2. Carbon reacts with some elements to form its binary compounds.



3. Carbon reduces some metallic and non-metallic oxide to their respective metals and non-metals.



4. Carbon is oxidized to carbon(iv)oxide by strong oxidizing acids.



14.7 Coal and Coke

Coal and coke are among the examples of amorphous carbon as already mentioned.

Coal:

Coal is a black impure carbon formed from vegetations buried under the earth for many millions of years. These plants under high pressure underwent metamorphosis into the hard black rock, which are now known as coal.

Uses of coal

In the home, it is used as a domestic fuel. It is used for cooking. It is used in power stations for generating electricity. Railway locomotive engines run on coal. Coal is used in making or manufacturing many chemicals and drugs.

Types of coal

Four main types of coal are:

1. Peat
2. lignite (brown coal)
3. Bituminous coal (soft coal)
4. Anthracite (hard coal)

Peat is not a true coal. It is the first stage of coal formation. Peat consists of about 50% carbon.

Lignite (brown coal) is a further stage in the process of coal formation. It consists of about 60% carbon.

Bituminous coal (soft coal) is the most plentiful of the four types of coal. It consists of about 80% carbon.

Anthracite (hard coal) is the final product of the coal making process.

. It consists of over 90% carbon. Anthracite is the oldest and hardest form of coal.

Coal formation: Peat → Lignites → Bituminous → Anthracite

Destructive distillation of coal

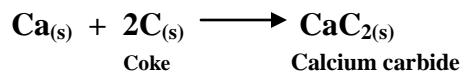
Destructive distillation of coal is heating coal to a very high temperature (about 1,250°C) in the absence of air so that it decomposes into four main products namely: coal gas, coal tar, ammonical liquor and coke. These four products are known as the products of destructive distillation of coal. The first three products are volatile. The only residue left after they have all been distilled off (and collected at different temperatures) is coke. Coal is used for the generation of heat energy for domestic purposes and chemical industries.

Coal gas is a mixture of these gases: ethane, methane, carbon (ii) oxide, hydrogen, sulphur (iv) oxide (very small) and hydrogen sulphide (very small). It is used as fuel in industries. Coal tar is a thick, brownish black liquid. It is a mixture of many organic chemicals like toluene, naphthalene, benzene, phenol etc. It is used in the production of many chemicals such as disinfectants, dyes, drugs etc. Ammonical liquor is made up of ammonia. It is used in the making of ammonium compounds. For instance ammonium tetraoxosulphate(vi), NH_4SO_4 , which is manufactured with this liquor is used as a fertilizer.

Coke

It has already been pointed out that coke is the only residue left after all the volatile substances had been distilled off after the destructive distillation of coal. Coke is a light, porous, black carbon. Coke is amorphous in that it contains about 90% amorphous carbon. It is used in the making of calcium carbide used in the production of ethyne.

Distillation of coke: Coke is a light, porous, black carbon. Coke is amorphous in that it contains about 90% amorphous carbon. It is used in the making of calcium carbide used in the production of ethyne.



Coke is used in the extraction of metals from their ores. It reduces the oxides of the less active metals (e.g ZnO, Fe₂O₃, CuO, PbO etc) to the metals.

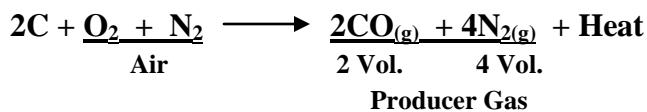


Coke is also used as fuel.

Gasification of coke

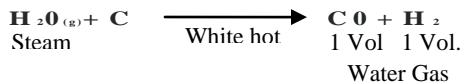
Coke can be converted or turned into gas by making it to combine with either the oxygen of the air or with steam.

Producer gas: Producer gas is formed when air is passed into thick layer of red-hot coke.



Producer gas is therefore a mixture of carbon (ii) oxide and nitrogen in the ratio of 1:2

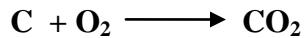
Water gas: Water gas is formed when steam is blown over white-hot coke.



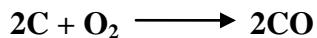
Water gas is a mixture of carbon (ii) oxide and hydrogen in the ratio of 1:1. Both producer gas and water gas are used as fuel in heating. Water gas has more calorific value than producer gas because both carbon(ii)oxide, CO and hydrogen, H₂ are combustible. Town gas is a name often given to a mixture of coal gas and water gas.

14.8 Oxides of Carbon

When carbon is heated in air it forms two types of oxides. They are: carbon (iv) oxide, CO₂ and carbon(ii)oxide, CO. If oxygen is in excess supply, CO₂ is formed. If the oxygen is in limited supply CO is formed.



Excess supply



limited supply

Carbon (iv) oxide is one of the naturally occurring compound of carbon. It is one of the gases that make up the air around us. Air contains 0.03% of CO₂. Carbon(iv)oxide does not support

combustion. It is important to green plants for making their food by the process of photosynthesis.

14.9 Laboratory Preparation of Carbon (iv) oxide

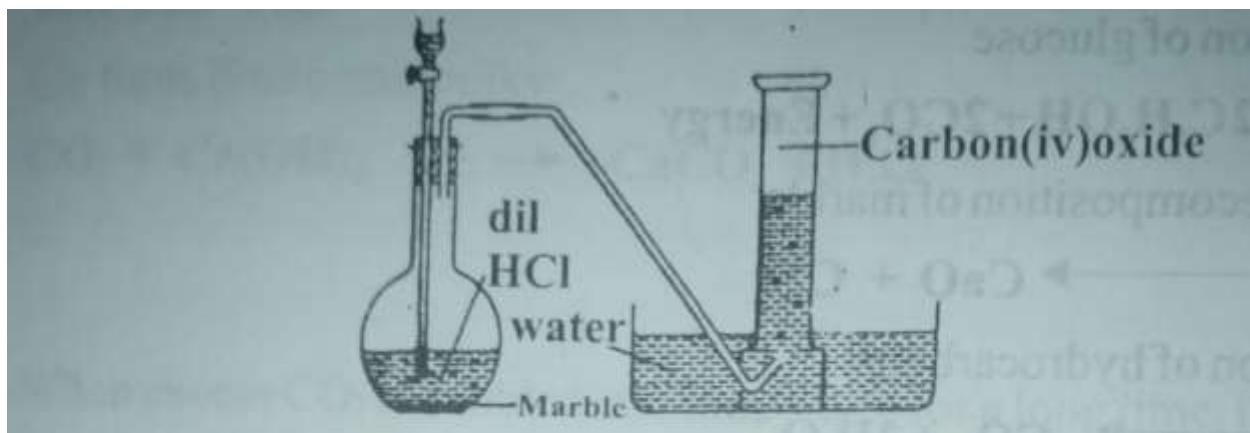


Fig 14.3(a) upward delivery or downward displacement of air over water

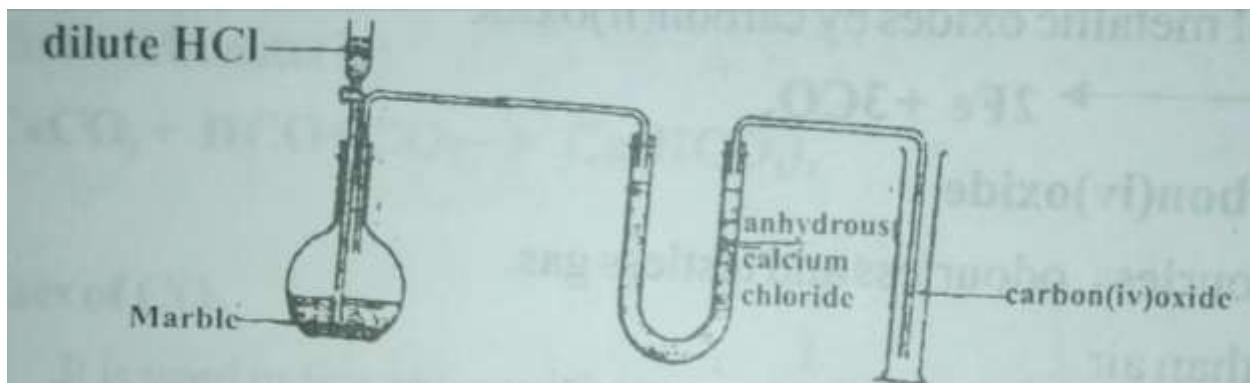
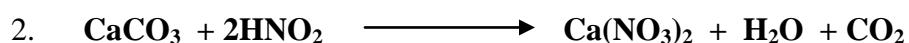
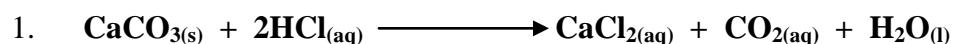
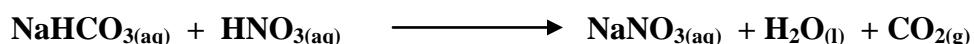


Fig 14.4 (b) downward delivery of upward displacement of air.

Carbon(iv)oxide or carbon dioxide is prepared in the laboratory by the action of dilute HCl or HNO₃ on marble or limestone.



it is also prepared by the action of the above named acids on hydrogen trioxocarbonate(iv) of sodium or potassium.



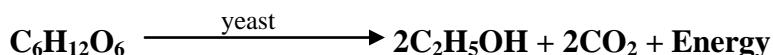
Method: Place some limestone or marble into a flat bottomed flask. Pour dil. HCl or HNO₃ through the thistle funnel.

Observation: There will be rapid effervescence and a colourless gas is produced. The gas is then collected over water or by downward delivery.

Note: Dil. H₂SO₄ is not used in this reaction because it coats the marble with a layer of insoluble calcium tetraoxosulphate(iv), which will terminate or stop the reaction after some seconds.

Other reactions that can result to the liberation of CO₂ are as follows:

- i. Fermentation of glucose



- ii. Thermal decomposition of marble



- iii. Combustion of hydrocarbons



- iv. Reduction of metallic oxides by carbon(ii)oxide



Properties of Carbon(iv)oxide

- a. CO₂ is a colourless, odourless and tasteless gas.

- b. It is denser than air

- c. It is slightly soluble in water. It dissolves in water to form trioxocarbonate (iv) acid

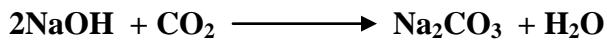
It is therefore an acid anhydride of trioxocarbonate (iv) acid.

- d. It is an acidic oxide. It turns wet blue litmus paper faint red or claret red.

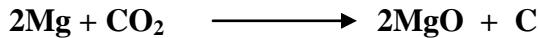
- e. It can either be solidified or liquefied. It solidifies at 78°C and the solid form is known as dry ice.

- f. It neither burns or supports combustion.

- g. Caustic soda solution absorbs CO₂ to form sodium trioxocarbonate (iv) and water.

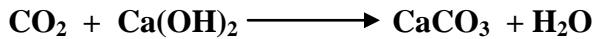


Burning magnesium continues to burn in a jar of CO₂ because the flame is so hot that it can decompose CO₂ into carbon and oxygen and the latter supports the burning.

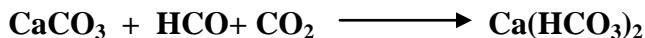


Test For CO₂

CO₂ turns lime water milky



When excess CO₂ is bubbled into lime water for a long time, the milky solution becomes clear again due to the formation of soluble calcium hydrogen trioxocarbonate(iv)



Uses of CO₂

1. It is used in fire extinguishers.
2. Green plants use it in the making of their food.
3. It is used in the manufacture of these compounds NaHCO₃, Na₂CO₃, PbCO₃, (NH₄)₂SO₄
4. In the baking of bread, it is used as a leavening agent.

14.10 Carbon Cycle

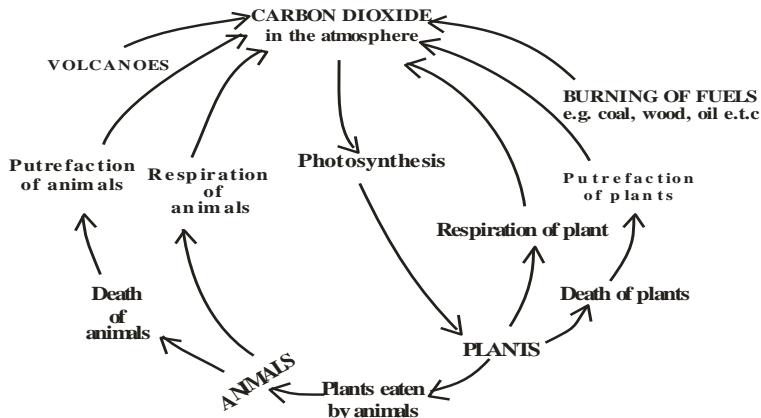
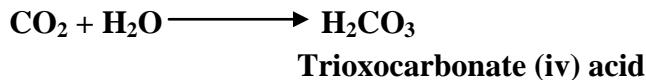


Fig 14.5

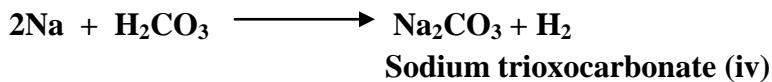
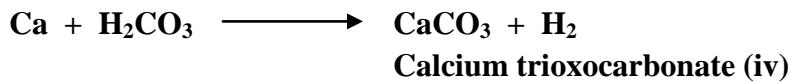
Carbon cycle is the process by which the atmospheric CO₂ is removed and re-introduced into the atmosphere. It is the process by which CO₂ circulates in nature. The volume of CO₂ in the air remains almost constant because the rate at which CO₂ is removed from air by the process of photosynthesis and by its dissolution in rain water is equal to the rate at which the gas is re-

introduced into the environment by plant, animal, burning of carbon compounds and decomposition of carbonates in the soil.

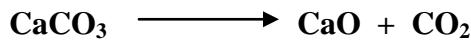
CO_2 dissolves in rain water to form a weak acid known as trioxocarbonate(iv)acid.



The weak acid on entering the soil, it reacts with the metals, to form the metallic trioxocarbonate (iv).



When the soil gets heated by the sun, the temperature of these trioxocarbonate (iv) and trioxocarbonate (iv) and hydrogen trioxocarbonate (iv) is raised and they decompose to release CO_2 into the atmosphere.



14.11 Carbon(II)Oxide (CO)

CO is prepared by dehydrating methanoic acid or ethanedioic acid with concentrated H_2SO_4 .

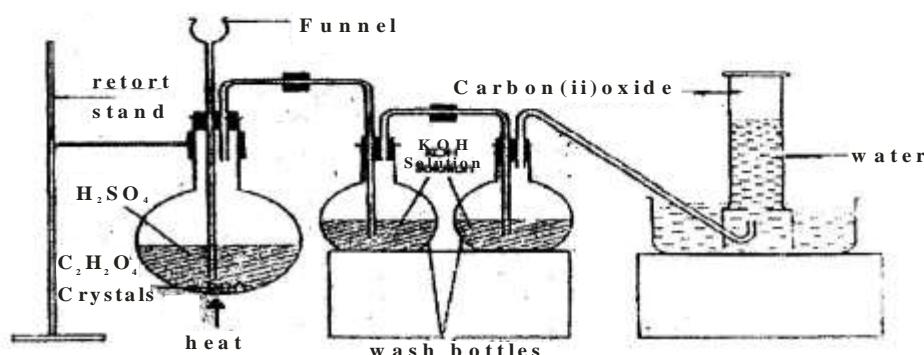
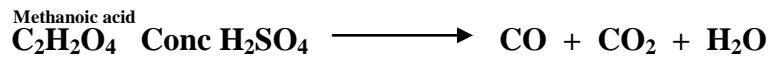
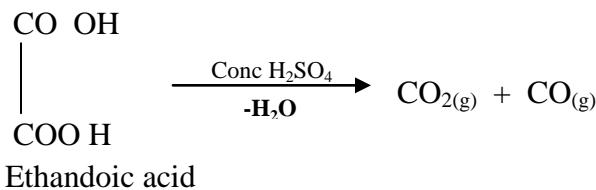


Fig 14.6 Laboratory preparation of CO.

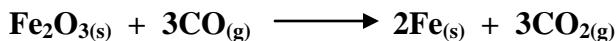
The apparatus used is shown above. Some crystals of Ethanedioic acid ($\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$) is placed in a strong round bottomed flask. Concentrated H_2SO_4 is poured down the thistle funnel. The mixture is warmed gently. As the crystals dissolve, effervescence occurs. A mixture of equal volume of CO and CO_2 is evolved. The gas mixture is passed through KOH or NaOH which dissolves CO_2 to form trioxocarbonate (iv) salt of potassium or sodium, CO is collected over water.



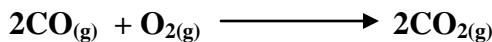
Conc. H_2SO_4 is not a catalyst in this reaction. It acts as a dehydrating agent.

Properties of CO

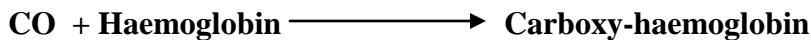
1. It is a colourless, odourless and tasteless gas
2. It is lighter than air
3. It is a neutral oxide
4. It is insoluble in water
5. It is soluble in ammonical copper(i) chloride
6. It is a strong reducing agent.



7. It is oxidized to CO_2 when it burns in air with a blue flame.



8. It combines readily with haemoglobin to form carboxy-haemoglobin (a stable compound), which prevents haemoglobin from doing its usual work of carrying oxygen. Haemoglobin reacts with oxygen to form oxy-haemoglobin (an unstable compound). The formation of carboxy-haemoglobin when CO is inhaled in a large dose, makes one to die of suffocation.

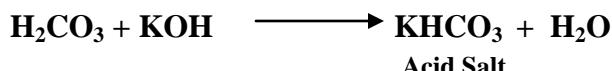
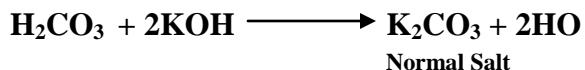


Uses of Carbon (ii) oxide, CO.

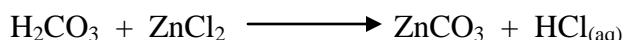
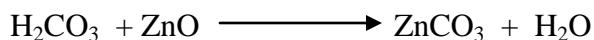
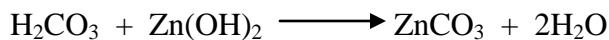
1. It is used as fuel
2. It is used a reducing agent

14.12 Trioxocarbonate (IV) Salt

H_2CO_3 acid is dibasic and forms two types of salts. The salts are trioxocarbonate (iv).



Trioxocarbonate (iv) salt is formed when H_2CO_3 reacts with free metal, metallic hydroxide, metallic oxide or other metallic salts



Properties of Trioxocarbonate (iv) Salts

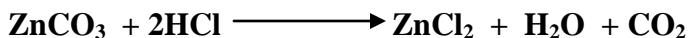
1. All trioxocarbonate (iv) salts are insoluble in water except Na_2CO_3 , K_2CO_3 and $(\text{NH}_4)_2\text{CO}_3$.
2. All common trioxocarbonate (iv) break down on heating to liberate CO_2 with the exception of Na_2CO_3 , K_2CO_3 and BaCO_3



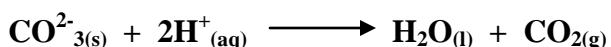
3. Na_2CO_3 , K_2CO_3 and $(\text{NH}_4)_2\text{CO}_3$ dissolved in water to form the corresponding hydroxide.



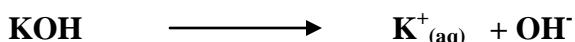
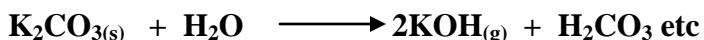
4. All trioxocarbonate (iv) react with any dilute acid to form CO_2 , water and salt.



This equation can be represented ionically as:



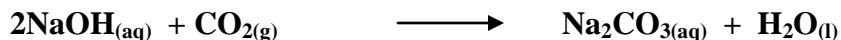
5. Sodium, ammonium and potassium trioxocarbonate (iv) Dissolves in water to form a weak acid and a strong base.



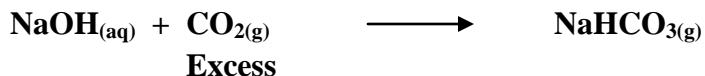
KOH is a strong base but H_2CO_3 is a weak acid. The resultant solution is alkaline to litmus because the number of OH^- is greater than that of the hydrogen ions.

14.13 Preparation of Baking Soda (NaHCO_3)

Baking powder is a common name for NaHCO_3 , but its IUPAC nomenclature is sodium hydrogen trioxocarbonate (iv). NaHCO_3 can easily be prepared in the laboratory by bubbling excess CO_2 into conc. NaOH solution. NaHCO_3 is then precipitated as a white powder.



Or



NaHCO_3 is slightly soluble in water; it is filtered, washed with a little cold distilled water and dried between filter papers.

Uses of NaHCO_3

- It is used in manufacturing effervescent tablets used to relieve indigestion (or constipation)
- It is used in the production of baking powder, which comprises NaHCO_3 and tartaric acid. The two ingredients react when they liberate CO_2 , which causes the dough to rise during the baking process.

2. It is used in the production of baking powder, which comprises NaHCO_3 and tartaric acid. The two ingredients react when they come in contact with water to liberate CO_2 , which causes the dough to rise during the baking process.

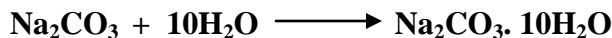
14.14 Laboratory Preparation Of Soda Ash (Na_2CO_3)

Sodium trioxocarbonate (iv), Na_2CO_3 is prepared in the laboratory by thermal decomposition of NaHCO_3 .



14.15 Laboratory Preparation Of Washing Soda $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$

Sodium trioxocarbonate (iv) decahydrate commonly known as the washing soda is prepared in the laboratory by dissolving Na_2CO_3 in water, concentrating the solution and allowing crystallization to take place.



14.16 Industrial Preparation of Soda Ash (NaCO_3)

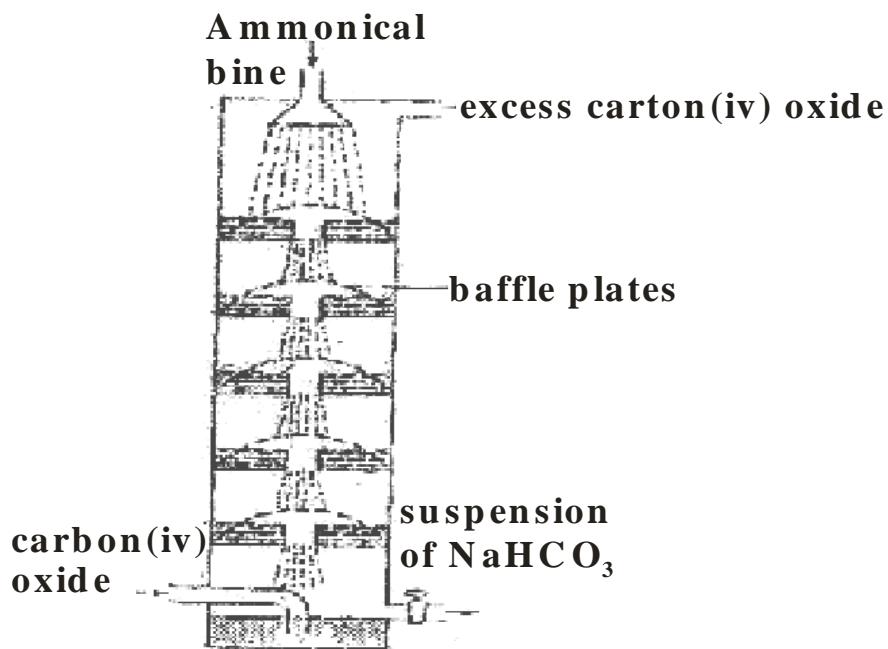
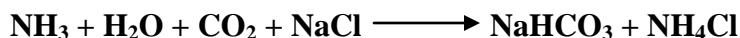
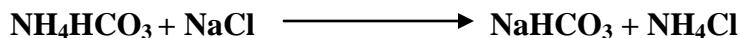
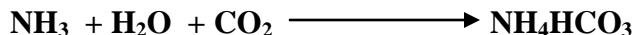


Fig 14.7 solvay process

Na_2CO_3 is manufactured in a large quantity by the solvay process. First of all, a mixture known as the ammoniacal brine is made by saturated a concentrated solution of sodium chloride (brine) with ammonia gas or by bubbling in enough quantity of NH_3 gas into brine. The ammoniacal

brine (NaCl solution + NH₃) is made to trickle down a solvey tower. CO₂ is at the same time forced into the tower and made to move upwards.

The perforated dome-shaped baffled-plates, slow down the flow of the ammoniacal brine and gas (CO₂). They also create greater surface areas for reaction. These reactions which occur in the solvey process is exothermic and so the tower has to be cooled.



NaHCO₃ is filtered, rinsed with distilled water and heated to yield an anhydrous salt (Na₂CO₃) known as soda ash. Na₂CO₃ is dissolved in water and crystallized to form the washing soda (Na₂CO₃. 10H₂O)

14.17 Economic Advantage And Efficiency of the Solvey Process

The principal raw materials are cheap and are easily obtained as follows:

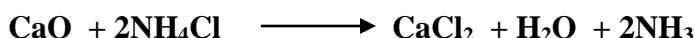
- i. Brine is obtained from sea-water or salt deposits.
- ii. CO₂ is obtained from the thermal decomposition of limestone.



About 50% of CO₂ required in the process is recovered while heating NaHCO₃.



- iii. The CaO obtained from number (ii) above reacts with NH₄Cl (left after the precipitation of NaHCO₃) to liberate NH₃ gas.

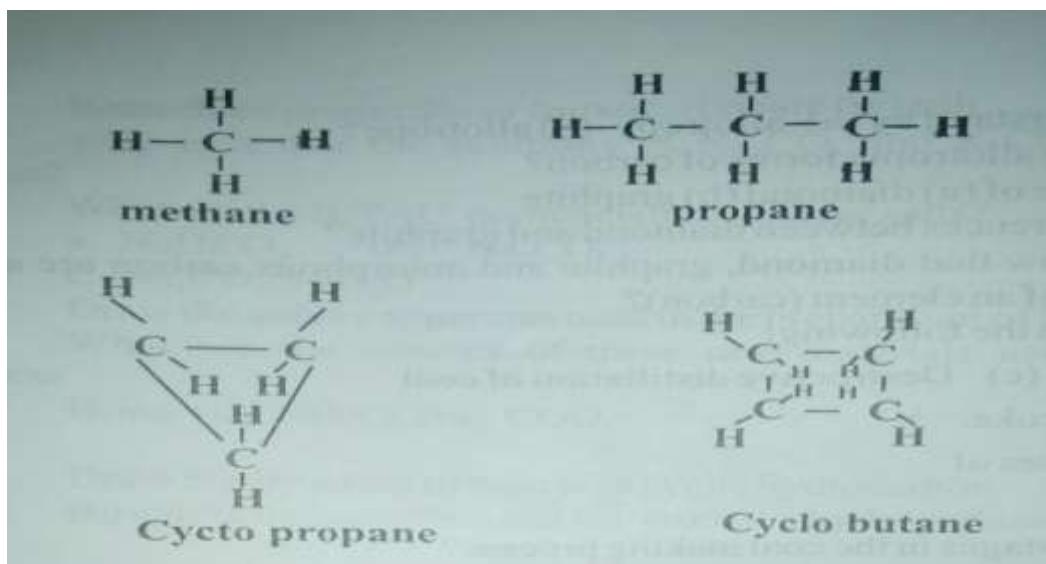


Uses of Na₂CO₃

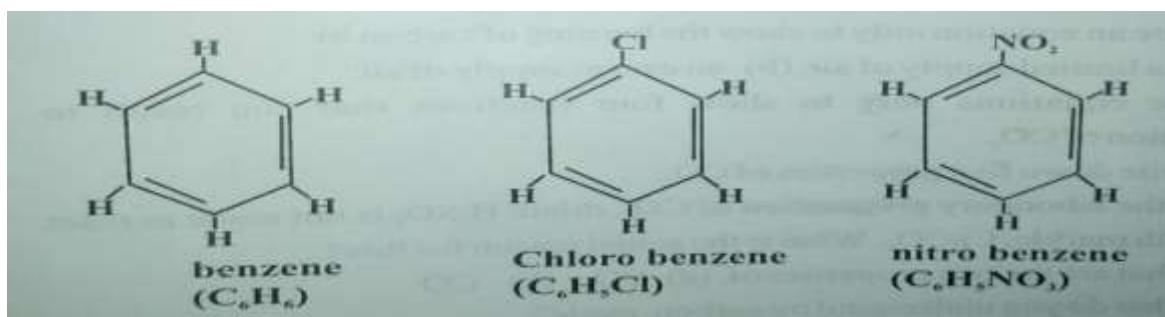
1. It is a useful reagent in the qualitative and volumetric analysis.
2. It is used in the manufacture of glass.
3. It is used in the softening of hard water.
4. It is use in the making of caustic soda (NaOH).

14.18 Hydrocarbon and its Main Classes

The name ‘hydrocarbon’ is derived from the two elements - hydrogen and carbon. Hydrocarbons are organic compounds that contains only carbon and hydrogen. Some examples of hydrocarbons are; methane, CH_4 ; Propane, C_3H_8 ; butane, C_4H_10 ; benzene, C_6H_6 . Hydrocarbons are classified into two groups. They are; the aliphatic and the aromatic hydrocarbons. Aliphatic hydrocarbons are composed of carbon chains. Aliphatic compounds with either straight or branched chain hydrocarbons are also called the acyclic compounds. Aliphatic compounds where the end carbons atoms join together to form rings are known as cyclic compound. Aromatic hydrocarbons are all cyclic compounds that contain benzene ring. They have sweet odour and are used in the manufacture of perfume. Examples of aliphatic hydrocarbon are;



Examples of aromatic hydrocarbons are;



14.19 Crude Oil and Natural Gas: See the fractional distillation of petroleum and major products in Part III, Chapter 2 of this book.

Exercises

1. What do you understand by (a) allotropy (b) allotrope?
2. What are the three allotropic forms of carbon?
3. Draw the structure of (a) diamond (b) graphite
4. What are the differences between diamond and graphite?
5. How can you show that diamond, graphite and amorphous carbon are all allotropic forms of an element (carbon)?
6. Write short note on the following:
 - a. Coal (b) Coke (c) Destructive distillation of coal (d) Gasification of coke.
7. Write down two uses of
 - a. Coal (b) Coke
8. What are the four stages in the coal making process?
9. Which of the two gaseous coke (producer gas and water gas) has more calorific value? Why?
10. Write an equation only to show the burning of carbon in
 - (a) a limited supply of air,
 - (b) an excess supply of air.
11. Use equations only to show four reactions that can result to the generation of CO_2 .
12. Write down five properties of CO_2 .
13. In the laboratory preparation of CO_2 , dilute H_2SO_4 is not made to react with marble, CaCO_3 . What is the actual reason for this?
14. What are the two properties of (a) CO_2 (b) CO
- 15a. What do you understand by carbon cycle?
 - b. Write two processes through which
 - i. CO_2 is removed from the atmosphere.
 - ii. CO_2 is re-introduced into the atmosphere.

16. Briefly discuss the effect of inhaling CO in a large dose.
17. Trioxocarbonate (iv) acid is dibasic and can form two type of salts. Using the chemical equation only show:
 - a. The two type of salts
 - b. Write the IUPAC nomenclature of the two salts
18. Write three properties of trioxocarbonate (iv) salt
19. Why is it that the solutions of Na_2CO_3 and K_2CO_3 are alkaline to litmus?
20. What are the IUPAC nomenclature of these salts?
 - a. NaHCO_3
 - (b) Na_2CO_3
 - c. $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
- 21a. Draw the solvey apparatus used in the preparation of Na_2CO_3 .
- b. What are the sources of these raw materials used in the solvey process
 - i. Brine
 - (ii) NH_4Cl
 - (iii) CO_2
22. Draw the structure of two; (a) a cyclic hydrocarbon
(b) cyclic hydrocarbon and (c) aromatic hydrocarbon

PART 2

CHAPTER 1

THE PERIODIC TABLE

KEY

- Atomic number
- Symbol of element
- Name of element
- Relative atomic mass (approximate)

GROUP	1	2	3	4	5	6	7	0
1	H Hydrogen							He Helium
2	Li Lithium	Be Beryllium						Ne Neon
3	Na Sodium	Mg Magnesium	Transition Elements					Al Aluminum
4	K Potassium	Ca Calcium	Sc Scandium	Ti Titanium	V Vanadium	Cr Chromium	Mn Manganese	Fe Iron
5	Rb Rubidium	Sr Strontium	Y Yttrium	Zr Zirconium	Nb Niobium	Mo Molybdenum	Tc Technetium	Ru Ruthenium
6	Cs Cesium	Ba Barium	La Lanthanum	Hf Hafnium	Ta Tantalum	W Tungsten	Re Rhodium	Os Osmium
7	Fr Francium	Ra Radium	Ac Actinium					

Fig 1.1 The Periodic Table

Read Periodicity and Periodic Table in Chapter 6 of Part I.

A period is a horizontal row of elements with the same number of shells. The elements in period one have only 1 shell and elements in period three have 3 shells.

Groups 1 and 2 are the S block elements. Their outermost or their valence electrons are located on the S - orbital.

Groups 3 to 0 are the P block elements. Their outermost or their valence electrons are on the P - orbitals. Transition metals are the d block elements.

Hydrogen has a unique position in the periodic table.

1.1 Trends in Properties Down the Groups and Across the Periods:

1. **The atomic volume:** is the volume occupied by 1 mole of an element. It can also be defined as the ratio of the relative atomic mass of an element to its density.

$$\begin{aligned}
 d &= m \\
 &\quad v \\
 dv &= m \\
 v &= m \quad \text{Where } d = \text{density} \\
 &\quad d \quad m = \text{relative atomic mass} \\
 &\quad v = \text{volume}
 \end{aligned}$$

Atomic volume decreases across the periodic table from left to right (i.e from group I to group 0). The reason for this is because the increasing atomic number (number of protons) has more and more attraction on the extra electrons which are being drawn closer and closer to the nucleus.

Down the group in the periodic table, the atomic volume increases with the increasing atomic numbers. This is because the net attractive forces are further reduced by the screening or shielding effect, of innermost shell electrons. The electron in the outermost shell therefore becomes less attracted to the nucleus. The result is that the electron is now located further away from the nucleus, thereby causing increasing atomic volume down the groups.

2. **Atomic and Ionic radii:** is the distance from the nucleus of an atom to the outermost shell. Atomic or ionic radii decreases with increasing atomic number. It is evident that the nuclear attraction increases more rapidly across a period and as a result the atomic radius decreases with increasing atomic numbers. Thus, the elements become progressively more electronegative from left to right.

The elements of any one group become decreasingly electronegative with increasing atomic number (ie down the group).

3. **Ionization Energy or Ionization Potential:** is the energy required to remove a valence electron from an atom of the element to form an ion. Ionization energy increases across the period from left to right. This is because as the atomic number increases, the atomic radius decreases. This causes the positive nucleus to exert stronger attraction to the outermost electrons. Hence the energy that will be required to remove such electron will be higher. Ionization energy decreases down a group because the outermost electrons are screened or shielded from the nucleus by the repelling effect of the inner electrons. There are 1st, 2nd, 3rd etc ionization energies. The first ionization energy of an element is the energy required to remove an electron from the outermost shell of a gaseous atom to produce a gaseous ion. Ionization energy is dependent on three factors. They are:

- i. The distance between the outermost electron and the nucleus.
 - ii. The screening or the shielding effect of the inner electrons.
 - iii. The size or the strength of the positive nuclear charge.
4. **Electron affinity:** is the energy released when an electron is added to a gaseous atom to form a gaseous ion. It is the opposite of ionization energy, because ionization energy deals with the loss of electrons, while electron affinity deals with the gain of electrons. Electron affinity decreases (ie increases in negative value) from left to right across a period and increases (ie decreases in negative value) from top to bottom within a group.
5. **Electronegativity:** is the tendency or ease with which an atom bonded in a molecule attracts an electron to itself. It can also be defined as a measure of the readiness of an element to accept electrons and change into negative ions.
- Electronegativity (EN) increases along the periods from group 1 to group 7 and decreases down the groups.
6. **Electropositivity:** is a measure of its readiness to give up electrons and form a positive ion. Electropositivity of metallic elements decreases down the electrochemical series (ECS). Electropositivity is the chemical opposite of electronegativity.

Exercises:

1. What do you understand these terms to mean?
 - (a) Periodicity of elements
 - (b) Periodic table.
2. Why is it that groups 3 to O are called the P block elements?
3. What are the three factors that affect the ionization energy of an atom?
4. Define these terms: a. Atomic volume b. Atomic radius
c. Ionization energy d. Electron affinity
e. Electronegativity f. Electropositivity

CHAPTER 2

CHEMICAL REACTIONS/ REACTANTS AND PRODUCTS

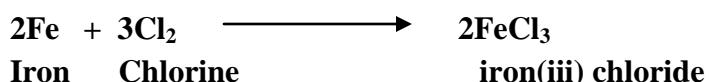
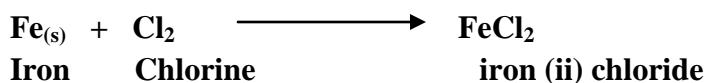
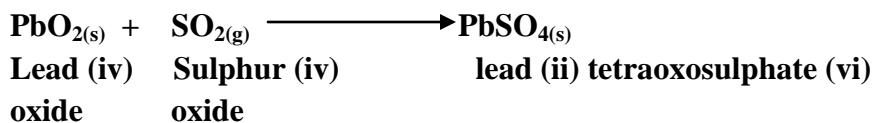
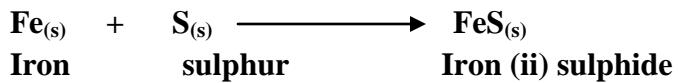
Substances undergo chemical reactions to form new chemical substances. The substances that undergo chemical reaction or chemical changes are called the reactants while the new chemical substances formed are called the products. Some chemical reactions that can be identified with ease in the chemistry laboratory are:

- | | |
|------------------------------|------------------------------|
| I Chemical combination | v. Catalytic reaction |
| ii. Decomposition | vi. Reversible reaction |
| iii. Double decomposition | vii. Thermal dissociation |
| iv. Displacement reaction | viii. Oxidation reaction |
| | ix. Reduction reaction. |

2.1 Chemical Combination

A chemical combination reaction is said to have taken place when two or more substances combine to form a single compound or a single substance.

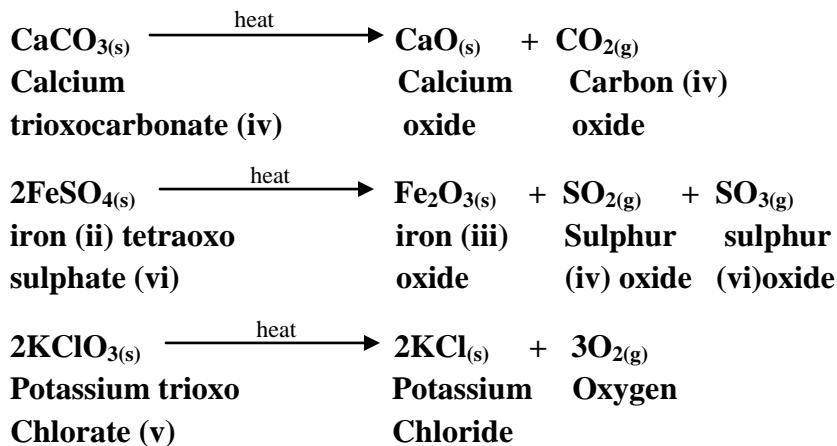
Example



2.2 Decomposition Reaction

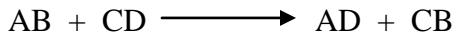
This is a chemical reaction which occurs when a compound splits up into simpler substances. If the decomposition is brought about by the application of heat, the reaction is known as the thermal decomposition reaction.

Example:

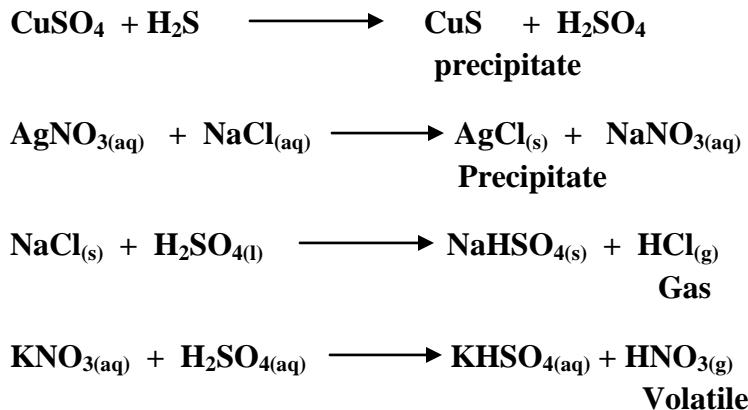


2.3 Double Decomposition Reaction

This is a chemical reaction in which the two reactants are both decomposed and the two new substances (products) are formed by an exchange of radicals.



Example:

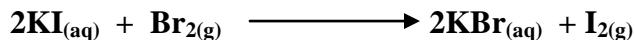
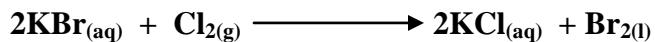
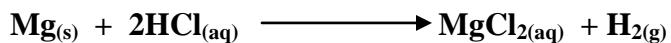


NOTE: Frequently, the two reactants are soluble in water while one of the products is soluble and the other is either volatile or insoluble.

2.4 Displacement

Displacement reaction occurs when an element (or a radical) takes the place of another in a compound. Ability of an element (or a radical) to displace another is dependent on their relative positions in the electrochemical series.

Example:



2.5 Catalytic Reaction

A catalytic reaction is any type of chemical reaction which employs a catalyst either to quicken or retard the rate of a chemical reaction. We therefore have catalytic decomposition, catalytic combination reaction, catalytic redox reaction etc.

A catalyst is a substance which alters the rate of a chemical reaction but remains chemically and quantitatively unchanged at the end of the reaction. A catalyst is specific in action. Examples of catalyst for some common catalytic reactions are:

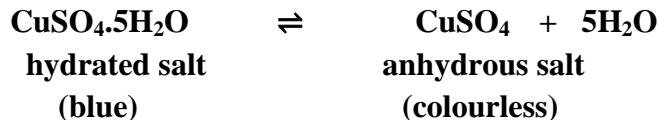
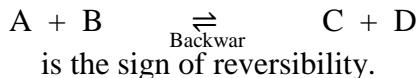
Reaction	Catalyst
a. Heating of potassium trioxochlorate (v) $2\text{KClO}_3 \longrightarrow 2\text{KCl} + 3\text{O}_2$	Manganese (iv) oxide (MnO_2)
b. Synthesis of Ammonia $\text{N}_{2(g)} + 3\text{H}_{2(g)} \longrightarrow 2\text{NH}_{3(g)}$	Finely divided iron
c. Synthesis of sulphur (iv) oxide $2\text{SO}_{2(g)} + \text{O}_{2(g)} \longrightarrow 2\text{SO}_3$	Vanadium (v) oxide
d. Decomposition of hydrogen peroxide $2\text{H}_2\text{O}_2 \longrightarrow 2\text{H}_2\text{O} + \text{O}_2$	Manganese (iv) oxide

We have positive and negative catalysts. Positive catalyst increases the rate of a chemical reaction but negative catalyst lowers the rate of a chemical reaction.

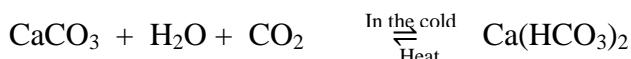
2.6 Reversible Reaction

This is the reaction that can be made to go in both forward and backward directions either by changing the condition of the reaction or by keeping them constant.

Forward

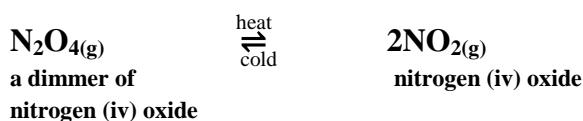
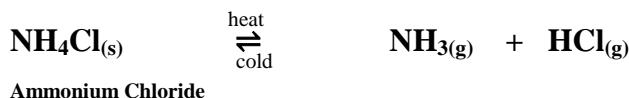


It is important to note that the conditions for both the forward and the reverse reactions in the above examples are the same. Other reversible reactions occur under different conditions. eg



2.7 Thermal Dissociation

This is a reversible reaction brought about by the application of heat.

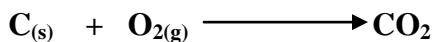


This reaction differs from thermal decomposition in that it is not reversible.

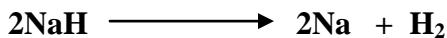
2.8 Oxidation

Oxidation reaction occurs when.

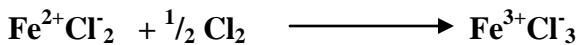
1. Oxygen is added to a substance eg



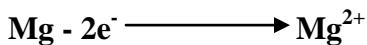
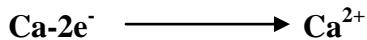
2. Hydrogen is removed from a substance eg



3. There is an increase in oxidation number of a substance eg.



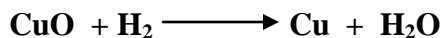
4. An element or ion loses electron or electrons eg



2.9 Reduction:

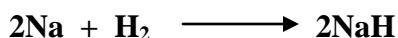
Reduction reaction occurs when

- Oxygen is removed from a substance eg



Copper oxide is reduced to copper.

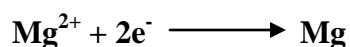
- Hydrogen is added to a substance eg.



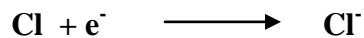
- There is a decrease in oxidation number of a substance eg



- An element or ion gains electron or electrons eg



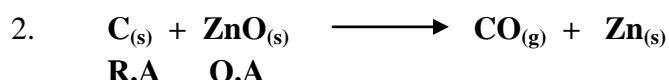
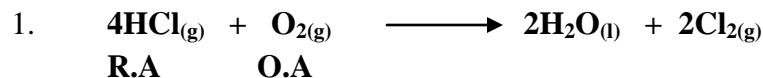
Magnesium ion is reduced to magnesium atom.



Chlorine atom is reduced to chlorine ion.

Reduction is the opposite of oxidation. In any reaction, where there is reduction, there is also corresponding oxidation and the reaction is called redox reaction. Redox reaction is therefore an abbreviated form of oxidation-reduction reaction.

Examples of redox reaction are:



An oxidizing agent (o. A) is

- An oxygen donor
- A remover of hydrogen
- An acceptor of electron

A reducing agent (r. A) is

- i. An oxygen acceptor
- ii. A hydrogen donor
- iii. An electron donor

2.10 Rate of Chemical Reaction

The field of study concerned with the rate of chemical reaction is known as chemical kinetics. The rate of a chemical reaction is the number of moles of reactants converted to products per unit time or the number of moles of product formed per unit time.

Rate of reaction = Change in concentration of reactant or product
The time taken for the change.

2.11 Factors that Affect the Rate of Chemical Reaction.

The following factors affect the rate of chemical reaction. They are:-

- 1. The concentration of the reactants
- 2. Temperature
- 3. Surface area of the reactants
- 4. Catalyst
- 5. Nature of the reactants
- 6. Pressure
- 7. Light.

Concentration:- If the concentration of the reactants decreases, the rate of the reaction decreases and if their concentration increases the rate of reaction also increases. It can be explained that with the decrease in the concentration of the reactants, the reacting molecules or particles become less crowded. The intermolecular distances of the particles become longer. This decreases the frequency of collision and hence the rate of reaction decreases. The reverse occurs if the concentration of the reactants is increased. Increase in pressure is equivalent to increases in concentration of gases.

Motivational experiment to show the effect of concentration on the rate of reaction:

Label 5 large test tubes numbers 1-5. Place them orderly in a test tube rack. Measure into test tubes (nos 2-4) these volumes of distilled water: 1.5 cm^3 , 3 cm^3 , 4.5 cm^3 , 6 cm^3 respectively. Do not pour water into test tube 1. Then measure into the 5 test tubes 1.5 cm^3 of conc HCl. Now 5

solutions of different concentrations have been made. Test tube - 1 contains the solution with the highest concentration while test tube 5 contains the one with the least concentration.

Measure 5 equal length (2cm each) of magnesium ribbon. Drop each Mg ribbon into the test tubes one after the other. Use a stop watch to record the length of time effervescence occur in each test tube (solution).

What conclusion can you draw from your observation?

Temperature: Increase in temperature of the reactants quickens the rate of a chemical reaction while decrease in temperature decreases the rate of a chemical reaction. The reaction rate is affected positively in two ways when the temperature of the reactants is raised.

Energy is supplied to the reactants in form of heat. They acquire greater energy. The number of particles with energies equal to or greater than the activation energy increases.

2. The average velocity (speed) of the reactants increases because they acquired greater kinetic energy. The reactants collide more frequently thereby increasing the rate of a chemical reaction.

Motivational experiment to show the effect of temperature on the rate of reaction: Measure about 15 cm³ of 1M hydrochloric acid into 5 small beakers labelled numbers 1 to 5. Measure out 5 equal length (3cm each) of magnesium ribbon. Drop each magnesium ribbon into the beakers one after the other. The solutions should be placed in the environment (oven) with the following constant temperatures: 25°C, 35°C, 45°C, 55°C and 65°C respectively. Use a stop watch to record the time duration of the effervescence. What inference can you draw from your observation?

the beakers one after the other. The solutions should be placed in the environment (oven) with the following constant temperatures: 25°C, 35°C, 45°C, 55°C and 65°C respectively. Use a stop watch to record the time duration of the effervescence. What inference can you draw from your observation?

Surface area: Surface area of contact affects the rate of a chemical reaction in a heterogenous system (eg. reaction between a solid and a liquid). Reaction occurs faster if the solid is in the powdered form than if it is in lumps. Grinding the solid increases its surface area and exposes the reacting particles in it.

Motivational experiment to show the effect of surface area:

Get 3 lumps of CaCO₃ (about 5 grams by mass). Leave one in lump. Break one into smaller lumps. Grind one into powder. Measure into 3 small beakers about 10cm³ of dilute hydrochloric acid. Label the beakers numbers 1 to 3. Now drop a lump of CaCO₃ (5g) into beaker 1. Drop

smaller lumps (5 g) into beaker 2. Drop the powdered CaCO_3 (5g) into beaker 3. These should be done one after the other. Again record the time duration of effervescence with a stop watch. What conclusion can you make from your observation?

Catalyst: Catalyst is a substance that quickens or slows down the rate of a chemical reaction but remains chemically unchanged at the end of the reaction. A positive catalyst increases the reaction rate by lowering the activation energy of the reaction. It lowers the activation energy and makes it possible for a greater number of molecules or reactants to possess the required activation energy.

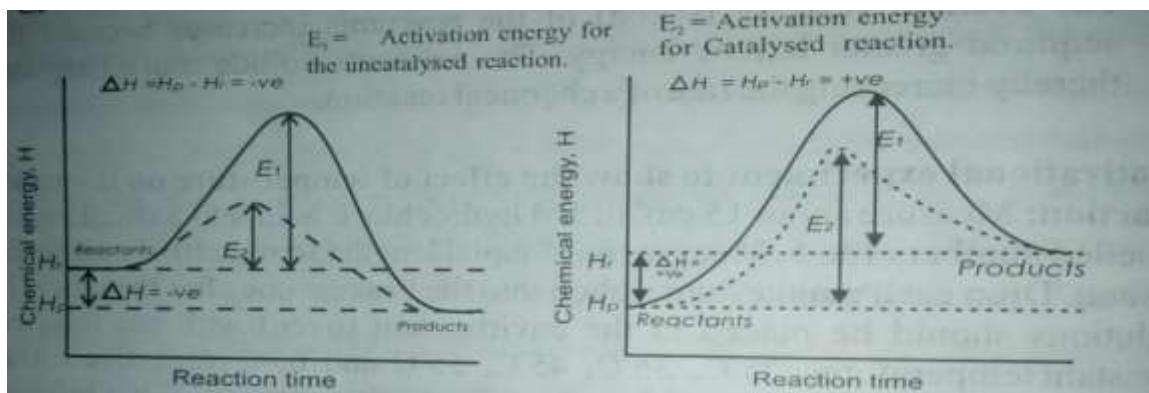


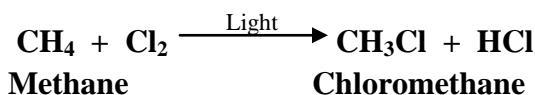
Fig. 2.11 Effect of Catalyst on activation energy.

The decomposition of KClO_3 and H_2O_2 to liberate oxygen is catalysed by manganese (iv) oxide.

Nature of reactants: The rate of reaction is affected by the chemical nature of the reactants. Al, Fe and Au will act or react differently when a piece of each is dropped in dilute HCl acid. If a piece of Al is added to dilute HCl, a rapid effervescence occurs. Hydrogen gas is liberated. A piece of iron placed in dilute hydrochloric acid shows a slow evolution of gas. The effervescence is not rapid. A piece of gold made to come in contact with dilute HCl acid shows no evidence of reaction. Al is the most reactive of the three metals. Au is the least reactive.

Light as a source of energy catalyses the rate of reaction of some reaction. Eg.

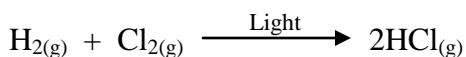
- Reaction between methane and chlorine



- Reaction between chlorine and hydrogen: This reaction is very negligible at ordinary pressure. It is slow in day light but very explosive in sunlight at room temperature. The light breaks up Cl_2 to chlorine radicals.explosive in sunlight at room temperature. The light breaks up Cl_2 to chlorine radicals.



The radicals react rapidly with hydrogen molecule to form hydrogen chloride gas.



2.12 Collision Theory: Energy and Reaction Rates:-

According to the collision theory, for a chemical reaction to occur, the reacting particles must collide. It is not all the collisions will result to a chemical reaction. Only the effective collision will cause a chemical reaction to occur. In other words, for a chemical reaction to occur, the collision must be effective. Effective collision is possible only if the reacting particles possess the necessary activation energy of reaction. The rate of a chemical reaction therefore depends on frequency of effective collision.

Chemical reaction involves breaking of bonds and formation of new bonds, and these require energy. It therefore means that the reacting particles must possess the minimal energy required for the breaking and formation of bonds. The energy barrier which the reacting particles must possess in order to react is known as the activation energy of the reaction.

During the reaction between two stable gaseous molecules (Hydrogen gas and Iodine gas), the covalent bonds between the two gases must first of all be broken before the formation of the new bond of the product.

Bond-breaking: $\text{H} - \text{H} \longrightarrow \text{H} + \text{H}$ (free atoms)

Bond-breaking: $\text{I}-\text{I} \longrightarrow \text{I} + \text{I}$ (free atoms)

Formation of new bond: $2\text{H} + 2\text{I} \longrightarrow 2\text{H}-\text{I}$

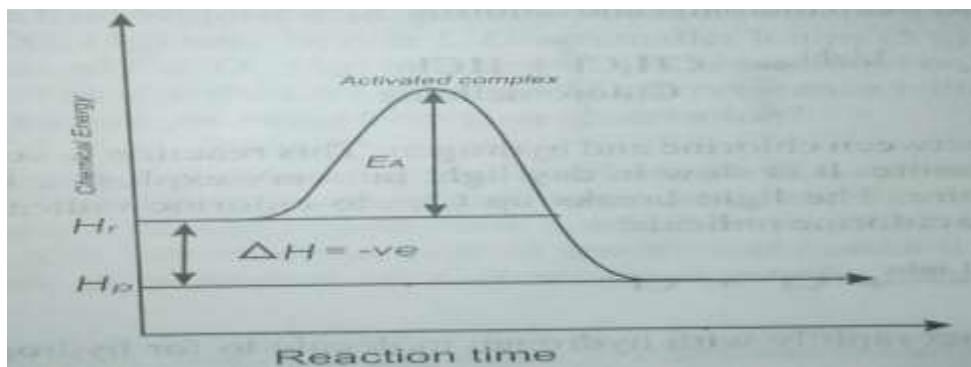


Fig. 2.21: Energy Profile diagram

H_r = Energy content of reactant

H_p = Energy content of product

E_A = Activation energy of the reaction

$$\Delta H = \text{Heat of reaction.}$$

If the energy of the colliding molecules is less than the activation energy, they merely rebound from each other and no reaction will therefore occur. Every reaction has its own activation energy. A reaction which has a high activation energy will take place if energy is supplied to it in form of heat, electrical or light energy. A reaction with low activation energy will occur spontaneously at low temperature.

2.13 Exothermic and Endothermic Reaction also Energy Profile

13.1 Motivational Experiment: Exothermic reaction:

1. Pour some water in a small beaker. Put into it some solid NaOH or KOH. Stir well with a stirrer. Pass it round for the students to feel the beaker with their palm.
2. Put 10cm³ of water into large test tube. Pour into it small quantity of concentrated H₂SO₄. Let the students feel the test tube with their palm.
3. Add NaOH solution to the dilute HCl in a test tube. Pass round for the students to feel the test tube with their palm.

After each experiment let the students report their observation.

Exothermic reaction:- is the reaction in which heat is liberated to the surrounding or environment.

The ΔH for exothermic reaction is negative. $\Delta H = -Ve$

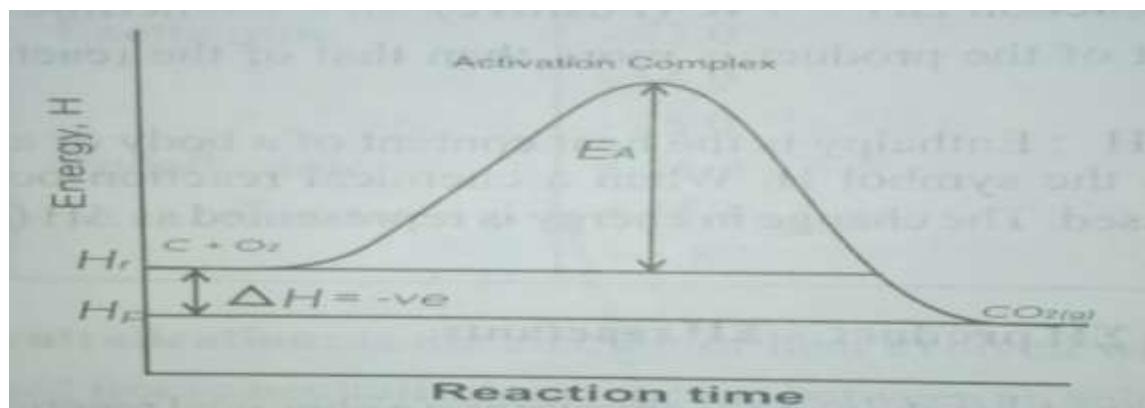


Fig. 2.3: Energy profile for exothermic reaction.

Endothermic reaction:

An endothermic reaction is the reaction in which heat is absorbed from the surrounding.



Motivational Experiment:

Dissolve some quantity of NH₄Cl crystals in some water poured into a test tube. Pass the solution round for the students to feel the test tube. Ask them individually to say what they observed.

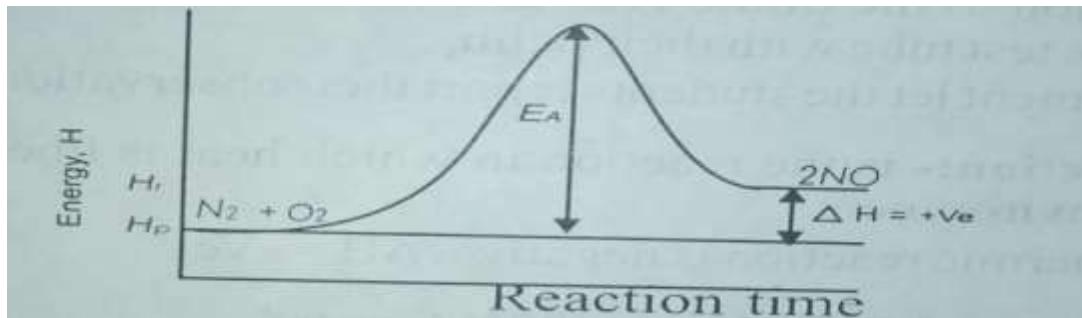


Fig. 2.4: Energy Profile For Endothermic Reaction

In endothermic reaction $\Delta H = +\text{Ve}$ (Positive). In endothermic reaction, the total heat content of the product is more than that of the reactants.

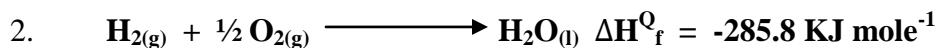
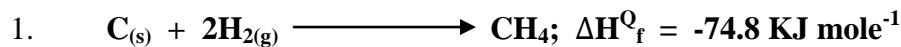
Enthalpy: H, ΔH : Enthalpy is the heat content of a body or a system. It is represented with the symbol H. When a chemical reaction occurs, heat is absorbed or released. The change in energy is represented as ΔH (Delta H). Δ means change.

$$\Delta H = \Sigma H_{\text{product}} - \Sigma H_{\text{reactants}}$$

Therefore, ΔH is the heat change accompanying a chemical reaction.

Standard enthalpy of formation (ΔH_f^{Q}):

Standard enthalpy of formation is the heat change accompanying the formation of one mole of a compound (a chemical substance) from its elements in their standard states. The standard state of an element is that in which it is at 1 atmosphere and 25°C. At this state all elements have enthalpies of zero.



Standard enthalpy of combustion:

Enthalpy of combustion is the heat change accompanying the complete combustion of one mole of a substance in oxygen.



Combustion is an exothermic reaction. The instrument used in measuring heat change in the laboratory is bomb calorimeter.

Enthalpy of solution: Heat of solution is the heat evolved or absorbed when one mole of a substance is dissolved in so much water that further dilution will result in no detectable heat change.

Solute	Type of heat Change	?H (KJ mole -1)
NaOH	Exothermic	- 43.0
KOH	“	- 54.9
Na ₂ CO ₃	“	- 25.0
NH ₄ NO ₃	Endothermic	+ 26.3
NH ₄ Cl	“	+ 15.1
NaCl	“	+ 3.8

Heat of neutralization: is the amount of heat evolved when one mole of water is formed from a neutralization reaction between an acid and an alkali or a base. It can also be defined as the amount of heat evolved when one mole of hydrogen ion (H⁺) from an acid reacts with one mole of hydroxyl ion (OH⁻) from a base to form one mole of water.

Heats of neutralization of acids and alkalis

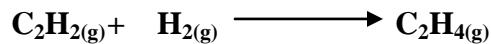
Acid	Alkali	Heat of neutralization KJ mole ⁻¹
HNO ₃	NaOH	57.1
HCl	KOH	57.3
HNO ₃	KOH	57.3
HCN	NaOH	12.0
HCl	NH ₄ OH	51.5
½ H ₂ SO ₄	NaOH	66.5
HNO ₃	½ Ca(OH) ₂	58.4

Calculation on Enthalpy Change:

1. Calculate the enthalpy change accompanying the following reaction given that;

$$\Delta H_f^{\circ} \text{ of } C_2H_2 = +227 \text{ KJ mole}^{-1}$$

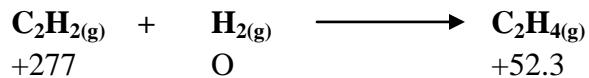
$$\Delta H_f^{\circ} \text{ of } C_2H_4 = +52.3 \text{ KJ mole}^{-1}$$



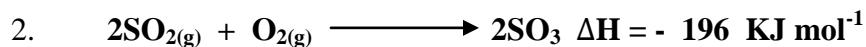
Solution

$$\Delta H \text{ of a reaction} = \sum \Delta H_f^{\circ} \text{ of product} - \sum \Delta H_f^{\circ} \text{ of reactants}$$

$$\text{Thus:- } \Delta H = \sum \Delta H_f^{\circ} \text{ product} - \sum \Delta H_f^{\circ} \text{ reactant}$$



$$\Delta H = + 52.3 - (+227 + O) = - 174 \text{ KJ mol}^{-1}$$



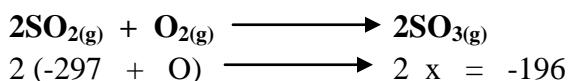
Given that ΔH_f^{Q} of $\text{SO}_2 = -297 \text{ KJ mol}^{-1}$

Calculate the heat of formation of sulphur (vi) oxide.

Solution

$$\Delta H_f^{\text{Q}} \text{ SO}_2 + \Delta H \text{ of combustion of } \text{SO}_2 = \Delta H \text{ of } \text{SO}_3$$

$$\text{Let the } \Delta H_f^{\text{Q}} \text{ SO}_3 = x \text{ KJ mol}^{-1}$$



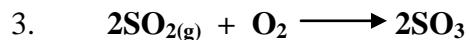
$$2x - (-594) = -196$$

$$2x = -196 - 594$$

$$2x = -790$$

$$x = \frac{-790}{2}$$

$$X = -395 \text{ KJ mole}^{-1}$$



$$\text{Given: } \Delta H_f^{\text{Q}} \text{ SO}_2 = -297 \text{ KJ mole}^{-1}$$

$$\Delta H_f^{\text{Q}} \text{ SO}_3 = -395 \text{ KJ mol}^{-1}$$

Solution



$$\Delta H = 2(-395) - [2(-297) + \text{O}]$$

$$= -790 - (-594)$$

$$= -790 + 594$$

$$= -196 \text{ KJ mole}^{-1}$$

4. Calculate the heat of formation of carbon(iv) oxide given that



$$\text{and } \Delta H_f^{\text{Q}} \text{ CO} = -108 \text{ KJ mole}^{-1}$$

Solution



$$2x - 2(-108) = -285$$

$$2x + 216 = -285$$

$$2x = -501$$

$$x = \frac{-501}{2}$$

$$x = -250.5 \text{ KJ mole}^{-1}$$

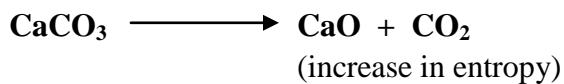
2.14 Spontaneity of Reaction; Entropy And Free Energy

Entropy is the degree of disorderliness of a system. It is represented by a symbol S. Entropy can be referred to as measure of the randomness, disorderliness, mixed-upness or scatteredness of the molecules or particles of a substance at a definite temperature and pressure. According to the second law of thermodynamics, any process which occurs in nature is accompanied by an increase in total entropy, eg when living creatures die, their bodies are converted by decay into simple gaseous products (such as CO₂ and NH₃) with increased entropy. Entropy is dependent on temperature and it is zero at absolute zero temperature.

The entropy of a substance can be increased by:

1. Rise in temperature.
2. Increase in volume eg. Melting of solid and vapourization of liquid are two situations which cause decrease in molecular complexity by increasing the entropy.

In chemical system, an increase in entropy is associated with decrease in molecular complexity or an increase in the number of gaseous molecules.



S is change in entropy accompanying a chemical reaction. If q is the heat absorbed by 1 mole of a substance at kelvin temperature, T, It follows that:

$$\frac{q}{T} = \Delta S$$

$$q = T \Delta S$$

When ΔS is positive = Increase in entropy.

When ΔS is negative = Decrease in entropy

Free energy, G: The total heat energy of a substance is called its enthalpy and is represented by H. Of this only part represented by G could be converted into useful work. G is called the free energy . Free energy, G is the heat energy of a substance that can be converted into useful work. If G_1 and G_2 are free energies before and after a chemical change, then change in free energy, G is given as:

$$G_1 - G_2 = G$$

For changes carried out at constant temperature and pressure.

The difference between ΔH and ΔG is the amount of energy involved in the entropy change. Thus If T is kelvin temperature

$$\Delta H - \Delta G = \Delta TS$$

$$\Delta G = \Delta H - \Delta T S$$

A chemical reaction is probable only if the value of G is negative. If $\Delta G = 0$, the reaction is in equilibrium.

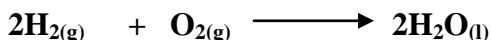
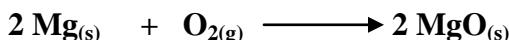
If $\Delta G = +Ve$, the reaction is not probable.

2.15 Negative value for G can arise in different ways:

1. ΔH is zero and ΔS is positive:

In some changes heat is neither absorbed nor evolved and the driving force of the reaction come from the increase in entropy Eg. A perfect gas expanding into a vaccum, mixing together of two gases or liquid at constant volume, transfer of solvent to a solution through a semi-permeable membrane.

2. ΔH is negative and ΔS is negative: Here - T ΔS is positive. For a change to be possible, the reaction must be exothermic and ΔH must be more negative than - T ΔS



In both cases there is a decrease in the number of gaseous molecules.

Monoclinic \longrightarrow Rhombic

The transition is exothermic and there is a decrease in entropy because the atoms are brought closer together. Rhombic sulphur has a higher density than Monoclinic sulphur.

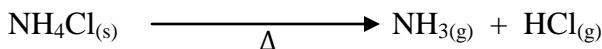
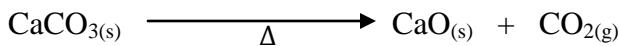
3. **ΔH is negative and ΔS is positive:**

When an exothermic reaction is accompanied by an increase in entropy, the two factors combine to give a high negative value for ΔG



4. **ΔH is positive and ΔS is positive:**

Here the change is endothermic and ΔG can only be negative if $T\Delta S$ is more positive than ΔH . That is, ΔH is less than $T\Delta S$. This may be possible at ordinary temperature but since $T\Delta S$ increases with rise in temperature, the condition may be fulfilled at higher temperature. Thus CaCO_3 and NH_4Cl split when strongly heated.



5. **ΔH is positive and ΔS is negative:** For any reaction which has ΔH positive and ΔS negative the ΔG must be positive and the reaction therefore is not possible.

System in equilibrium and factors affecting them, le Chatelier's principle

2.16 Chemical Equilibrium:

Chemical equilibrium is a condition in the course of a reversible chemical reaction in which no net change in the amounts of reactants and products occurs. A reversible chemical reaction is one in which the products, as soon as they are formed, react to produce the original reactants. At equilibrium, the two opposing reactions go at equal rates, or velocities, hence there is no net change in the amounts of substances involved.

The system in equilibrium has the following properties:-

1. Rate of forward reaction usually represented with r_1
2. Rate of backward reaction usually represented with r_2
3. Equilibrium constant usually represented with K .

For the reversible reaction $A \rightleftharpoons B + C$, the rate of forward reaction is expressed mathematically as

$R_1 = k_1 [A]$ where k_1 is the rate constant for the forward reaction and $[A]$ represents the concentration of A. The rate of backward reaction is expressed mathematically as $r_2 = k_2 [B]$

[C] where k_2 is the rate constant for the backward reaction while [B] and [C] represent the concentration of B and C respectively.

At equilibrium, $r_1 = r_2$, therefore.

$$k_1[A] = k_2[B][C]$$

$$\frac{k_1}{k_2} = \frac{[B][C]}{[A]} = K$$

Equilibrium constant is therefore the ratio of the rate of forward reaction to the rate of backward reaction. At equilibrium, G is zero.

reaction to the rate of backward reaction. At equilibrium, G is zero.

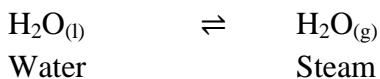
Le Chatelier's Principle states that: If a chemical system is in equilibrium and one of the factors involved in the equilibrium is altered, the equilibrium will shift so as to tend to annul the effect of the change.

This statement means that if one of the factors tending to keep a system in equilibrium is affected by either increasing or decreasing it, the other factors will change their values in order to restore the position of equilibrium.

2.17 Factors Affecting the Position of Equilibrium:

They are temperature, pressure and concentration.

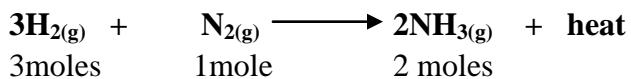
The effect of temperature change on the position of equilibrium is exemplified with the equilibrium between water and steam.



The forward reaction is an endothermic reaction. It is accompanied by an absorption of heat. If the temperature of the system is increased, this will cause the production of more steam (equilibrium position will move to the right) but if the temperature is decreased more water will be formed from the condensation of steam (equilibrium position will move to the left). Condensation of steam is exothermic. Raising the temperature causes endothermic reactions to occur but lowering the temperature leads to exothermic reactions.

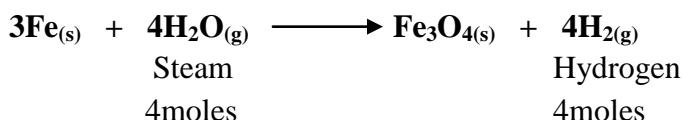


The effect of pressure change on the position of equilibrium is very significant only in reactions involving gases. Gas pressure is caused by the collision of the gas molecules on the walls of the container. In a container of constant volume, the gas pressure is directly proportional to the number of molecules present. The greater the number of molecules, the greater the pressure exerted. In the Haber process 4 moles of gaseous reactants react to yield 2 moles of NH₃.



Reduction in volume leads to drop in the pressure of the system. It can be said that if the pressure of the system is increased by introducing more of the gaseous reactants in the correct ratio, the equilibrium will shift so as to lower the pressure. This means that increase in pressure will favour the forward reaction in which there is a pressure reduction. For optimum yield of ammonia, high pressure of about 200 atmosphere is maintained.

In a reversible reaction where there is no change in the number of molecules (left to right), pressure change will have no effect on the equilibrium eg.



Effects of a change in pressure is summarised thus:

Types of reaction	Effect of increase in total pressure	Effect of decrease in total pressure
Increase in number of molecules left to right eg. $2O_{(g)} \rightleftharpoons 3O_{(g)}$ $N_2O_{(g)} \rightleftharpoons 2NO_{(g)}$	Position of equilibrium moves to the left eg. More O ₂ in equilibrium mixture. Eg. More NO ₂ in equilibrium.	Position moves to the right eg. More O ₂ in equilibrium mixture. Eg. More NO ₂ in equilibrium.
Decrease in the number of molecules left to right eg. $N_{(g)} + 3H_2 \rightleftharpoons 2NH_{(g)}$	Position of equilibrium moves to the right eg. More NH ₃ in equilibrium mixture.	Position of equilibrium moves to the left eg. More N ₂ and H ₂ in equilibrium mixture.
No change in number of molecules, left to Right. eg. $H_{(g)} + I_{(g)} \rightleftharpoons 2HI_{(g)}$	No effect. Position of equilibrium Maintained	No effect, position of equilibrium Maintained.

Exercises:

1. Briefly explain these types of chemical reaction and give an example of each.

- a. Decomposition reaction b. Displacement reaction
- c. Redox reaction d. Thermal dissociation reaction
- 2a. What do you understand these terms to mean?
- i. Reducing agent (R.A) ii. Oxidizing agent (O.A)
- b. Write a chemical equation showing both the reducing and oxidizing agents by writing the abbreviation R.A and O.A under them.
3. Define (a) ‘reactant’ (b) ‘product’
4. What do you understand by: “Rate of chemical reaction”?
5. What are the factors that affect, the rate of Chemical reaction?
6. Briefly explain a motivational experiment to show that area affects the rate of chemical reaction.
7. Use equations only to represent two reactions that are catalysed by light.
8. Write three ideas pointed out in the collision theory.
9. What is activation energy?
10. Draw an energy profile diagram and label the following: H_r ,
 H_p , E_A and ΔH .
11. What will happen:
- (i) If the energy of the reacting particles (reactants) is less than the activation energy?
(ii) If the energy of the reactants is greater than the activation energy?
12. Define these terms.
- (a) Exothermic reaction. (b) Endothermic reaction
13. Write a compound or a substance and briefly explain how it can be used to demonstrate a motivational experiment on:
- (a) Exothermic reaction (b) Endothermic reaction.
13. Define the following:
- (a) Enthalpy (b) Standard enthalpy of formation.
- (c) Enthalpy of solution (d) Heat of neutralization.

14. Calculate the heat of formation of carbon (iv) oxide given that



and $H_f^{\text{Q}} \text{ CO} = -108 \text{ KJ mole}^{-1}$

i. If the energy of the reacting particles (reactants) is less than the activation energy?

ii. If the energy of the reactants is greater than the activation energy?

15. What do you understand these terms to mean?

(a) Entropy of a substance.

(b) Enthalpy

(c) Free energy or Gibbs free energy.

16. What are the factors that can increase the entropy of a system?

17. Briefly explain these types of chemical reaction and give an example of each.

a. Decomposition reaction b. Displacement reaction

c. Redox reaction d. Thermal dissociation reaction

- 18a. What do you understand these terms to mean?

I. Reducing agent (R.A) ii. Oxidizing agent (O.A)

b. Write a chemical equation showing both the reducing and oxidizing agents by writing the abbreviation R.A and O.A under them.

19. Define (a) ‘reactant’ (b) ‘product’

20. What do you understand by: “Rate of chemical reaction”?

21. What are the factors that affect, the rate of chemical reaction?

22. Briefly explain a motivational experiment to show that surface area affects the rate of chemical reaction.

23. Use equations only to represent two reactions that are catalysed by light.

24. Write three ideas pointed out in the collision theory.

25. What is activation energy?

CHAPTER 3

MASS- VOLUME RELATIONSHIP

3.1 Two Methods Used in Chemical Calculations

1. Proportion Methods

Most chemical calculations are based on direct and inverse proportions.

1. Direct proportion:

Consider two variables: **A**, number of eggs, and **B**, cost of eggs. An increase in the number of egg leads to an increase in the cost of eggs; i.e. **A** is directly proportional to **B**.

Mathematically: $A \propto B$

2. Inverse proportion:

Consider two variables: length **I**, of a burning candle, at a particular time **t**, decreases the length **I**, of the burning candle; i.e.

t is inversely proportional to **I**.

Mathematically: $t \propto 1/I$

Calculations From the First Principle

When a mathematical problem is solved using the proportion method, it is said to be from the **first principle**, or from the **definition**. The method involves identification of a statement of fact.

The following example gives the steps to follow when solving problems involving direct proportion...

Worked Example

A doze of egg costs 300 naira.

- a. What will be the cost 45 eggs?
- b. How many eggs will cost 500 naira? (1 dozen = 12)

Answer

From the First principle: Method 1

- a. To find the cost of 45 eggs:

The statement of fact is: 12 eggs cost 300 naira (known).

The two variables involved are: number of eggs and the cost, in naira.

What is given in the question is the number of eggs; what is required (unknown) is the cost, in naira.

The four basic steps to follow are:

Step I: State or restate the statement of fact such that the given variable comes before the required variable; i.e number before cost of eggs:

12 eggs cost 300 naira.

Step II: Find the cost of 1 egg, by diving the cost 12 eggs by 12.

(Cost of one egg will be less.)

Step III: Find the cost of 45 eggs required, by multiplying the cost of 1 egg by 45.

(Cost of 45 eggs will be more than that of one egg)

i.e. 45 eggs cost $45 \times 300 / 12$ naira.

Step IV: Evaluate the expression and attach the correct unit, i.e.

$(45 \times 300) / 12 = 1,125$ naira.

Note: With adequate practice, it is possible to omit Step II, so that Step III is obtained from **Step I** by multiplying 45 by $300 / 12$.

- b. To find the number of eggs costing 500 naira.

The two variables involved are: number of eggs and cost. The statement of fact remains

12 eggs cost 300 naira.

Step I: The statement of fact is restated such that the given variable comes before the required variable, i.e 300 naira is the cost of 12 eggs.

Step III: Find the number of eggs, which costs one naira by dividing the number of egg by 300, i.e.

1 naira is the cost of $12 / 300$ eggs/

Step III: Find the number of eggs costing 500 naira by multiplying the cost of 1 egg by 500, i.e

500 naira is the cost of $500 \times 12/300$ eggs.

Step IV: Evaluate the expression and attach the correct unit. i.e.

$(500 \times 12)/300 = 20$ eggs.

Notes: In solving problems by proportion method:

Each statement is a complete sentence that contains at least two variables.

The unknown variable to be found comes last, i.e. at the end of the statement.

Method 2 (*Alternative Proportion Method*)

This proportion method involves six steps.

a. To find the cost of 45 eggs:

The statement of fact is: 12 eggs cost 300 naira.

Step I: The variables in the statement of fact are stated as being equivalent to one another, i.e.

$12 \text{ eggs} \equiv 300 \text{ naira}$.

Step II: Let x represents the required cost of 45 eggs. These two variables are stated as $45 \text{ eggs} \equiv x \text{ naira}$ (Not $x \text{ naira} \equiv 45 \text{ egg}$)

Step III. Like terms in steps **I** and **II** are divided:

$$\frac{12 \text{ eggs}}{45 \text{ eggs}} \quad \frac{300 \text{ naira}}{x \text{ naira}} \quad \text{or} \quad \frac{12}{45} \quad \frac{300}{x}$$

Step IV: Cross-multiply, such that the unknown x is one the left-hand-side of the expression, i.e.

$$12 \times x = 300 \times 45$$

Step V: Make x the subject of the formula;

$$x = (300 \times 45)/12$$

Step VI: Evaluate the final expression, and attach the correct unit of x , i.e $x = 1125$ naira.

b. To find the number of eggs costing 500 naira.

Let x represents the number of eggs.

Step I: $12 \text{ eggs} \equiv 300 \text{ naira}$ (given)

Step II: $x \text{ eggs} \equiv 500 \text{ naira}$ (required)

Step III: Cross-multiply:

Step IV: $300 \times x = 12 \times 500$

Step V: $x = (12 \times 500)/300$

Step VI: $x = 20$ eggs.

Use of Mathematical Formulae

A faster method of solving a mathematical problem in chemistry is to use an appropriate mathematical formula, which connects three or more variables. The required or unknown variable is made the subject of the formula, followed by correct substitution, then, evaluation.

A mathematical formula can be obtained via the proportion relationship of the various variables, e.g.

One mole of an element is equal to its relative atomic mass, expressed in grams. This is a statement of fact. The two variables involved are the amount, n in moles, and the mass, m is directly proportional to the amount, n .

Mathematically: $m \propto n$

The constant of proportionality is the molar mass, M of the element. The molar mass, M of a known chemical substance is constant, i.e.

$$\frac{m}{n} = M$$

$$\frac{\text{Mass, } m \text{ (g)}}{\text{Amount, } n \text{ (mol)}} = \text{Molar mass, } M \text{ (g mol}^{-1}\text{)} \quad \dots\dots 1$$

$$\text{Or Amount, } n \text{ (mol)} = \frac{\text{Mass, } m}{\text{Molar mass, } M} \quad \dots\dots 2$$

$$\text{Or Mass, } m = \text{Amount, } n \times \text{Molar mass, } M \quad \dots\dots 3$$

Note: In solving a mathematical problem in chemistry, the following steps should be taken:

1. List the given and required variables and the constants.
2. Provide a mathematical formula, which connects the variables
3. Do the correct substitution; then, make the unknown (required) variable the subject of the formula.
4. Evaluate to three significant figures, and insert the correct unit – when necessary.

Significant Figures

The term significant figure is used to describe the importance attached to a particular numeral or figure in a number.

In chemical calculations, the final answers are expected to be expressed to three significant figures.

In a decimal fraction:

The first significant figure is the first non-zero numeral when counted from the left-hand-side, e.g. in the number 0.0234, the first significant figure is 2, the second is 3, while the third is 4;

Zeros that are between non-zero significant figures are significant, e.g., 0.003005 is to 4 significant figures;

Zeros after a non-zero unit are significant, e.g. 0.0900 is to 3 significant figures;

When rounding off a number to a specific significant figure, the numerals 0,1,2,3 and 4 directly following the last digit of the required significant figure are ignored, while numerals 5,6,7,8 and 9 are rounded up to a whole number and added to the last digit, e.g. 0.07654 is 0.0765 when expressed to 3 significant figures, while 0.09876 is 0.0988 to 3 significant figures.

Example

One significant figures: 1;9; 0.1 ; 0.02 ;0.003 ; 0.004 ; 1 x 10³ ; 9 x 10⁻⁵ ; etc.

Two significant figures: 10 ; 99 ; 1.0; 9.9 ; 0.20 ; 0.030 ; 0.0044 ; 1.0 x 10³ ; 9.0 x 10⁵; etc.

Three significant figure: 101 ; 1.00 ; 2.34 ; 0.0860 ; 0.00504; 6.02 x 10²³ ; 9.65 x 10⁻⁵ ; etc.

Four significant figures: 1001 ; 100.0 ; 11.11; 0.1000 ; 0.01000 ; 0.09876; etc.

Relative atomic mass, relative molecular mass, standard temperature and pressure: Read chapter 6.5 and of this book.

3.2 Mass, Mole and Mass-Volume Relationship

Common Standard Units Of Measurement

The students are reminded of the standard units of measures, which are familiar to them, e.g.

M e a s u r e	U n i t
Mass	Grams, g
Length	C m
Area	C m ²
Volume	C m ³
Set of twelve	Dozen

Mole as a Standard Unit of Measurement:

A mole of a chemical substance is like a bag of rice. As a bag of rice is a collection of grains of rice, so also, a mole of a chemical substance is a collection of atoms, molecules, ions, electrons etc (ie elementary chemical units or particles).

If certain masses, in grams, of two elements are taken, such that they are equal to their respective atomic masses, the two elements will contain the same number of atoms (elementary units).

The relative atomic mass of carbon is **12.0**, whereas that of sulphur is 32.0. If as many carbon atoms are taken as to weigh 12.0g, and as many atoms of sulphur are taken as to weigh **32.0g**, both elements will contain the same number of atoms.

Inferences:

A mole has a specific number of **elementary units**.

Equal **amount**, in **moles**, of two or more chemical substances will contain the same number of **elementary units**.

Introducing the Avogadro's constant, L

The number of elementary units in one mole of a chemical substance is constant. It is called (**Avogadro constant, L**, and is numerically equal to **6.02×10^{23}** .

Inter-relationships: Concept of Molar Mass

There is the need for the students to be aware of the relationships between the atoms and the mole, on one hand, and the molecules and the mole on the other. This will enable the students to differentiate between an atom of a chemical substance, a molecule of a chemical substance, and a mole of a chemical substance.

Atoms and Moles

A mole is a collection of atoms of an element.

When the relative atomic mass of an element is expressed in grams, it is equal to the mass of one mole of its atoms. The mass of one mole of a chemical substance is called **molar mass, M**; i.e.

Molar mass = mass of one mole

Its S.I. unit is grams per mole (g/mol or g mol⁻¹)

The mass of one mole of carbon – 12 is 12.0 g; i.e.

The molar mass of carbon is 12.0 g/mol. Hence, the molar mass of sodium is 23.0 g/mol, while that of sulphur is 32.0 g/mol. Therefore:

1 mole of carbon atoms weighs 12.0 g and contains 6.02×10^{23} atoms.

1 mole of sodium atoms weighs 23.0 g and contains 6.02×10^{23} atoms.

1 mole of sulphur atoms weighs 32.0 g and contains 6.02×10^{23} atoms.

A mole is not a number, it is the S.I. unit of amount of a chemical substance. Hence the amount of carbon in 12.0 g of carbon atoms is 1.0 mole.

When comparison is being made between two or more chemical substances, any unit of mass can be used in expressing the molar mass; i.e.

12.0 kilograms of carbon will contain the same number of atoms as in 23.0 kilograms of sodium.

12.0 tonnes of carbon contain the same number of atoms as in 23.0 tonnes of sodium

Molecules and Moles

A molecule is composed of two or more atoms chemically bonded to form a unit. For instance:

H_2 represents one hydrogen molecule containing two hydrogen atoms bonded together chemically.

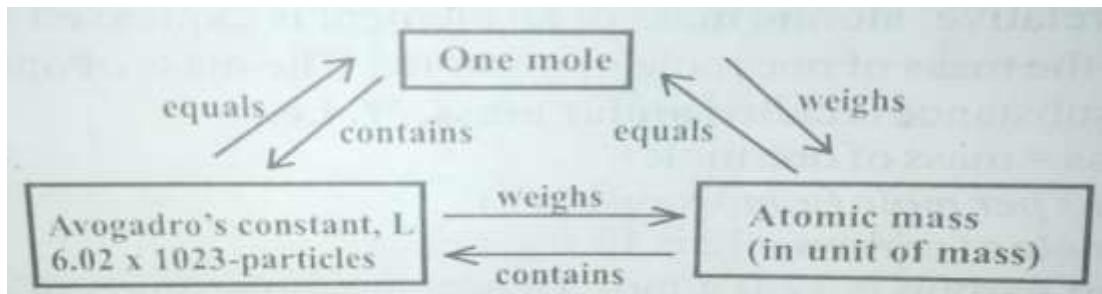
In terms of moles: H_2 represents one mole of hydrogen molecules, and contains two moles of hydrogen atoms. It weighs 2.0 g and contains 6.02×10^{23} molecules.

H_2O represents one mole of water molecules; it contains two moles of hydrogen atoms and one mole of oxygen atoms. It weighs 18.0 g and contains 6.02×10^{23} molecules.

HNO_3 represents one mole of trioxonitrate (V) acid; it contains one mole of hydrogen atoms, one mole of nitrogen atoms, and three moles of oxygen atoms. It weighs 63.0 g, and contains 6.02×10^{23} molecules.

H_2SO_4 represents one mole of tetraoxosulphate (VI) acid; it contains two moles of hydrogen atoms, one mole of sulphur atoms, and four moles of atoms of oxygen. It weighs 63.0 g, and contains 6.02×10^{23} molecules.

Figure 3.1 shows the relationship between the relative atomic mass of an element, in grams; its amount, in mole; and the Avogadro constant, L.



Step -wise approach to the Teaching of mole concept

Step 1: Introduction

In order to reduce students' mathematical problems, the teacher should use atoms, molecules and compounds whose masses are even numbers or are easily divisible by 5 or 10 in illustrating chemical calculations.

The following substance have masses that are even numbers, or multiples of 5 or 10: $\text{H}_2 = 2$; $\text{O}_2 = 32$; $\text{CH}_3 = 30$; $\text{MgO} = 40$; $\text{CaCO}_3 = 100$; $\text{NaHSO}_4 = 120$, $\text{NO} = 30$.

Step 2: Mass-Mole Relationship in elements and compounds.

Worked examples

The following worked examples will be used to illustrate the steps to be taken when solving problems involving mass-mole relationship from the first principle.

1. How many moles are in 15.0 g of carbon atoms? ($\text{C} = 12$)

Answer

The required variable is amount of carbon, $n = ? \text{ mol}$.

The given variable is mass of carbon, $m = 15.0 \text{ g}$.

The constant is molar mass of carbon, $M=12\text{g mol}^{-1}$

From the First Principle: Method 1

Note: In the statement of fact, the known or given variable (mass of carbon) is stated first, then, the required variable (amount).

12g of carbon atoms equal 1 mole

1 g of carbon atoms equals $1/12$ mol

$\therefore 15$ g of carbon atoms equal $15 \times 1/12$

$$= 1.25 \text{ mol (3 sig. fig)}$$

Method 2

Let x represents the amount, in moles.

1 mol of carbon = 12g (Statement of fact)

x mol of carbon = 15g (Required)

Cross-multiply such that the unknown is on the left:

$$12 \times x = 1 \times 15$$

$$X = 15/12 = 1.25 \text{ mol}$$

2. Calculate the mass of iron present in 0.025. mol of iron filings. (Fe = 56.0)

Answer

The require variable is mass, $m=?\text{g}$

The given variable is amount, $n=0.025 \text{ mol}$

The constant is the molar mass, $M=56 \text{ g/mol}$

From the First Principle: Method 1

1 mol of iron weighs 56g

$\therefore 0.025 \text{ mol weighs } 56 \times 0.025 = 1.40 \text{ g}$

Method 2

Let x represents the mass of iron filings.

1mol of iron = 56g

$$0.025 \text{ mol of iron} = x \text{ g}$$

Cross multiplying: $1 \times x = 56 \times 0.025$

$$X = 1.40 \text{ g}$$

3. If 0.123 mole of an element X weights 8.0g. What is the relative atomic mass of X?

Answer

The given variable are: amount $n=0.123$ mol and the mass, $m = 8.0\text{g}$.

What is required is the constant, molar mass, M in g/mol, i.e. mass of one mole.

From the First Principle

$$0.123 \text{ mol of X weighs } 8.0\text{g}$$

$$\therefore 1 \text{ mol of X weighs } 8.0 = 65.0\text{g}$$

$$0.123$$

Hence, relative atomic mass of X = 65.0 (no unit)

4. Given that 0.101 mole of an oxide X_2O weights 14.4g Calculate the :

- Molar mass of X_2O .
- relative atomic mass of X. ($O = 16.0$)

Answer

- a. To calculate the molar mass of X_2O

Given: amount = 0.101 mol; mass = 14.4g

From the First Principle

Let x represents the mass of 1 mole of X_2O .

$$0.101 \text{ mol} \equiv 14.4 \text{ g}$$

$$1 \text{ mol} \equiv x \text{ g}$$

$$\text{i.e } 0.101 \times x = 14.4 \times 1$$

$$x = 14.4 / 0.101 = 143 \text{ g}$$

Hence, the molar mass of $X_2O = 143 \text{ g/mol}$

- b To calculate the atomic mass of X. [O=16]

Since the molecular mass of $X_2O = 143$; then,

$$X_2O = 143$$

$$2X + 16 = 143$$

$$2X = 143 - 16 = 127$$

$$X = 127/2 = 63.5$$

Hence, the relative atomic mass of X is 63.5 (no unit)

5. Calculate the amount of sodium in 13.3g of pure anhydrous Na_2CO_3
[Na = 23.0; C = 12.0; O = 16.0]

To calculate the amount of Na in Na_2CO_3

This is mass-mole relationship.

Molar mass of $Na_2CO_3 = (23 \times 2) + (12 \times 1) + 16 + 3 = 46 + 12 + 48 = 106g$

mol^{-1} 106g (1mol) of Na_2CO_3 contain 2 mol ($2 \times 23g$) of Na.

i.e 106g of Na_2CO_3 contain 2 mol of Na.

\therefore 13.3g of Na_2CO_3 contain $(2 \times 23g)/106$ mol of Na

0.25 mol of Na

Revision Exercises

1. What is the molar mass of a substance, if 0.4 mole of the substance has a mass of 25.0 g?
A 6.3g B 40.0g C 62.5g D 2.5g (UTME)
2. What is the amount of gold present in 15.5g of a pure gold wedding ring? (Au = 197).
3. Calculate the mass of the substance present in:
 - a. 0.105 mol of $Na_2CO_3 \cdot 10H_2O$
 - b. 1.20 mol of $KMnO_4$ [H = 1; C = 12; O = 16; Na = 23; K = 39; Mn = 55]

Step 3: Calculations from Equations

Mass –Mole Relationships

A balanced chemical equation describes quantitatively, the amount, in moles, of the reactants used-up and the products obtained from a reaction. Such an equation is said to be stoichiometric because, the mole ratio of the reactants and/or products can be obtained from it.

Stoichiometric equations can be used to estimate the quantities of reactants required for, and/or the yield of products expected from a chemical reaction.

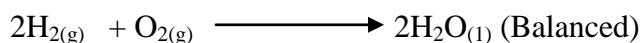
The steps to follow are:

1. Write a balanced equation for the reaction.
2. Write the amounts, in moles, under each of the relevant reactants and products
3. Convert the amount (in moles) to mass (in grams, Kilograms or tones) as required.
4. Find the quantity required by direct proportion. The statement of fact with at least two variables is obtained from the balanced equation. Remember: the variable to be found comes at the end of the statement.
5. Express your Final answer to three significant figures.
6. Finally, insert the appropriate unit.

Examples

a. Formation of Water

Hydrogen burns in air or oxygen to form water:



2 Molecules 1 molecule 2 molecules

2 moles 1 mole 2 moles

2(1x2)g (16x2)g 2(2+16) g

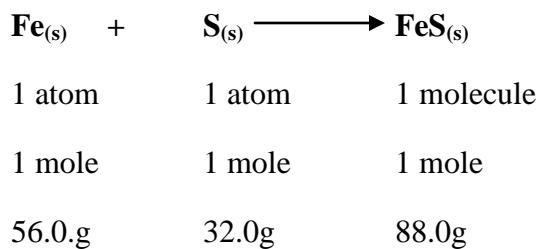
4.0g 32.0g 36.0g

The following information can be obtained from the balanced equation of the reaction:

- ? In elementary terms: Two molecules of hydrogen gas require One molecule of oxygen gas to produce Two moles of water.
- ? In terms of mass: 4.0g of hydrogen gas require 32.0g of oxygen gas to produce 36.0g of water.

b. Formation of Iron (II) sulphide

Iron (II) sulphide is obtained by heating a mixture of iron filings and powdered sulphur:



The following relationships can be obtained from the equation above.

In elementary terms: **One** atom of iron combines with one atom of sulphur to give one molecule of iron (II) sulphide.

Mole-mole relationship: **1 mole** of iron reacts with 1 mole of sulphur to give 1 mole of FeS.

Mass-mass relationship: **56 g** of iron react with 32 g of sulphur to give 88 g of Fe S.

Mass –mole relationship:

1 mole of Fe reacts with 32g of S.

56g of Fe produce 1 mole of FeS.

88g of Fe S are obtained from 1 mole of S, etc.

Worked Examples

The following worked examples will be used to illustrate the steps to be taken when solving problems involving mass- mole relationship from balanced equations of reactions, from the first principle.

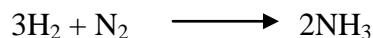
6. In the reaction: $3 \text{ H}_2 + \text{N}_2 \longrightarrow 2\text{NH}_3$
 - a.. How many moles of hydrogen gas will react with 2.0 moles of nitrogen gas?
 - b. How many moles of nitrogen will produce 3.0 moles of ammonia?
 - c. How many moles of ammonia will be obtained from 4.5 moles of hydrogen?

Answer

- a. To calculate the amount of H₂ molecules

Given: 2.0 moles of N₂

This is mole-mole relationship between N₂ and H₂ From the equation of reactions:



3mol 1 mol 2 mol

1 mole of N₂ requires 3 moles of H₂,

∴ 2.0 moles of N₂ will require (3x2.0) moles of H₂

$$= 6.00 \text{ moles.}$$

b. To calculate the amount of N₂

Given: 3.0 moles of NH₃

2 moles of NH₃ are obtained from 1 mole of N₂

∴ 3.0 moles of NH₃ are obtained from (1x3.0)/2 moles

$$= 1.50 \text{ moles of N}_2$$

c. To calculate the amount of NH₃

Given 4.5 moles of H₂

This is mole-mole relationship between H₂ and NH₃ from the equation of reaction

3 moles of H₂ will produce 2 moles of NH₃

∴ 4.5 moles of H₂ will produce (2x 4.5)/3 moles

$$= 3.00 \text{ moles of NH}_3.$$

7. Zinc reacts with copper (II) tetraoxosulphate (VI) according to the following equation:



a. State the type of reaction involved.

b. Determine the amount of zinc that would react completely with 48.0g of copper (II) tetraoxosulphate (VI) [CuSO₄ = 160; Zn=65]

Answer

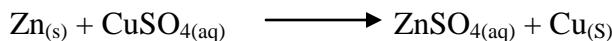
a. It is a displacement of redox reaction.

b. To determine the amount of zinc, in moles

Given: 48.0g of CuSO₄ Required: Amount of Zn.

This is mass- mole relationship between CuSO₄ (in grams) and Zn (in moles).

From the balanced equation of reaction:



1 mole 1 mole

160g

Note: The value given as the molar mass of Cu SO₄, i.e 160g/mol, must be used in the calculations.

160g (1 mole) of Cu SO₄ required 1 mole of Zn.

∴ 48.0g of CuSO₄ will require $(1 \times 48.0)/160$ mol.

$$= 0.300 \text{ mole of Zn}$$

8. Calculate the amount of sulphur (IV) oxide required to deposit 12.0g of sulphur on reacting with excess hydrogen sulphide [S= 32.0; O = 16.0; h=1.0 (WAEC)

Answer

Note: When one of the reactants in a reaction is in excess, it enable the other reactants to be completely used up. There is no need to insert the account or mass of reactant in excess in the balanced equation. (C = 12; O = 16; Ca = 40)

11. In order to determine the percentage of NaHCO₃ in a brand of baking powder, 10.0 g of the sample was heated to constant mass to produce 1.65 g of CO₂ according to the equation:



Calculate the

- mass of NaHCO₃ in 10.0 g of the sample.
- Percentage of NaHCO₃ in the sample

12. When 30.0 g of rusted iron nails reacted with excess of CuSO₄ solution, 0.50 mole of copper was deposited according to the equation:



Calculate the mass of iron in the rusted iron nails. (Fe =56)

Steps 4: Concentration of a solution: Mass-Mole-Liquid Volume Relationships

An aqueous solution of an acid or a base is obtained by dissolving the substance in water.

The amount n of the substance, in moles, in a given volume V of the solution determines its concentration C .

1. When a small quantity of an acid or a base is dissolved in a large volume of water then, the solution obtained is said to be dilute.
2. When a large quantity of an acid or a base is dissolved in a small volume of water, then, the solution obtained is said to be concentrated.

Recall that: **solute + solvent = solution**.

The term concentration is analogous to the term sweetness of a solution. For instance: a solution that contains a (one) cube of sugar dissolved in 100cm^3 of water, will be less sweet, of low concentration or a dilute sugar solution, compared, with a sugar solution that contains 10 cubes of sugar in 100cm^3 of water will be less sweet, of low concentration, or a dilute sugar solution, compared with a sugar solution that contains 10 cubes of sugar in 100cm^{30} of water. It will be sweeter and a concentrated sugar solution.

A standard solution

A standard solution is a solution of known concentration. It is prepared if the substance can be obtained in a pure state and its chemical composition is known.

For instance: a solution known to contain 1,0 g of pure sodium hydroxide, NaOH , in 250cm^3 of solution, is a standard solution.

A standard solution of a pure substance is prepared by weighing accurately a known mass of the substance, dissolving it in water, and making up the solution to a definite volume in a volumetric (or standard) flask.

Units of Concentration

The concentration of a solution can be expressed, quantitatively, in three standard forms:

1. **Mass concentration**, p is the mass m of the solute in grams, present in 1dm^3 of solution. It is expressed in grams per dm^3 (g/dm^3 or g dm^{-3}).

Note: The density, d of a solution can be expressed in grams per cm^3 (g cm^{-3}).

2. **Molar concentration** C is the amount n of the solute, in moles, present in 1dm^3 of solution. It is expressed in moles per dm^3 (mol/dm^3 or mol dm^{-3})

Recall that one mole of a chemical substance is numerically equal to the molar mass M expressed in grams.

Molar concentrate is sometimes called morality M. A solution contains one mole of a substance in 1dm^3 of solution, is called a one-molar solution, 1.0M. M stands for moldm^{-3} . The use of the term morality is obsolete.

3. **Percentage concentration** is the mass m (or volume v) of a substance in 100 grams (or $10\%\text{cm}^3$) of solution. For instance:

a. 10% (m/m) NaOH means a solution, which contains 10 g of pure NaOH in 90 g of water (or 100 g of solution).

B. 10% (m/v) NaOH means 10g of pure NaOH in 100cm^3 of solution

C. 90% (v/v) ethanol means 90cm^3 of ethanol in 10cm^3 of water (or 100 cm^3 of solution).



1 mol 3 mol

$$3 \times 32 = 96\text{g}$$

To calculate the amount of SO_2 required for 12g S

This is mass- mole relationship between mass of S and amount of SO_2

Given: 12.0g of sulphur, Required: Amount of SO_2

96g of sulphur required 1 mol of SO_2

12g of sulphurs require $(1 \times 12)/96$ mol of SO_2

$$= 0.125 \text{ mol } \text{SO}_2$$

Revision Exercises

4. What amount of oxygen will be required for the complete combustion of 3 moles of methane according to the following equation?



A. 2mol B. 3 mol C. 4mol D. 5 mol E. 6 mol. (WAEC)

5. Hydrogen chloride gas reacted with oxygen gas to yield water and chlorine gas. What is the mole ratio of the hydrogen chloride gas to oxygen? (WAEC)

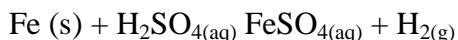
6. In the reaction



- a. Interpret the reaction in terms of moles.
 - b. How many moles of the metal will react with three moles of hydrochloric acid?
 - c. How many moles of hydrocholric acid will produce three moles of hydrogen gas?
7. $2\text{Na}_{(s)} + 2\text{H}_2\text{O}_{(l)} \rightarrow 2\text{NaOH}_{(aq)} + \text{H}_2{}_{(g)}$

From the equation above, calculation the mass of sodium hydroxide produced by 2.3 g of sodium.

- a. 0.40g b. 0.80 g c. 4.00 g d. 8.00 g (H = 1, O = 16, Na = 23) (UTME)
8. Iron reacts with H_2SO_4 according to the equation.



Calculate the mass of FeSO_4 that would be formed by 0.5 mole of Fe. (H = 1, S = 32, Fe = 56) (WAEC)

9. What amount of copper will be deposited when 28g of iron filings react with excess CuSO_4 solution according to the following equation?



- A. 0.10 mol B. 0.50 mol C. 1.00mol D. 2.00mol

Purity of a Reactant

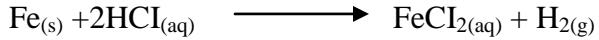
A balanced chemical equation gives the amount of the pure reactants consumed and the amounts of the pure products formed.

If an impure reactant is used in a chemical reaction, the products obtained will be pure. This is because only pure portion of the impure reactants will undergo chemical reaction. The amount of the pure products is then used to determine the extent of purity of, or the quantity of impurity in the reactants.

Worked example

When 10.0 g of razor blade completely reacted with an excess solution of hydrochloric acid, 0.110 mole of FeCl_2 was obtained. Calculate the mass of iron in the razor blade. (Fe = 56)

Answer



1 mol (56 g)

To calculate the mass of iron that will produce 0.110 mol of FeCl₂

This is a mass-mole relationship. Form the balanced equation of reaction:

$$1 \text{ mole of FeCl}_2 = (56 \times 0.110) \text{ g of Fe}$$

$$= 6.16 \text{ g}$$

i.e. 6.16g of pure iron is in 10.0 g of the razor blade.

Revision Exercises

10. Calculate the mass of impurity in a sample of CaCO₃ if 0.050 Mole of CO₂ was formed when 7.0 g of the sample reacted with excess dilute HCl.



The parameters of a solution

The three parameters that are used to describe a solution are: its concentration C; the amount n (or mass m) of solute in the solution; and the volume V of the solution.

Mathematical Relationships between the parameters

1. **Molar concentration:** The S.I. unit is mole per dm³ dm (mol/dm³ or mol dm⁻³)

$$\text{i.e } \quad \mathbf{C} = \mathbf{n/V}$$

$$\text{or } \quad \mathbf{V} = \mathbf{n/C}$$

$$\text{or } \quad \mathbf{n} = \mathbf{CV}$$

2. **Mass concentration:** The S.I. unit is grams per dm³ (g/dm³ or g dm⁻³)

Mass concentration p = Mass m (g) Volume V (dm³)

$$(g \text{ dm}^{-3})$$

3. Molar and mass concentrations are related by the mathematical expression:

Mass conc = Molar conc x Molar mass

$$(\text{g}/\text{dm}^3)(\text{mol}/\text{dm}^3) \quad (\text{g}/\text{mol})$$

Or, molar concentration = $\frac{\text{Mass concentration}}{\text{Molar mass}}$

$$\text{Or, molar mass} = \frac{\text{Mass concentration}}{\text{Molar concentration.}}$$

In order to compare the concentrations of two or more solutions of the same or different chemical substances, their concentrations must be expressed in the same units.

Varying the parameters

Consider three different solutions:

A contains 1 cube of sugar in 100 cm^3 of solution;

B contains 2 cubes of sugar in 100 cm^3 of solution;

C contains 2 cubes of sugar in 200 cm^3 of solution.

Comparing A and B

The concentration (sweetness) of B is twice that of A

Reason: An increase in the amount of solute n in a given solution of volume V. leads to an increase in the concentration C of the resulting solution.

i.e. $C \propto n$ (V constant)

Comparing B and C

The Concentration (sweetness) of B is twice that of C.

Reason: An increase in the volume V of a solution by the addition of pure water, leads to a decrease in the concentration C of the resulting solution, since the amount of solute n before and after dilution is constant.

i.e. $n \propto V$ (Constant)

Comparing A and C

A and C are of the same concentration (sweetness).

Reason: An increase in the amount of solute n in a given solution V requires the addition of more water to the resulting solution, if the concentration C is to remain unchanged.

i.e. $n \propto V$ (C constant)

Calculations involving Aqueous Solutions of Acids and Bases

Titration technique is used to determine the average volume of the aqueous solution of an acid required to neutralize completely, a known volume of the aqueous solution of the base.

Worked examples

10 A solution contains 1.0 g of pure NaOH in 250cm³ of solution. Calculate its concentration in:
a. g dm⁻³ b. mol dm⁻³ (NaOH=40g mol⁻¹)

a. To find the concentration in g dm⁻³, i.e mass of NaOH in 1000cm³ (1dm³) of solution.

Given: 1.0 g of NaOH dissolved in 250 cm³ of solution

required: Mass of NaOH dissolved in 1 dm³ (1000 cm³)

Method 1: From the first principle

The statement that connects the two variables: mass, in grams (required) and volume of solution (given) is:

250 cm³ of solution contain 1.0 g of NaOH (given).

.. 1000cm³ of solution contain (1.0 x 1000)/ 250 g

$$(1.0 \times 4) \text{ g} = 4.0 \text{ g NaOH.}$$

i.e mass concentration in mol dm⁻³ i.e. amount in 1 dm³ of solution.

(From the First Principle)

The statement of fact involves three variables. It is stated in such a way that the required variable is mentioned last. In this problem the required variable is the concentration in mol dm⁻³. The statement of fact is:

40g of NaOH in 1000 cm³ solution is 1.0 mol dm⁻³.

Hence, 1 g of NaOH in 1000cm³ of solution is (1.0 x 1/40) mol dm⁻³

.. 1 g of NaOH in 250 cm³ of solution is (1.0 x 1/40 x 1000/250) mol dm⁻³

$$= (1.0 \times 1/40 \times 4) \text{ mol dm}^{-3}$$

$$= 0.100 \text{ mol dm}^{-3}$$

11. Calculate the mass of anhydrous Na₂CO₃ present in 23.0 cm³ of 0.100 mol dm⁻³ solution (Na₂CO₃ = 106 g mol⁻¹)

(From the First Principle)

The statement of fact is:

100cm³ of 1.0 mol dm⁻³ Na₂CO₃ contain 106g.

Hence, 25cm^3 of 1.0mol dm^{-3} Na_2CO_3 contain $(106 \times 25/1000)$ g.

.. 25 cm^3 of 0.10 mol dm^{-3} Na_2CO_3 contain $(106 \times 25/1000 \times 0.10/1)$ g,

$$= 0.265 \text{ g.}$$

Exercises

1. What is a standard solution?
2. State the difference between a weak acid and a dilute acid; WAEC
3. A solution of sodium hydroide containing 6.0 g in 250 cm^3 of solution has a concentration of A. 0.04mol dm^{-3} B. 0.60 mol dm^{-3} C. 0.96mol dm^{-3} D. 0.15 mol dm^{-3} ($\text{NaOH} = 40 \text{ g mol}^{-1}$)
4. A contains 2.5g of HCl in 250 cm^3 of solution; B contains 8.5g HCl in dm^3 of solution.
Which solution is more concentrated and why?
5. Calculate the
 - a. amount of acid present in 250cm^3 of a 0.20 moldm^{-3} H_2SO_4 .
 - b. Mass of KOH present in 250 cm^3 of a 0.10 mol dm^{-3} solution ($\text{KOH} = 56.0$)
 - c. Volume of 1.0 mol dm^{-3} HNO_3 solution that would contain 0.10 mole of the acid.
6. How many grams of H_2SO_4 will be required for the preparation of 0.175 dm^{-3} of 6.0 mol dm^{-3} solution? n ($\text{H}_2\text{SO}_4 = 98.0 \text{ g mol}^{-1}$)
A. 206g b. 103g C. 98.1g d. 51.5g (WAEC)

CHAPTER 4

ACID/BASE REACTION

4.1 Acid-Base Titration

An acid reacts with a base to form salt and water only. This reaction is a neutralization reaction. TITRATION is the estimation of the volume of an acid, that will neutralize a given volume of a base.

An End Point is the point of neutrality. It is the point at which an acid completely, neutralized the base.

A Standard Solution is a solution whose concentration is known. This is a solution of known concentration. A Buffer Solution is a solution whose pH value does not alter or change for small additions of acid or alkali.

4.2 Apparatus Used in Titration

The following apparatus are used during the titration experiment in the laboratory:

Burette: The burette contains the acid and it is used to measure the volume of acid that will neutralize a given volume of base.

Pipette: It is used to measure the volume of base.

Conical flask: It contains the base. A measured volume of base is poured into it during titration.

Indicators: Indicators are used to know whether a given solution is acidic or basic. They are substances that show one colour in an acid and another colour in a base.

Indicators	Colour in acid	Colour in base	Colour in neutral solution
1. Litmus paper	Red	Blue	Purple
2. Methyl orange	Red	Yellow	Orange
3. Phenolphthalein	Colourless	Pink	Colourless

Other equipments used in acid/base titration are as follows:

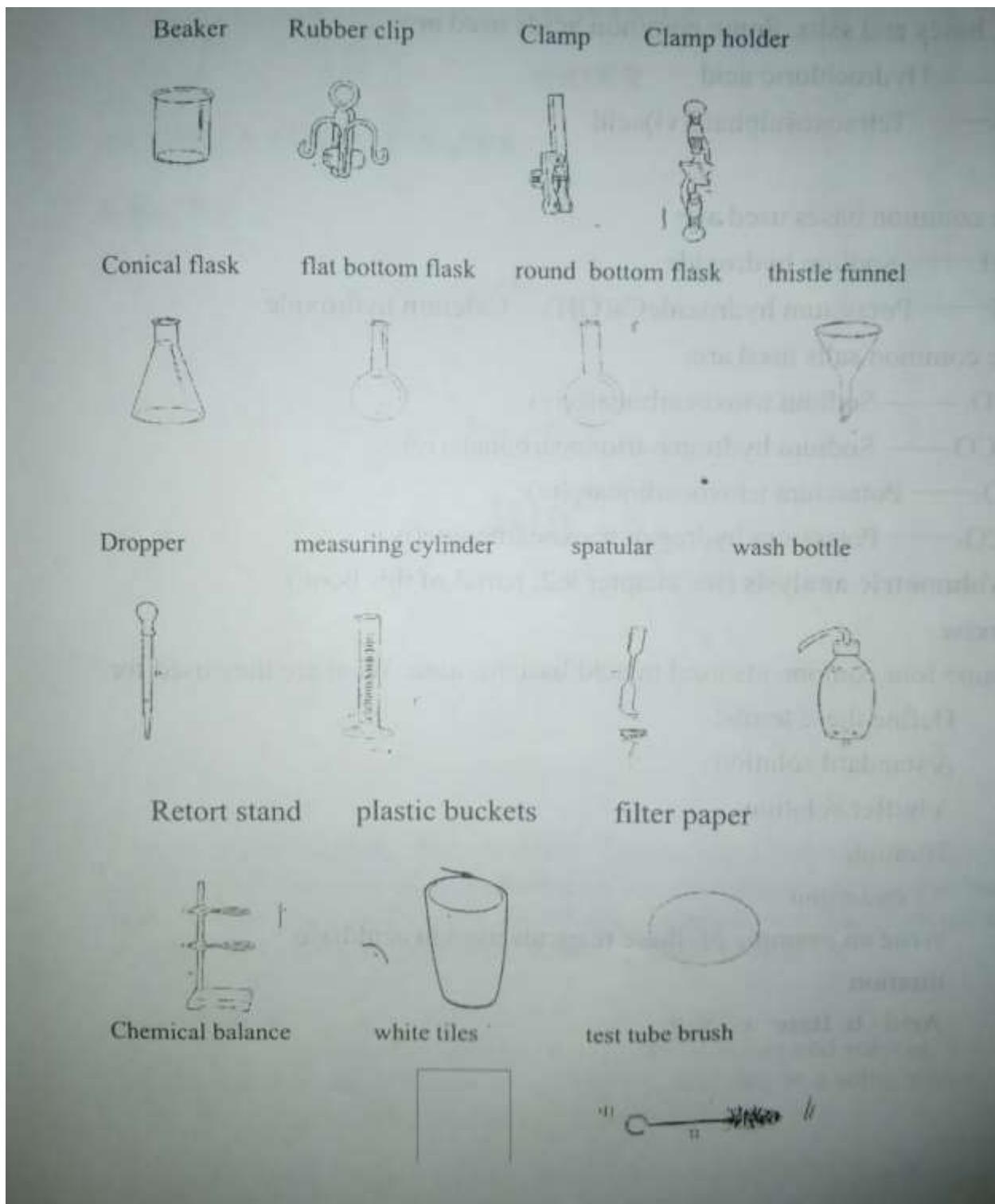


Fig 4.1

4.3 Reagents: Some of the reagents used during acid/base titration are acids, bases and salts. Some common acids used are

HCl-----Hydrochloric acid

$\text{H}_2\text{S0}_4$ -----Tetraoxosulphate(vi)acid

Some common bases used are:

NaOH-----Sodium hydroxide

KOH-----Potassium hydroxide $\text{Ca}(\text{OH})_2$ Calcium hydroxide

Some common salts used are:

Na_2CO_3 -----Sodium trioxocarbonate (iv)

NaHCO_3 -----Sodium hydrogen trioxocarbonate(iv)

K_2CO_3 -----Potassium trioxocarbonate(iv)

KHCO_3 -----Potassium hydrogen trioxocarbonate(iv)

4.4 Volumetric analysis (see chapter 9.2, part 3 of this book)

Exercise

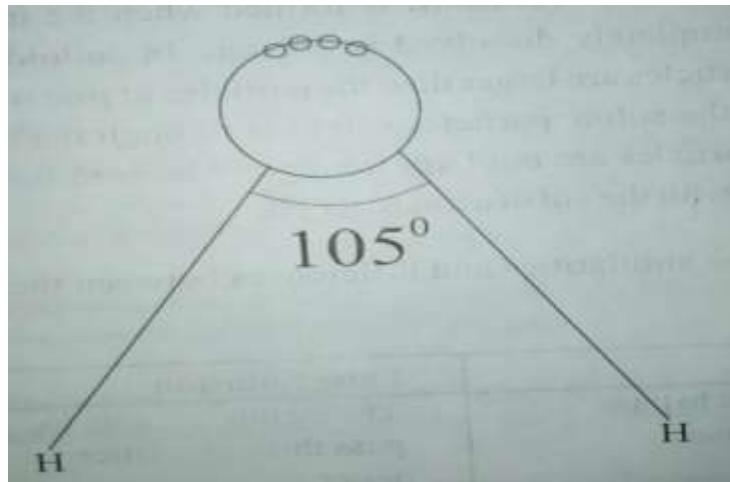
1. Name four equipments used in acid/base titration. What are they used for.
2. Define these terms;
 - a. A standard solution
 - b. A buffer solution
 - c. Titration
 - d. An end-point
3. Write an example of these reagents used in acid/base titration
 - a. Acid
 - b. Base
 - c. Salt

CHAPTER 5

WATER

5.1 STRUCTURE OF WATER

Water, H_2O



The molecular formula of water is H_2O . The molecule of water is V-shaped. In the water molecule the oxygen atom has two bonding pairs of electrons and two lone pairs. Due to mutual repulsion, the two bonding pairs are still brought more closely together. The H-O-H angle is only 105° .

5.2 Solution

Solution is a uniform or homogenous mixture of solute and solvent. Solute is the substance that dissolves in solvent. A solute can be a solid, a liquid or a gas. Solvent is a substance that dissolves solute.

It was pointed out that water is a universal solvent. It is ionic in its molecular structure. When a solid dissolves in water, the resultant solution is called an aqueous solution.

5.3 Types of Solution: There are two types of solution. They are true solution and false solution.

True solution is formed when solute is completely dissolved in solvent. This means that the particles of solid dissolve in such a way that they are able to get in between the solvent particles. The solid particles are called crystalloids because they (the simple molecules, ionic, or particles) are too small to be seen by the naked eye. The solute particles are so small that they can pass through the pores of a filter paper with ease eg brine.

False solution or colloidal solution: is formed when the individual solute particles are not completely dissolved in solvent. In colloidal solution, the individual solute particles are larger

than the particles of true solution. Though (like true solution) the solute particles can pass through the pores of a filter paper. The solute particles are not large enough to be seen by the naked eye. They can be viewed with the aid of a microscope.

The following are the similarities and differences between the true and false solutions:

True Solution	False Solution
1. The solute are able to pass through a filter paper.	The solute are able to pass through a filter paper.
2. The solution does not show light scatter or Tyndall effect.	The solution shows Tyndall effect.
3. Solute can diffuse through a medium eg H ₂ O, air etc	The solute cannot diffuse through a Medium.
4. The solution can be dialysed ie it can pass through a semi-permeable membrane.	The solution cannot be dialysed

NOTE: Tyndall effect means that colloidal particles, unlike the crystalloids of true solutions, are large enough to block the path of light rays and so scatter them.

Dialysis is the passage of both water (solvent) and crystalloids (solute) through a semi permeable membrane.

Suspension: A suspension is a heterogeneous mixture of undissolved particles in a given medium. The particles are large enough to be seen by the naked eye. The particles will eventually settle down if the mixture is left to stand for some time. Suspension can be separated using filter paper.

Examples of suspension, are: muddy water, sulphur in water etc.

5.4 Solubility:

Saturated solution: A saturated solution is a solution which contains as much solute as it can dissolve at that temperature in the presence of undissolved solute particles.

Super saturated solution: A super saturated solution is one which contains more solute than it can hold at that temperature.

Unsaturated solution: An unsaturated solution is a solution which does not contain as much solute as it can dissolve at that temperature.

Solubility: of a solute in a solvent is defined as the maximum amount of the solute in moles or grams that will saturate 1dm^3 of the solvent at a given temperature in the presence of undissolved solute.

Increase in temperature generally increases the solubility of the substances. The solubility of gases decrease with increase in temperature ie Gases are more soluble in cold water than in hot water.

5.5 Determination of Solubility:

Question 1

When 50cm^3 of a saturated solution of sugar at 40°C was evaporated to dryness, 34.2g of dry solid was obtained. What was the solubility of the sugar at 40°C ?

(Molecular mass of sugar is 342g)

Answer:

50cm^3 of solution contain 34.2g

$$1000\text{cm}^3 \text{ will contain } 34.2\text{g} \times \frac{1000}{50}$$

Solubility in gdm^{-3} = 684g

$$\text{Solubility in mol dm}^{-3} = \frac{\text{Solubility in g dm}^{-3}}{\text{Molar mass of sugar}}$$

$$= 684\text{g dm}^{-3}$$

$$342\text{g mol}^{-1}$$

$$= 2 \text{ mol dm}^{-3}$$

Question 2:

If 50g of calcium trioxocarbonate (iv) saturate 500cm^3 of water at 60°C , calculate the solubility of the salt at that temperature. ($\text{Ca} = 40$, $\text{C} = 12$, $\text{O} = 16$).

Answer:

500cm³ of water is saturated by 50g of CaCO₃

$$1000\text{cm}^3 \text{ of water is saturated by } 50 \times \frac{1000}{500}$$

$$= 100\text{g / dm}^3 \text{ or}$$

$$100\text{g dm}^{-3}$$

$$\text{Solubility in mol dm}^{-3} = \frac{\text{Solubility in gdm}^{-3}}{\text{Molar mass of CaCO}_3}$$



$$40 + 12 + (16 \times 3) = 100$$

$$100\text{g} = 1 \text{ mol dm}^{-3}$$

$$100\text{g}$$

Question 3:

If 40g of sodium hydroxide (NaOH) saturates 100cm³ of water at 50°C, calculate the solubility of sodium hydroxide at 50°C

(Na = 23, O = 16, H = 1)

Answer:

100cm³ of H₂O is saturated by 40g of NaOH

$$1000\text{cm}^3 \text{ of H}_2\text{O is saturated by } 40 \times \frac{1000}{100}$$

$$= 400\text{g / dm}^3$$

Convert to mole

$$\text{Mole} = \frac{\text{Mass in g}}{\text{Molar mass}}$$

Molar mass of NaOH

$$23 + 16 + 1 = 40$$

$$\therefore \frac{400\text{g}}{40\text{g}} = 10\text{M}/\text{dm}^3 \text{ or } 10\text{M dm}^{-3}$$

Conversion of mole to gram

To convert mole to gram, multiply the number of mole by the molar mass or the atomic mass.

$$\text{Mole} = \frac{\text{Mass in gram}}{\text{Molar mass}}$$

$$\therefore \text{Mass in gram} = \text{Mole} \times \text{Molar mass}$$

Question 4:

How many grams are there in 0.5 mole of tetraoxosulphate (vi) acid.

(H = 1, S = 32, O = 16)

Answer:



$$(1 \times 2) + 32 + (16 \times 4)$$

$$2 + 32 + 64 = 98$$

$$\begin{aligned}\text{Mass in gram} &= \text{Mole} \times \text{Molar mass} \\ &= 0.5 \times 98 \\ &= 49.0\text{g or } 49\text{g.}\end{aligned}$$

Graph of solubilities:

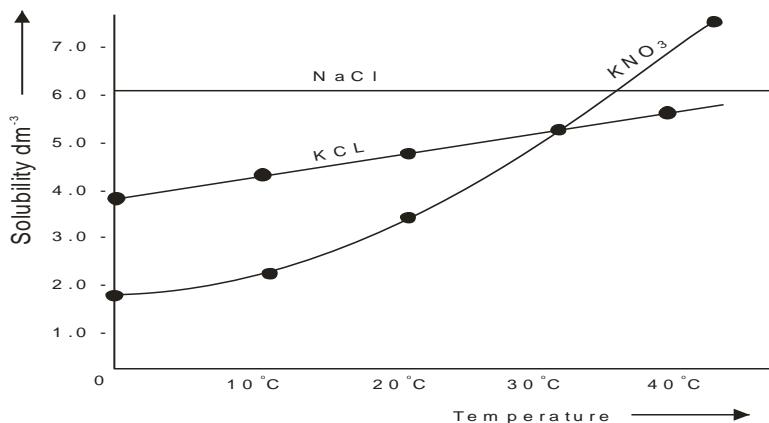


Fig. 5.2: Solubility Curve of NaCl, KCl and KNO₃

Look at the above graph of solubility and answer the following questions.

- i. Which salt does the solubility increases most rapidly with rise in temperature?
- ii. At what temperature does KNO₃ and KCl have the same solubility?
- iii. Which salt shows little or no change in solubility with rise in temperature?

5.6 Hardness of Water and removal of hardness

A hard water is a water that cannot lather easily with soap. Hard water is caused by the presence of soluble metallic salts. The most common of these are hydrogen trioxocarbonate (iv), tetraoxosulphate (vi) and chlorides of calcium and magnesium. Hard water cannot produce lather easily because the salt in water causing the hardness react chemically with soap. The cleansing dirty scum is produced. Lather will not form until all the salts are removed in form of scum.

Types of hardness of water: Two types of hardness of water are: Temporary hardness and Permanent hardness.

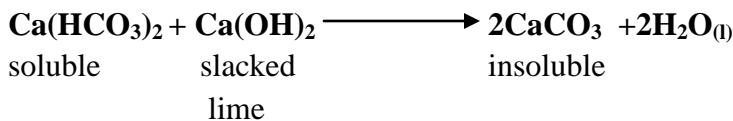
Temporary hardness: Temporary hardness is that hardness of water which can be removed or softened by boiling. Temporary hardness is caused by hydrogen trioxocarbonate (iv) of calcium or magnesium ie (Ca(HCO₃)₂) or Mg(HCO₃)₂.

When the water containing calcium or magnesium hydrogen trioxocarbonate (iv) is heated, the acidic salt decomposes into insoluble calcium or magnesium trioxocarbonate (iv) and carbon (iv) oxide.



CaCO₃ is insoluble and cannot cause hardness of water.

Temporary hard water can also be removed by adding slacked lime ie Ca(OH)₂.

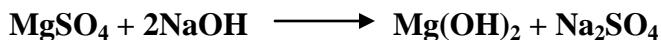
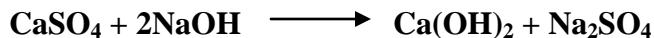


Permanent hardness: Permanent hardness of water is that hardness which cannot be removed by boiling. It is caused by dissolved calcium tetraoxosulphate (vi) or magnesium tetraoxosulphate (vi) ie CaSO₄ or MgSO₄. These salts causing the hardness are not decomposed by boiling. The permanent hardness of water can therefore be removed by the addition of chemicals eg.

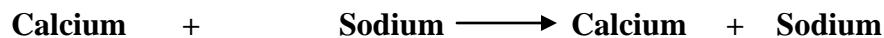
- i. Addition of washing soda: Na₂CO₃.10H₂O



ii. Addition of caustic soda: NaOH



iii. Addition of permutit or zeolite: This is also known as ion exchange process or permitit method. The ion exchanger contains sodium aluminium trioxosilicate (iv) ie NaAlSiO_3 . When the hard water passes through a resin container, the sodium ions enter the solution while the unwanted calcium and magnesium ions take their place in a double decomposition reaction.



The common name for NaAlSiO_3 is sodium zeolite or sodium permuitit.

Advantages of hard water:

It contains calcium compound which is essential for our teeth and bones. Hard water has pleasant taste unlike pure water which is tasteless.

Disadvantages of hard water:

Hard water makes the dyeing of cloths difficult.

It forms scum with soap and wastes soap.

It forms fur or scale deposit in kettles and boiler pipes.

Water cycle:

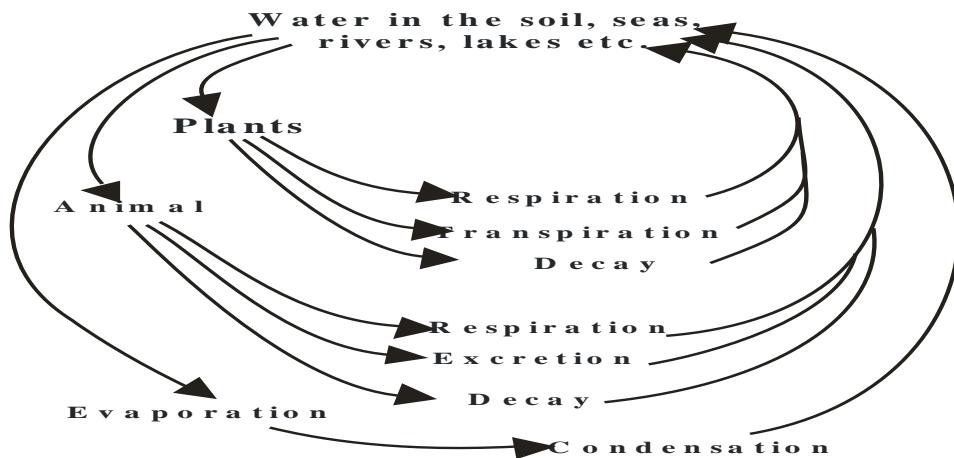


Fig. 5.3: Water cycle

Water cycle is the way water circulates in nature.

5.7 Municipal water supply

Water should be kept clean in order to prevent water-borne diseases such as cholera, diarrhoea and typhoid. Adequate water supply is essential to our health. Clean or pure water does not contain any disease causing germs. It is clear and colourless. It has neither taste nor odour. Pipe-borne water is used in urban areas and in some rural areas. It is the responsibility of the government to provide clean and pure drinking water for its citizens. However, many rural communities contribute to the cost of their pipe-borne water supply through self-help efforts. Pipe-borne water is clean and pure for drinking because it is treated chemically. In some rural areas many people depend on various sources of water supply for their drinking water. These sources include streams, springs, well and rivers. The surroundings of these sources of water supply should be kept clean so as to avoid contamination of water. The water we obtain from most of these sources is not safe for drinking. In order to make the water safe, it must be given water treatment processes.

5.8 Purification of Water

Some examples of treated water are: distilled water, pipe-borne water for township and chlorinated water for use in swimming pools. Treated water is prepared for special purposes.

Pipe-borne water can be purified using the following methods:

Coagulation, sedimentation, filtration and disinfection.

Coagulation: This is a method whereby chemical known as coagulating substances (such as silica gel) is added to the water which has been already tapped from the sources, to remove any colloidal particles.

Filtration: In this method, organic matter is separated from the water.

Sedimentation: Sediments which are formed during the coagulation
Are separated from the bore hole water.

Disinfection: After all these, chemicals are added to the water so that the water will be free from germs. The chemicals used include: silica of soda, hydrated lime, liquid chlorine, alum etc.

Exercises:

1. Define these terms:
 - (a) Solution
 - (b) Solvent
 - (c) Solute
 - (d) True solution
 - (e) False solution
 - (f) Suspension.
2. What do you understand these terms to mean?
 - (a) Solubility
 - (b) Saturated solution
 - (c) Unsaturated solution
 - (D) Super saturated solution
3. If 20g of sodium hydroxide saturates 100cm³ of water at 30°C. Calculate the solubility of sodium hydroxide at 30°C.
(Na = 23, O = 16, H = 1)
4. If 100cm³ of a saturated solution of sugar at 20°C yield 34.2g of the solid sugar on evaporation to dryness, calculate the solubility of the sugar at 20°C.
(Molecular mass of sugar is 342g).
5. Draw the molecular structure of water.
6. Briefly describe two methods of purifying water.
- 7a. Why is water called a universal solvent?
8. Write two chemical properties of water.

- 9a. What is hard water?
10. Write one method of softening:
 - (i) Temporary hard water. (ii) Permanent hard water.
11. Write (a) two advantages of hard water.
(b) two disadvantages of hard water.

CHAPTER 6

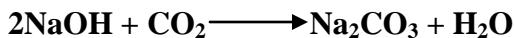
AIR

6.1 Composition of Air

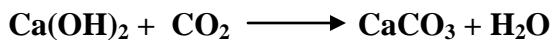
Air is a mixture of gases. The percentage composition of air by volume is as follows. Nitrogen 78%, Oxygen 20%, Carbon (iv) Oxide 0.03%, Rare gases 1%, Water vapour-variable.

6.2 Evidence That Supports that Air is a Mixture

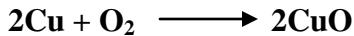
- If air is liquefied at a very high pressure, pure nitrogen and oxygen can be separated out at different temperature by fractional distillation of the liquefied air.
- If air is passed through a solution of sodium hydroxide (caustic soda), the CO₂ in the air is absorbed.



- If Air is passed through lime water, the CO₂ in the air reacts with it and turns limewater milky.



- When air is passed over heated copper, the O₂ in the air oxidizes copper to copper (II) oxide.



These experiments showed that air is a mixture of gases because.

- The constituents of the air can be separated by physical methods.
- After the separation, the constituents of the air still retain their individual properties. For instance, the CO₂ can still turn lime water milky and O₂ can oxidize Cu to CuO.
- If the correct proportion of each of the components of air are mixed together under standard temperature and pressure, they combine physically (not chemically) to form air.
- Air cannot be represented by a chemical formula because it is a mixture and it has slightly variable composition from one environment to the other.

6.3 Nature of Air

Air is matter because (1) it has mass and occupies space.

Experiment to show that air has mass

Tie the mouth of a deflated balloon with a string. Weigh the balloon together with the string on a chemical balance.

Record the mass. Now inflate the balloon and weight it together with the same string. If you compare the two masses, you will see that the mass of the inflated balloon is greater than that of the deflated one. The difference in mass is due to the air inside the inflated balloon.

Experiment to Show that Air Occupies Space

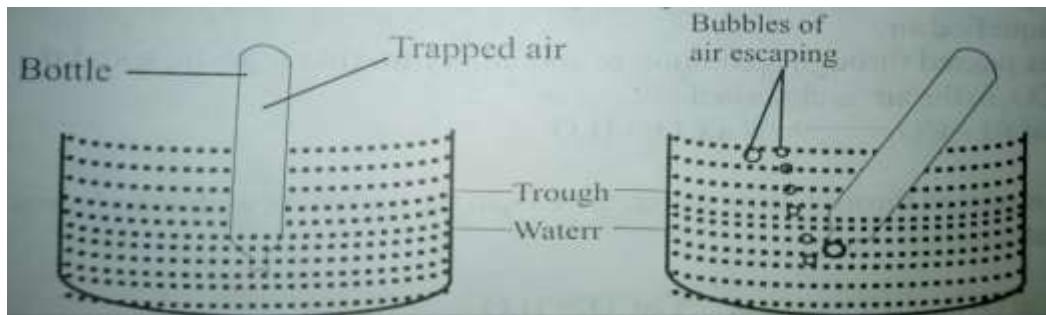


Fig. 6.1(a)

Fig. 6.1(b)

Fill a trough with water. Invert a bottle over it and push it straight down into the water. What do you notice? Now, tilt the bottle in the water gradually. What do you observe?

One will observe that when the inverted bottle is pushed straight down into the water, air is trapped inside the bottle. As a result, the bottle is only partially filled with water and the water level in it is below that in the trough. When the bottle is tilted, air bubbles begin to escape from the bottle through the water in the trough to the surface. At the same time, water gradually enters the bottle to take the place of the escaping air. The bottle eventually becomes completely filled with water when all the trapped air has escaped.

6.4 Properties of Air.

It has been said that air is a mixture of gases, so the properties of air will be the sum of the properties of the individual gases that make it up. The individual gases that make up air are; Oxygen, Nitrogen, Carbon (iv) Oxide, Water Vapour and Rare gases.

- Properties of oxygen; Read Chapter 8, part 2 of this book.
- Properties of nitrogen; Read Chapter 10, part 2 of this book.
- Properties of carbon(iv) oxide; Read Chapter 14, part 1 of this book.

6.5 Bunsen Flame

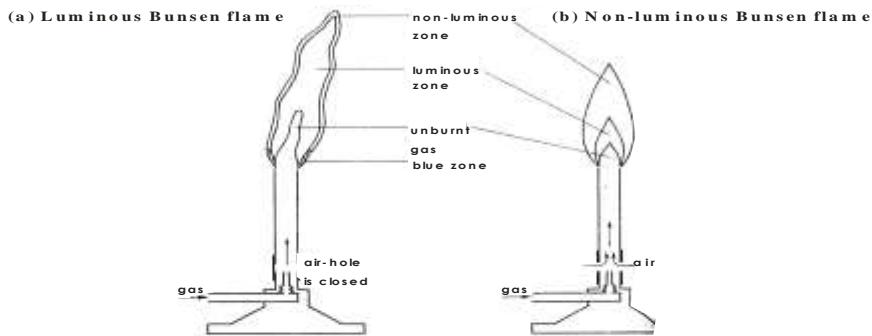


Fig 6.1 Bunsen Flame

An air hole or an air inlet is built at the base of the bunsen burner so that a stream of air can be supplied to the flame together with the fuel gas. The air from this hole and the surrounding air help to allow for a more complete combustion of the fuel. The air-hole can be adjusted so that the flame can be made luminous or non-luminous. The Bunsen burner fuel is a mixture of hydrocarbon, gases, hydrogen and Carbon (II)Oxide. The products of burning or combustion are Carbon (iv) Oxide, water, Carbon (II) Oxide, soot etc. The air-hole at the base of the burner tube should be closed in order to produce a luminous Bunsen flame. In this case, the air supply to the flame for burning depends solely on the atmosphere surrounding the flame. Hence there is a very large bright yellow zone inside the flame where there is incomplete combustion. This results in the luminosity. The flame produced is high and wavy. It is not hot, and there are always large deposits of soot on the surfaces of any object held in it. The flame has four zones just like the candle flame.

The air hole should be kept open, for one to produce a non-luminous Bunsen flame. The air enters through the hole to mix with the fuel. The flame then depends on two sources for its air supply; the air coming in through the air hole and that surrounding the flame. There is then, complete combustion. The flame is non-luminous, hotter, cleaner, and more compact than the luminous Bunsen flame. The flame has only three zones. They are;

1. The unburnt gas zone which is much reduced in size when compared to that of the luminous bunsen flame.
2. The luminous zone or region of incomplete combustion which is also much smaller because of a sufficient supply of air to the inner region of the flame.
3. The outer most non-luminous zone or a region of complete combustion which has increased in size due to the fact that the flame has a sufficient supply of air.

Exercise

- 1a. Define air
- b. What are the constituent of air?
- c. What are the percentage compositions of air?
- d. Is air matter? Explain.
2. Briefly describe how the constituents of air can be separated.
3. What properties can be used to show that air is a mixture of Gases.
- 4a. How many zones has a non-luminous Bunsen flame.
- b. Draw and label it correctly.

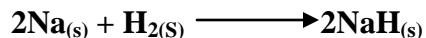
CHAPTER 7

HYDROGEN

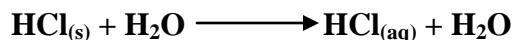
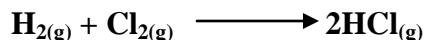
7.1 Hydrogen, Isotopes and Electronic Configuration

Hydrogen is one of the elements in the periodic table. Its atomic number is one and its electronic configuration is $1s1$. Hydrogen exists in three isotopic forms. The three isotopes of hydrogen are protium, deuterium, and tritium. They are represented as H; H or D and H or T respectively. Their relative atomic masses are 1, 2 and 3 respectively. Deuterium and protium are chemically similar except that D is slightly less reactive. Deuterium forms an oxide D_2O , which is very similar to an oxide H_2O , which is very similar to water, H_2O . Tritium is radioactive and is not always found in ordinary hydrogen.

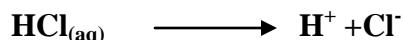
7.2 Uniqueness of Hydrogen: Hydrogen combines with some metals and non-metals to form electrovalent and covalent hydrides or compounds.



Hydrogen reacts with halogens (eg Chlorine) to form covalent chlorides which dissolves in water to form an acid.



In acid hydrogen behaves as a cation.



Hydrogen does not occur free in nature. It combines with several other elements forming water, acids and most organic compounds. It was called “flammable air” because it burns with blue flame to form water. The name hydrogen (i.e water producer) was given to it by Lavoisier. Hydrogen is produced when metals attack dilute mineral acids, water or steam. The readiness with which a metal reacts with an acid or water to liberate hydrogen depends on the position of the metal in the reactivity series. Metals which occur below hydrogen in the series will not liberate it from acids.

Increasing chemical activity	
K	These metals readily liberate H ₂ from dilute HCl and H ₂ SO ₄
Na	
Ca	
Mg	
Al	Al release hydrogen only with hot conc. HCl
Zn	
Fe	
Sn	
Pb	
(H)	
Cu	
Hg	
Ag	
Au	
K	Attack cold water to form alkalis and hydrogen gas
Na	
Ca	
Mg	
Al	Attack steam to form oxides and hydrogen gas
Zn	
Fe	
Sn	
Pb	
(H)	
Cu	No reaction with either cold water or steam
Hg	
Ag	
Au	

7.3 Laboratory Preparation of Hydrogen

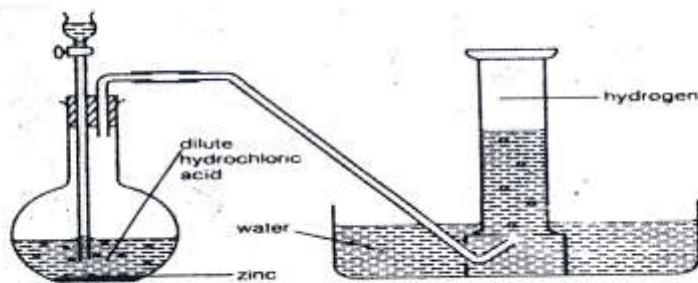
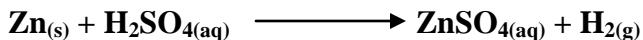
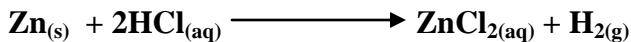


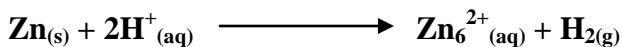
Fig. 7.1

Hydrogen is prepared in the laboratory by the action of dilute HCl on Zinc metal. Put some pieces of zinc into a flat bottomed flask. Add dil. HCl by means of a thistle funnel. There is effervescence and a gas is given off which is collected over water. If the Zn metal is pure (or if the reaction is slow), add few drops of CuSO₄ solution to catalyze the reaction. The dry gas is obtained by passing it over fused CaCl₂ or conc.

H₂SO₄ and is collected over mercury or by the downward displacement of air.



The above reactions can be represented ionically as:

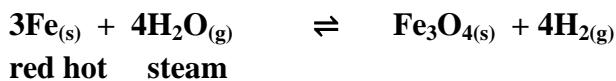


HNO₃ acid is not used in this reaction because of its oxidizing character. It will oxidize H₂ to water.

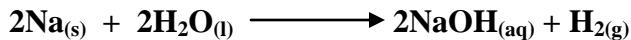
Note: Naked flames should not be left nearby because a mixture of H₂ and air explode when heated.

Other methods of obtaining hydrogen are:

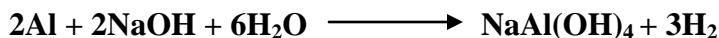
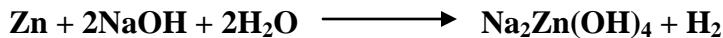
1. By the action of steam on heated iron.



2. By the action of Na on cold water. The reaction is vigorous.



3. Zinc and aluminium can react with caustic alkalis to form hydrogen



KOH can be used in place of NaOH

7.4 Industrial Preparation of Hydrogen

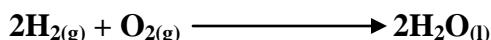
Hydrogen can be prepared in a commercial quantity by the electrolysis of acidified water (Read chapter 7.4 of part II). Hydrogen is also obtained as a by-product in the electrolysis of brine.

7.5 Test For Hydrogen

A mixture of hydrogen and air explode with a pop sound when flame is applied to it.

7.6 Properties of Hydrogen

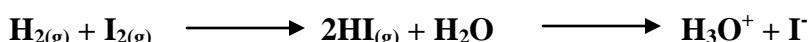
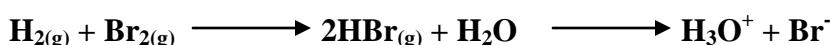
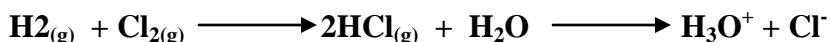
1. It is a colourless, odourless and tasteless gas.
2. It is lighter than air.
3. It liquidifies at -253°C
4. It is insoluble in water
5. It has no effect on the moist litmus paper
6. Hydrogen burns in air to form water



7. Hydrogen reduces the oxides of some metals (ZnO , PbO , CuO) to their respective metals while it is oxidized to water e.g

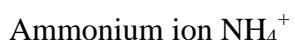
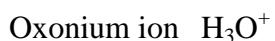


8. H_2 reacts with halogen to form halides

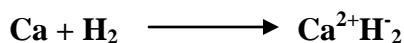
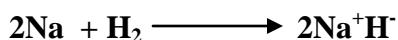


From these reactions we can deduce that;

- i. Hydrogen can form a covalent molecule e.g. HCl molecule, H-Cl, HBr, H-I and H-H.
- ii. Hydrogen can form a coordination compound with another molecule possessing a lone pair of electrons e.g.



9. Hydrogen can react with highly electropositive metals (e.g. Na, K, Li, Ca) to form hydrides.



Here hydrogen behaves as a non-metal by accepting electron from the more active metals.

Hydride is a binary compound of hydrogen. E.g Sodium hydride, NaH, Potassium hydride, KH, Calcium hydride, CaH₂

Uses of Hydrogen

1. It is used in the preparation of ammonia (Haber process)
2. It is used as fuel for industrial heating e.g. water gas
3. It is used in the manufacture of margarine. Vegetable oil is hydrogenated to margarine in the presence of Ni (Catalyst).

Exercises

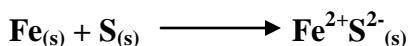
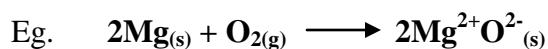
1. Use chemical equations only to show two methods of generating H₂ gas in the school laboratory
2. Write two:
A. Physical properties of hydrogen
B. Chemical properties of hydrogen
3. List the three isotopes of hydrogen and write their relative molecular masses.
4. What are the uses of hydrogen?

CHAPTER 8

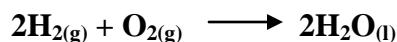
OXYGEN

8.1 General Properties of Oxygen Group

Elements in group 6 includes oxygen and sulphur which are non-metals. They are oxidizing agents because, they are electron acceptors.



Unlike sulphur, oxygen is soluble in water. Oxygen and sulphur do not react with water. Both oxygen and sulphur combine directly with hydrogen to form water and hydrogen sulphide respectively.



8.2 Electronic Structure and Bonding Capacity of Oxygen

Oxygen has the atomic number of 8. Its number of proton is 8. Its number of electrons is also 8. The atomic structure of oxygen is:

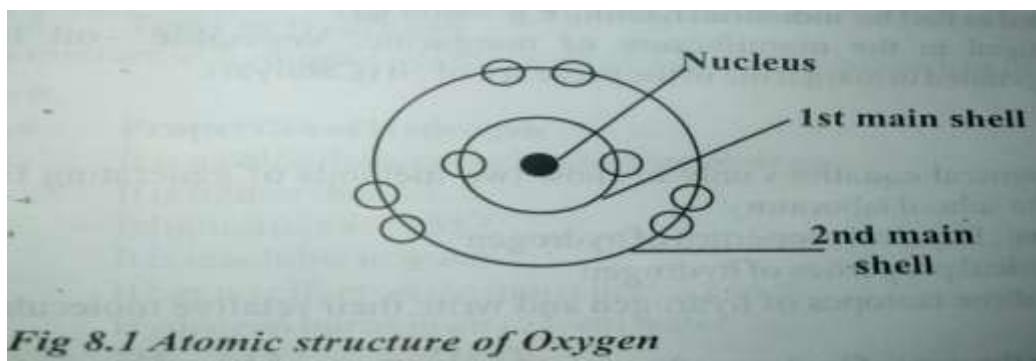


Fig 8.1 Atomic structure of Oxygen

The electronic configuration of oxygen is $1s^2, 2s^2, 2p^4$. Oxygen requires two electrons to attain octate state in its outermost main shell. It can also be said that oxygen requires two electrons to complete the maximum number of electrons, P-orbitals can carry. P-orbitals can carry a maximum of six electrons. An Oxidation number of Oxygen is -2. Its bonding capacity is 2 and its valency is also 2.

8.3 Laboratory Preparation of Oxygen

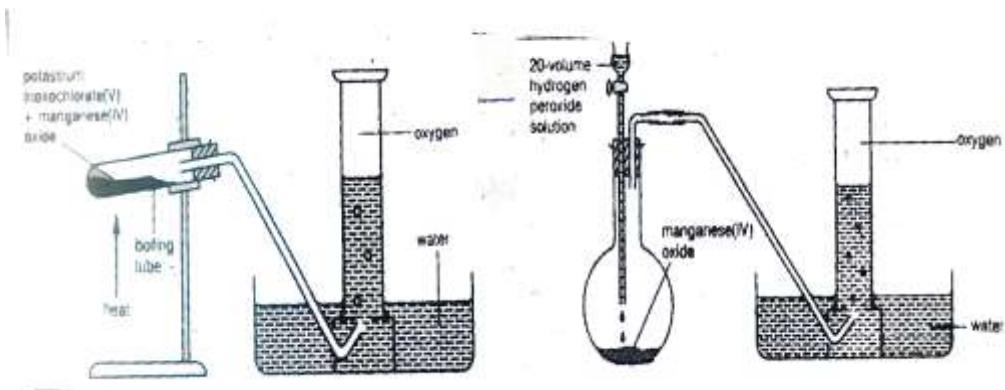
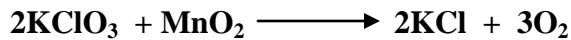


Fig 8.2 Preparation of oxygen from potassium trioxochlorate (iv)

Oxygen is prepared in the school laboratory by heating a mixture of potassium trioxochlorate (v) and manganese (iv) oxide. The latter is added as a catalyst to quicken the rate of decomposition of potassium trioxochlorate (v) when it is heated. The mixture is placed in hard glass tube and heated. A colourless and odourless gas given off is collected over water. Oxygen cannot be collected by displacement of air because it has about the same density as air.



this can be represented ionically as:

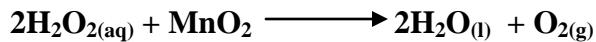


Other methods include the following:

Oxidation of hydrogen peroxide, H_2O_2 with acidified potassium tetraoxomanganate (vii), KMnO_4 . No heating is required for this reaction to occur.



Catalytic decomposition of H_2O_2 . No heating is required here too.



The dried gas, O_2 is obtained by passing it through conc. H_2SO_4 in a wash bottle or through anhydrous calcium chloride, CaCl_2 in a U-tube.

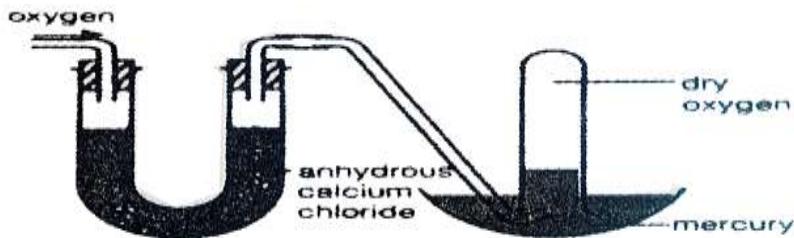


Fig 8.4 Drying oxygen

Industrially, oxygen is prepared by:

- i. Electrolysis of acidified water
- ii. The fractional distillation of liquid air

8.4 Fractional Distillation of Liquid Air

Fractional distillation of liquid air involves (a) the liquefaction of air and (b) the actual distillation of the liquid air in a fractionating column. In the liquefaction process, air is made to pass through caustic soda, which absorbs CO_2 in it. It is then compressed to about 200 atmospheres and cooled. The air is forced into a bigger chamber of low pressure. This causes continuous decrease in the temperature of the air which liquefies at 200°C . The liquid air is then channeled into a fractionating column. Heat is supplied to it. At -196°C , nitrogen gas is evolved. The temperature rises and at -183°C , the liquid oxygen with the higher boiling point vaporizes and it is collected and dried as usual.

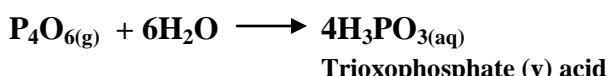
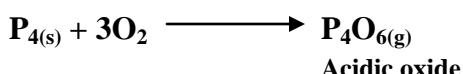
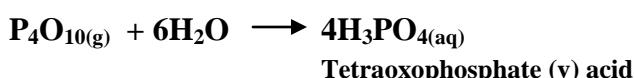
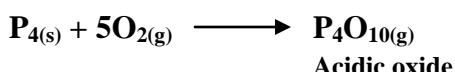
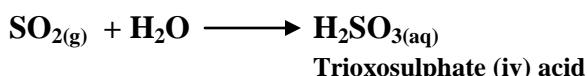
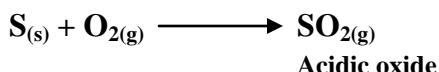
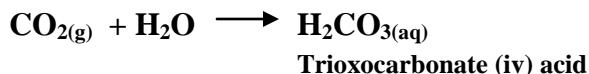
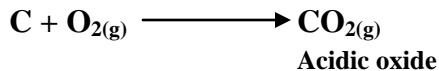
8.5 Physical Properties of Oxygen

1. Oxygen is a colourless, odourless and tasteless gas.
2. Oxygen gas exists as a diatomic molecule
3. It is about 1.1 times heavier than air.
4. It dissolves very slightly in water and has no effect on the moist litmus paper.

8.6 Chemical Properties of Oxygen

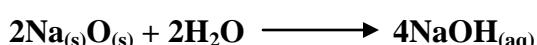
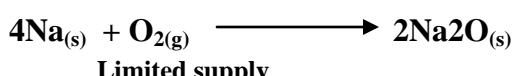
Oxygen is very reactive with some non-metals, metals and compounds.

Reaction with non-metals: It combines with carbon, sulphur and phosphorus to form their corresponding acidic oxides. Acidic oxides are also called acid anhydrides because their dissolution in water results in the formation of acids.

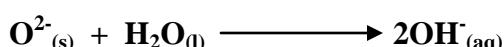
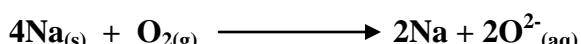


Reaction with metals: Oxygen combines with highly electropositive metals like Na, K and Ca to form their corresponding

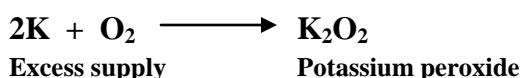
Basic oxides. Their oxides dissolve in water to form alkalis.



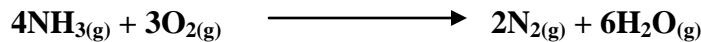
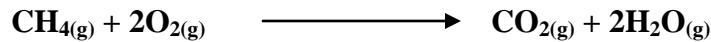
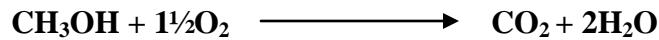
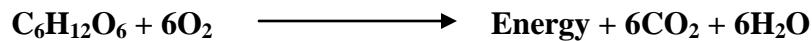
or



If Na and K are heated in an excess supply of oxygen, they burn with a bright yellow flame and lilac flame respectively to form their higher oxides instead of their basic oxides.

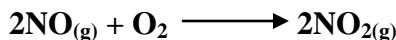


Reaction with some compounds: Oxygen combines with some compounds to form a variety of products.



8.7 Test For Oxygen

1. O₂ rekindles a glowing splint of wood. This distinguishes it from all gases except dinitrogen (i) oxide, N₂O. oxygen is distinguished from this gas by:
 - a. Having no smell. (N₂O has a pleasant or sweet smell).
 - b. Oxygen reacts with nitrogen (ii) oxide, NO to produce brown fumes of nitrogen (iv) oxide, NO₂.



Dinitrogen (i) oxide has no effect on nitrogen (ii) oxide.

Uses of Oxygen

1. Ethyne gas is made to burn in oxygen to produce oxy-ethyne flame for welding and cutting of metals at 2,200°C.
2. A mixture of oxygen and an esthetic is given to a patient during surgical operations.
3. It is used as an aid to breathing where the natural supply of oxygen is insufficient, e.g people at high altitude like pilots and astronauts in space make use of oxygen tents.

8.8 Classification of Oxides

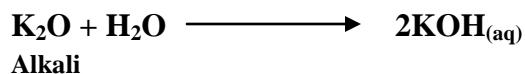
Oxide is a binary compound of oxygen e.g Zinc oxide, ZnO and calcium oxide, CaO.

Four important classes or types of oxide are the following:

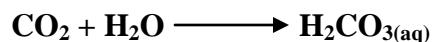
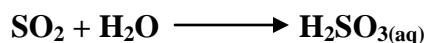
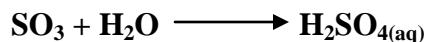
- i. Basic oxide
- ii. Acidic oxide
- iii. Amphoteric oxide

iv. Neutral oxide

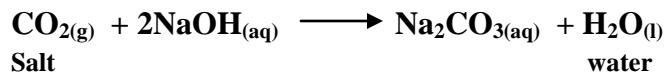
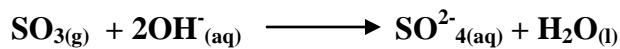
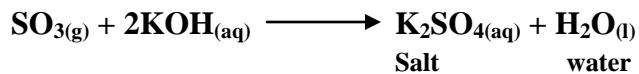
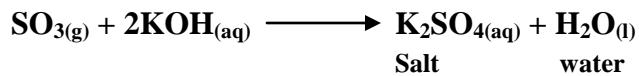
A Basic Oxide: is a metallic oxide. It reacts with an acid to form a salt and water only. Soluble metallic oxide (basic oxide) is called an alkali. Examples of alkali are calcium oxide, CaO; sodium oxide Na₂O; Potassium oxide, K₂O.



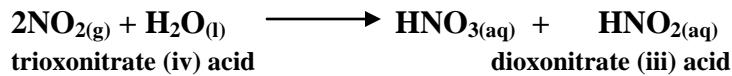
An acidic oxide: is a non-metallic oxide which when combines with water produces an acid e.g. SO₃, SO₂, CO₂ etc.



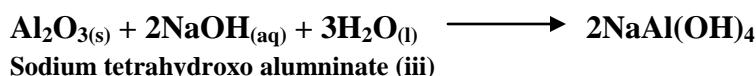
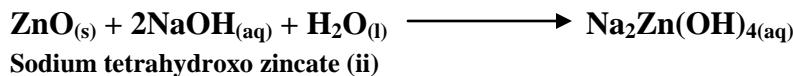
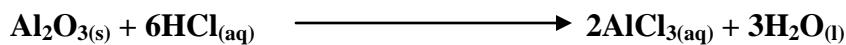
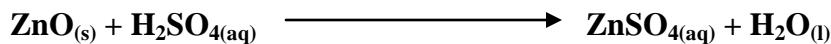
Acidic oxides also reacts with alkalis to form a salt and water only



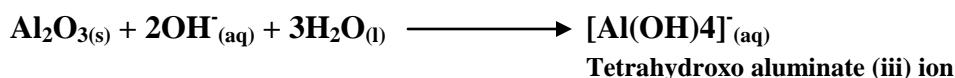
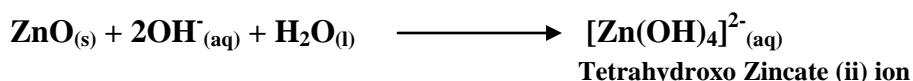
Acidic oxide is also called acid anhydride but nitrogen (iv) oxide, NO₂ is known as a mixed acid anhydride because it reacts with water to form two types of acid. It also reacts with caustic soda to form two types of salt.



Amphoteric oxide: is a metallic oxide which can show both basic and acidic properties. This means that it can react with both acid and alkali to produce a salt and water only.e.g ZnO and Al₂O₃



Na₂Zn(OH)₄ and NaAl(OH)₄ are complex salts. Its formation can be represented ionically as:

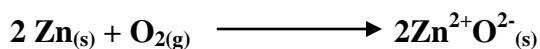
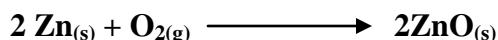


A neutral oxide: is an oxide which shows neither basic nor acidic character. In other words, neutral oxide is neutral to litmus. Examples of these oxides are: dinitrogen (i) oxide, N₂O, Carbon (ii) oxide, CO and water, H₂O.

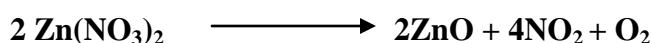
8.9 Zinc Oxide, ZnO: this compound is a white powder. It is yellow when hot

It is prepared by the following methods:

- i. Burning zinc in air



- ii. Dissolving Zn in dilute HNO₃ acid. Evaporating the Zn(NO₃)₂ solution so formed to dryness and heating the residue strongly.



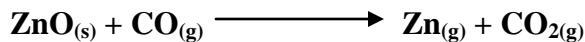
- iii. Thermal decomposition of zinc trioxocarbonate (iv).



- iv. Heating zinc hydroxide



ZnO is amphoteric. It dissolves in both acids and alkalis. It cannot be converted into metal by heating the oxide in a stream of hydrogen because hydrogen is lower than Zn in the electrochemical series. ZnO is reduced to Zn by heating it with coke or carbon (ii) oxide.



ZnO is used as a white pigment in paints. It is used as a filler in the manufacture of rubber. Pure ZnO is used in the making of cosmetic powders and creams.

THE BASIC OXIDES

K
Na
Ca } Oxides of these metals are soluble in water forming alkalis

Mg
Al
Zn
Fe } Oxides of these metals can be made from the metal by the action of trioxonitrate
Pb
Cu } (v) acid and then heat (Al is an exception)

Hg
Ag
Au } Oxide of these metals decompose when heated.

Exercises

1. Use chemical equations only to show two methods of generating H₂ gas in the school laboratory
2. Write two:
 - Physical properties of hydrogen
 - B. Chemical properties of hydrogen

3. Write two balanced chemical equations only to show how oxygen can be prepared in the laboratory.
4. What are the two suitable substances that can be used to dry oxygen gas?
5. Briefly describe how oxygen can be prepared in a commercial quantity.
6. Write two:
 - i. Physical properties of oxygen
 - ii. Chemical properties of oxygen
7. Oxides are classified into four.
 - i. what are they?
 - iii. Give an example of each.
8. What are the two importance of ZnO?

CHAPTER 9

HALOGENS

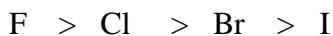
9.1 Electronic Configuration of Halogens

Halogens are in group 7 or VII in the periodic table. They require one electron each to attain an octate state. Their electronic configuration are as follows;

Main shell Subshell or orbital			1 s	2 s p	3 s p d	4 s p d f
Halogen	Symbol	Atomic Number				
Florine	F	9	2	2,5		
Chlorine	Cl	17	2	2,6	2,5	
Bromine	Br	35	2	2,6	2,6,10	2,5
Iodine	I	53	2	2,6	2,6,10	2,6,10,7

9.2 Properties of Halogens and Gradation Down the Group

Halogens are the non metals that form the group 7 of the periodic table. The halogens are Flourine, chlorine, Bromine, Iodine and Astatine, Halogens are sometimes called “Salt producer” because they react with metals to produce salts. They have 7 electrons each in their outermost shell and exhibit a co-valency of I. Flourine is the first gaseous member and also the most reactive of all. It can even react with the test tube. Their decreasing order of reactivity are as follows:



Halogens exist as diatomic molecules; F_2 , Cl_2 , Br_2 and I_2 with only weak Vander Waals forces holding the individual molecules together. These forces are broken down on warming. Iodine undergoes sublimation on heating. Being electron acceptors, the halogens are strong oxidizing agents. Halogens are univalent. They form electrovalent compounds with metals and covalent compounds with hydrogen. The hydrides of halogens are gases at room temperature. They dissolve in water to form strong acids.

9.3 Laboratory Preparation of Chlorine

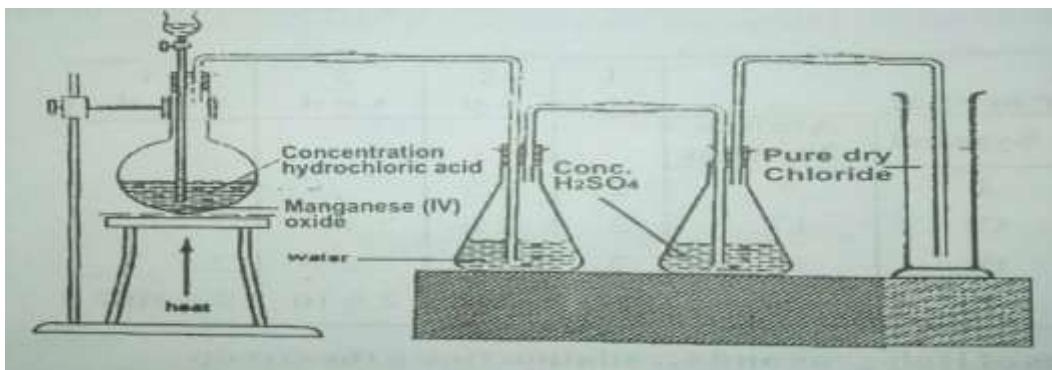


Fig. 9.1

Cl_2 is prepared in the laboratory by oxidation of conc HCl using the oxidizing agents - Manganese (iv) oxide (MnO_2) or Potassium tetraoxomanganate (vi). The reaction is as follows:

- i. $\text{MnO}_{2(s)} + 4\text{HCl}_{(aq)} \longrightarrow \text{MnCl}_2 + \text{H}_2\text{O} + \text{Cl}_2$
- ii. $2\text{KMnO}_{4(aq)} + 16\text{HCl}_{(aq)} \longrightarrow 2\text{MnCl}_{2(aq)} + 2\text{KCl} + 8\text{H}_2\text{O} + 5\text{Cl}_2$

Reaction (i) requires the application of heat unlike the reaction (ii).

MnO_2 is placed in the round bottomed flask. Conc. HCl is poured into the flask. Heat is applied to the mixture. Effervescence occurs and a greenish yellow gas is evolved with misty fumes of hydrogen chloride gas. The gas is passed through the first and second wash bottles. The H_2O in the first wash bottle absorbs the HCl gas. The conc. H_2SO_4 in the second wash bottle dries the gas. The gas is then collected by downward delivery or upward displacement of air.

In the above two reactions, conc. HCl acid is oxidized to Cl_2 .

Industrial preparation of Cl_2 :

Cl_2 is prepared in a commercial quantity by the electrolysis of brine.

Read Part II, Chapter 6.3 of this book. This is the electrolysis of brine using simple electrolytic cell. It is unfortunate to note that this simple cell results in a reaction between the Cl_2 produced and the resultant NaOH solution (NaOH formed in the cell after the electrolysis) to form sodium oxochlorate (i).

Two types of cells have therefore been devised for the commercial or industrial production of Cl_2 . They are:

- a. The mercury cathode cell
- b. The Diaphragm cell.

The mercury cathode cell:

Electrodes: The anode must be a graphite while the cathode is made of mercury. Ions present from

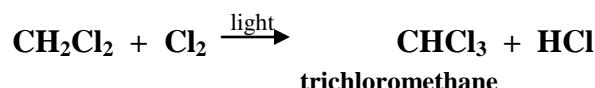
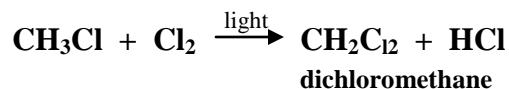
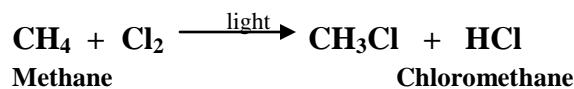


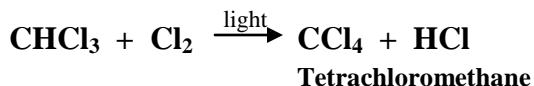
Cathode	Anode
<p>Both Na^+ and H^+ migrate to the cathode. Na^+ is discharged because, Hg likes to associate with it to form sodium amalgam</p> $\text{Na}^+ + \text{Hg}_{(l)} \longrightarrow \text{Na/Hg}$ <p>Sodium amalgam (Na/Hg) flows out into a trough where it combines chemically with H_2O to form NaOH and H_2. Mercury is once more recovered.</p> $\text{Na/Hg} + \text{H}_2\text{O} \longrightarrow \text{Hg} + \text{NaOH} + \text{H}_2$	<p>Both Cl^- and OH^- migrate to the anode. Cl^- is discharged due to concentration effect.</p> $\text{Cl}^- - e^- \longrightarrow \text{Cl}$ $\text{Cl} + \text{Cl} \longrightarrow \text{Cl}_2$

It is more preferable to use mercury Cathode cell than simple cell in the industrial preparation of Cl_2 . Here the quantity of Cl_2 produced is recovered. Cl_2 cannot react with NaOH because Na^+ associates with Hg to form Na/Hg .

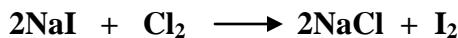
9.4 Physical and Chemical Properties of Chlorine

1. It is a poisonous greenish yellow gas. It is denser than air. It is slightly soluble in water. It has a choking and irritating smell.
2. Cl_2 undergoes substitution reaction with saturated hydrocarbons. Light energy catalyses the reaction.

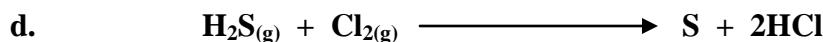
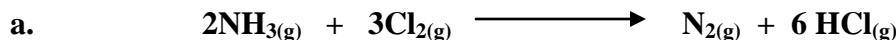
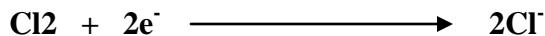




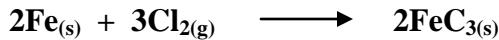
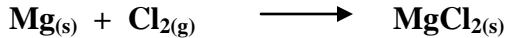
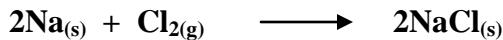
3. Cl₂ being more reactive than Br₂ and I₂ displaces them from their salts



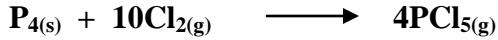
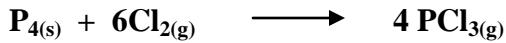
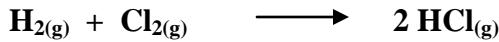
4. Chlorine being an electron acceptor is a strong oxidizing agent.



5. Chlorine combines with metals to form their corresponding chlorides



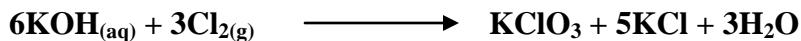
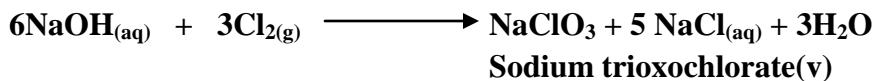
6. Chlorine reacts readily and directly with all non-metals except inert gases, oxygen, nitrogen and carbon.



- 7a. Cl₂ reacts with cold dilute alkali to form its corresponding oxochlorate (1) and its metallic chloride

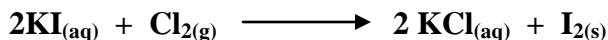


- b. Cl₂ reacts with hot concentrated alkali to form a mixture of the metallic trioxochlorate (v) and the metallic chloride.



Test for Cl₂ gas:

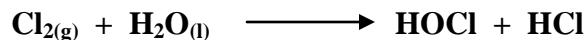
1. Colour - Greenish yellow
2. Cl₂ gas turns damp blue litmus paper red and then bleaches it.
3. Chlorine turns starch iodide paper blue.



The Cl₂ displaces Iodine from iodide paper. The free iodine then turns the starch blue.

Bleaching action of Cl₂

Cl₂ reacts with H₂O to form oxochlorate (i) acid.



The oxochlorate(i) acid is very reactive. It readily gives up its oxygen to the dye to form a colourless compound



After bleaching, the bleached material is treated with sodium trioxosulphate(vi) (Na₂S₂O₃) to remove any remaining Cl₂. It is then washed with water to remove HCl acid which is always formed when Cl₂ bleaches.

Uses of chlorine

1. It is used in the sterilization of water.
2. In the industry, it is used as a bleaching agent and in the commercial preparation of hydrochloric acid.
3. Cl₂ is used as a disinfectant and germicide.
4. It is used in the manufacture of the bleaching powder (CaOCl₂.H₂O). CaOCl₂.H₂O is used as disinfectant in latrines and refuse dumps.
5. It is also used in the production of chloroethane which is used in the making of polychloroethane (plastic).

9.5 Laboratory Preparation of Hydrogen Chloride

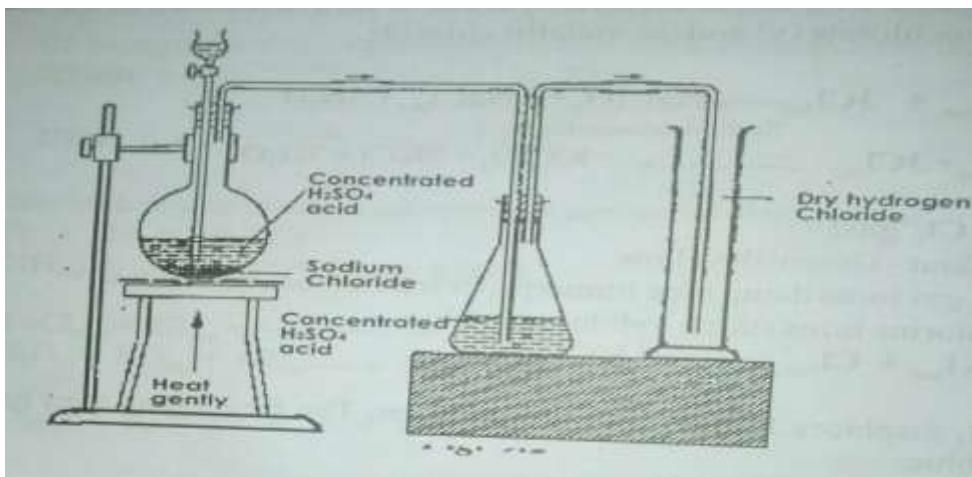
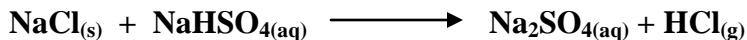
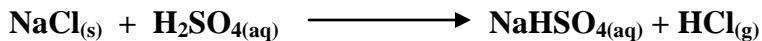
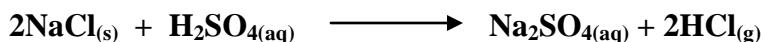


Fig. 9.2

The apparatus used is as shown above. The common salt placed in the round bottomed flask is added concentrated tetraoxosulphate (vi) acid through the thistle funnel. The mixture is warmed. Effervescence occurs and misty fume is observed. The gas is dried by passing it through a concentrated tetraoxosulphate (vi) acid. It is collected by downward delivery since it is heavier than air.



The overall equation is represented as:



Ionically.

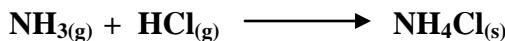


Industrially hydrogen chloride gas is prepared by direct combination reaction between hydrogen and chlorine gases obtained from the electrolysis of sodium chloride solution (brine).

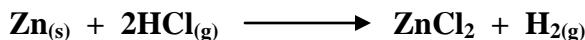
Properties of hydrogen chloride gas:

1. Pure hydrogen chloride is a colourless gas with irritating smell.
2. It dissolves in water to form an aqueous solution of hydrochloric acid.
3. It turns damp blue litmus paper red.

4. When ammonia is brought near a gas jar of hydrogen chloride, white dense fumes of ammonium chloride are produced.



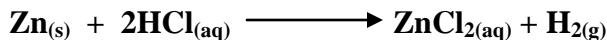
5. Hydrogen chloride reacts with many hot metals to liberate hydrogen and form the metallic chloride.



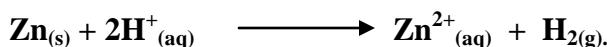
9.6 Properties of Hydrochloric Acid

Hydrochloric acid exhibits all the properties of acids.

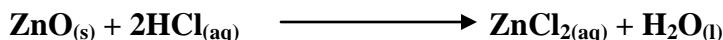
- Concentrated hydrochloric acid is corrosive to organic materials like skin, cloth, paper, wood etc. It turns blue litmus paper red.
- Metals which are higher than hydrogen in the activity series reacts with hydrochloric acid to liberate hydrogen gas.



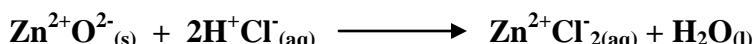
Ionomically.



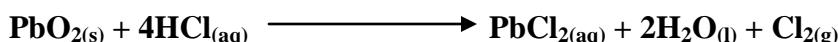
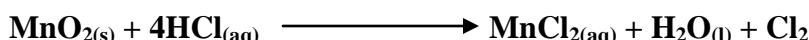
- It reacts with a base or an alkali to form a salt and water only.



Ionomically.



- Strong oxidizing agents like manganese (iv) oxide, potassium tetraoxomanganate (vii) and lead(iv) oxide react with the acid and oxidize it to chlorine.



Ionomically:





See the laboratory preparation of chlorine in chapter 9.1 of part II.

5. Hydrochloric acid acts as a precipitating agent. It precipitates the insoluble AgCl and PbCl₂ from the solutions of their trioxonitrate (v) salts.



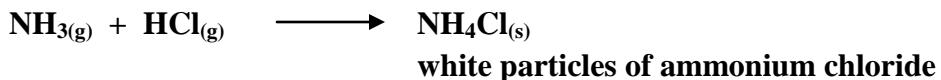
Test for hydrogen chloride:

An unknown gas is suspected to be hydrogens chloride if:

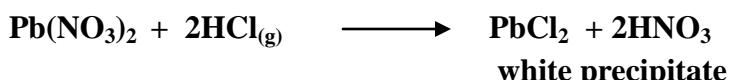
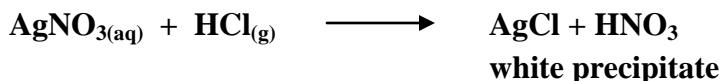
- i. The gas is colourless and has an irritating smell.
- ii. It fumes in moist air and turns blue litmus paper red.

After these preliminary tests, confirm the gas with these two reactions.

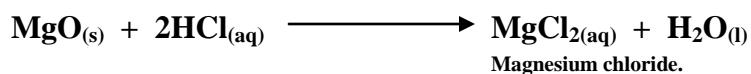
1. Reaction with ammonia.

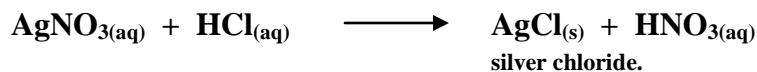


2. Reaction with AgNO₃ or Pb(NO₃)₂.



Chlorides are the binary compounds of chlorine. They are formed from various methods.





Uses of hydrochloric acid:

1. It is used in the manufacture of chlorine.
2. It is used in the pickling of metals. That is, it is used in the clearing of metal surfaces (to remove oxides from metals) prior to electroplating.
3. It is used as a precipitating agent in the qualitative analysis.
4. It is used in the preparation of chlorides.

Exercises:

- 1a. What are halogens?
- b. Write four examples of halogen.
2. Write a balanced chemical equation to show the laboratory preparation of chlorine.
3. Write two physical properties of chlorine.
4. Write three chemical properties of chlorine.
5. What are the uses of chlorine?
6. What are three tests that can be used to identify chlorine gas?
7. Write four general properties of halogens.
8. Draw the electronic configuration of (i) Fluorine (ii) Bromine (iii) Chlorine
9. Briefly describe the laboratory preparation of hydrogen chloride.
10. Use a balanced chemical equation only to show the reaction of hydrochloric acid
 - i with metal
 - ii. with an alkali
 - iii. With an oxidizing agent.
- 11a. Write three uses of hydrochloric acid.
 - b. What is a chloride?
 - c. Write two named examples of chlorides.

CHAPTER 10

NITROGEN

10.1 General Properties of Group VA Elements

Nitrogen and phosphorus are group VA elements. They show very similar properties because they have the same atomic number. They are electron acceptors. They form many covalent oxides with oxygen. Some examples of their oxides are; dinitrogen (iii) oxide, N_2O_3 ; phosphorus(iii)oxide; P_4O_6 ; dinitrogen(v)oxide, N_2O_5 and phosphorus(v)oxide, P_4O_{10} . These oxides react with water to form strong acid. These oxides are acidic. Nitrogen and phosphorus form similar hydrides and chlorides. These are; NH_3 , PH_3 and NCl_3 and PCl_3 . Group VA elements are also referred to as nitrogen family group.

Air is a mixture of gases. It is made up of nitrogen, oxygen, carbon (iv) oxide and the rare gases. The percentage by volume of these gases are:

Nitrogen	-	78%
Oxygen	-	21%
Carbon (iv) oxide	-	0.03%
Rare gases	-	1%

10.2 Laboratory Preparation of Nitrogen from Air

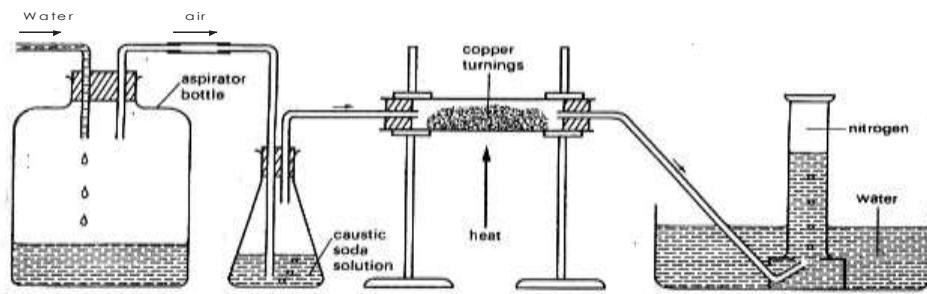
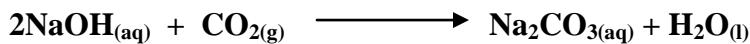
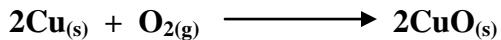


Fig. 10.1 Preparation of nitrogen from air

Nitrogen can be prepared from air by removing the other components of air. Air drawn from the aspirator bottle is passed through sodium hydroxide solution (caustic soda) to remove the carbon(iv) oxide.

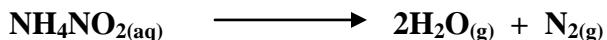


The resultant air free from CO_2 is then passed through heated copper turning to remove oxygen.

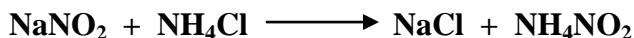


The gas left is Nitrogen and some traces of rare gases. The resultant gas is collected over water.

N₂ is also prepared in the laboratory from thermal decomposition of Ammonium dioxonitrate (iii).



NH₄NO₂ is unstable and its decomposition reaction is exothermic. Heating it is dangerous because the reaction may get out of control and explosion may occur. For this reason, N₂ is prepared instead by heating a mixture of sodium dioxonitrate (iii) and ammonium chloride. The two compounds react on heating to yield ammonium dioxonitrate (iii) which in turn decomposes to produce nitrogen.



NH₄NO₂ is decomposed as soon as it is formed. The N₂ prepared from air is heavier than that prepared from NH₄ NO₂ because it contains rare gases.

Industrial preparation of nitrogen:

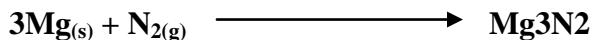
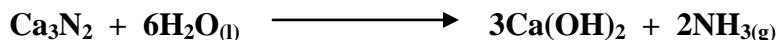
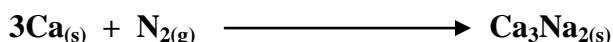
Nitrogen is produced in a commercial quantity by fractional distillation of liquid air. The air, is first passed through caustic soda solution to remove the carbon (iv) oxide. The resultant air which contains mainly O₂ and N₂ and some rare gases are compressed under high pressure and turned into liquid. The liquid is then separated into nitrogen and oxygen by fractional distillation. N₂ gas is evolved first at -196°C and is thus separated from oxygen which boils at -183°C.

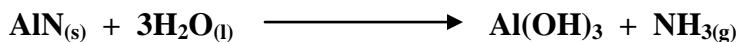
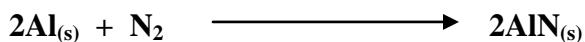
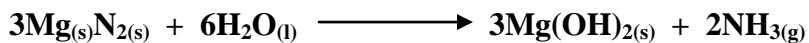
Properties of nitrogen:

It is slightly soluble in water. It is a colourless, odourless and tasteless gas. Impure N₂ is denser than air but pure N₂ is slightly lighter than air.

Nitrogen reacts with some metals, oxygen and hydrogen at a very high temperature.

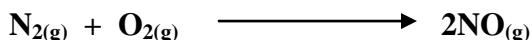
- i. **Reaction with some metals and the formation of metallic hydroxide on hydrolysis with warm water.**



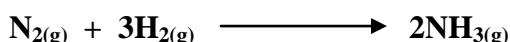


Nitrogen can also react with iron.

- ii. Reaction with oxygen: It reacts with oxygen at about 2000°C to form nitrogen (ii) oxide.



- iii. Reaction with hydrogen: At high temperature and pressure, nitrogen reacts with hydrogen in the presence of a catalyst to form ammonia gas.



Uses of nitrogen:

- i. It is used in the manufacture of ammonia which is an important chemical in the production of fertilisers.
- ii. It is used as a cooling agent or a refrigerant.
- iii. Plants absorb nitrogen in form of soil trioxonitrate (v) for the manufacture of plants protein.
- iv. It is used as a preservative to prevent rancidity (due to the oxidation of fat) in packaged foods.

10.3 Laboratory Preparation of Ammonia

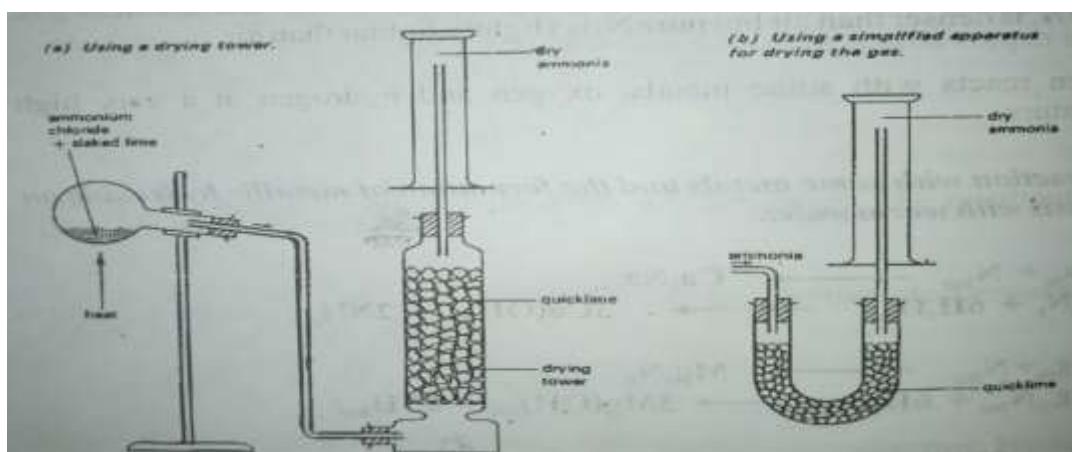
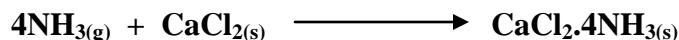
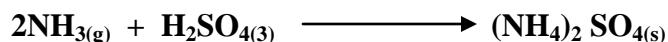
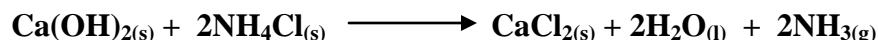


Fig. 10.2: Preparation of ammonia

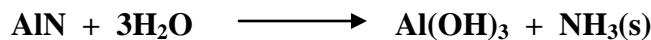
Ammonia is produced in the laboratory by heating ammonium chloride with any alkali. It is also prepared by heating any ammonium salt with any base. Usually in the laboratory, ammonia is prepared by heating NH_4Cl and $\text{Ca}(\text{OH})_2$. $\text{Ca}(\text{OH})_2$ is more preferable to NaOH because the latter is deliquescent. The ground mixture is placed in round bottomed flask connected to a drying tower filled with calcium oxide (quick lime). The gas is not dried by the usual drying agent because concentrated H_2SO_4 and fused CaCl_2 react with ammonia as follows:-



The gas is collected by upward delivery because it is lighter than air.



Ammonia can also be prepared by the reaction of the following nitrides with warm water: Aluminium nitride, Magnesium nitride and Calcium nitride.



Industrial Preparation of Ammonia

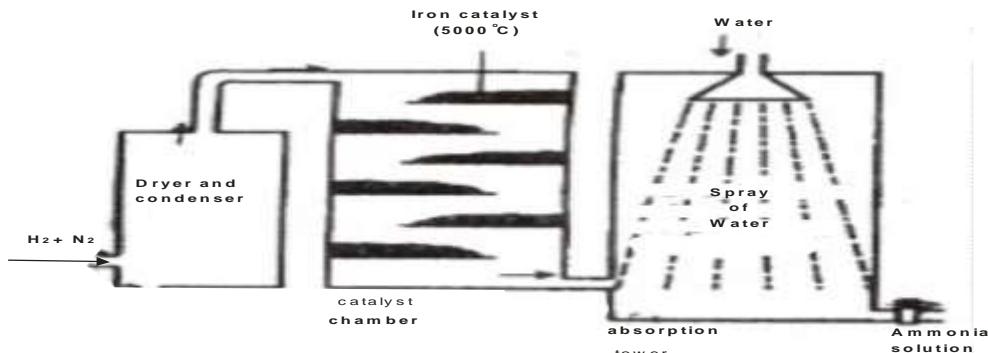
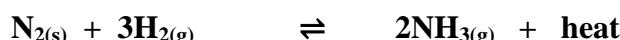


Fig. 10.3

Haber process is used in the production of ammonia in a commercial quantity. Nitrogen and hydrogen is mixed in the volume ratio of 1:3 respectively.



The reaction is reversible and the optimum yield of ammonia is obtained by maintaining these conditions.

- i. High pressure of about 250 atmosphere.
- ii. Low temperature of about 450 to 500°C
- iii. A catalyst (finely divided iron promoted with aluminium oxide).

Aluminium oxide or alumina increases the efficiency of the catalyst. After mixing N₂ and H₂ in the ratio specified above, the mixture is then passed through drier and condenser to dry and compress the gas. The liquified gas passes into the catalyst chamber where the reaction occurs. NH₃ formed dissolves in water in the next chamber and is tapped out as ammonia solution.

Another important commercial source of NH₃ is the destructive distillation of coal (see chapter 22.1 of part I). It is obtained in form of ammoniacal liquor.

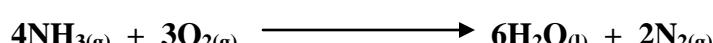
Test for ammonia gas: Ammonia is tested in the school laboratory as follows:

- i. It turns moist red litmus paper blue.
- ii. It has a chocking smell.
- iii. It forms white fumes with concentrated hydrochloric acid

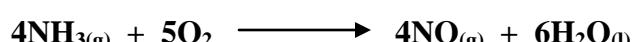


10.4 Physical And Chemical Properties of Ammonia

- 1. Ammonia is a colourless gas with a chocking odour.
- 2. It is very soluble in water. It dissolves in water to form ammonium hydroxide (an alkali).
- 3. It is an alkaline gas. It turns moist red litmus paper blue.
- 4. It burns in oxygen to form vapour and nitrogen.



Though NH₃ react with unlimited supply of air in the presence of hot platinum as a catalyst to form nitrogen (ii) oxide and water.

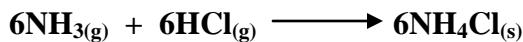
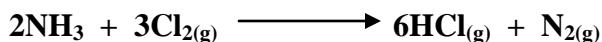


- 5. It reduces copper (ii) oxide to copper while it (NH₃) is oxidized to water and nitrogen.

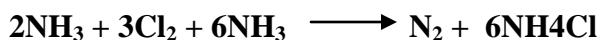


6. **Ammonia reacts with chlorine in two ways:-**

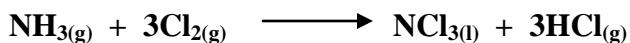
- i. If ammonia is supplied in an unlimited quantity, it reduces chlorine to hydrogen chloride. The $\text{HCl}_{(\text{g})}$ formed reacts with excess NH_3 to produce a white dense fume of ammonium chloride.



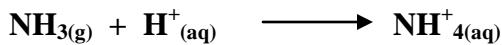
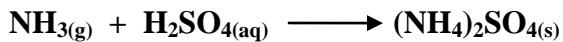
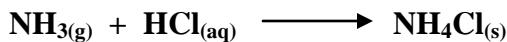
The overall reaction is given thus:-



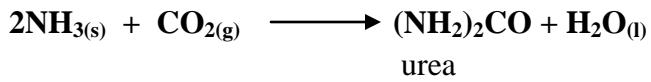
- ii. If chlorine is supplied in an unlimited quantity it oxidizes NH_3 to an oily liquid called nitrogen trichloride. It is a dangerous explosive.



7. Ammonia is a base because it can react with acids to form salts. It can also react with hydrogen proton to form an ammonium ion.

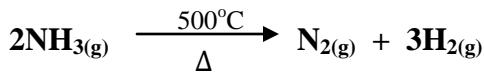


8. Ammonia combines directly with carbon (iv) oxide at a temperature of 150°C and at a pressure of 150 atmosphere to form an important organic compound known as urea.

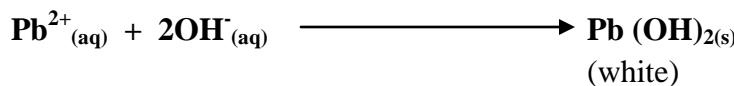
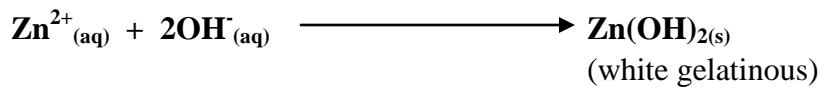
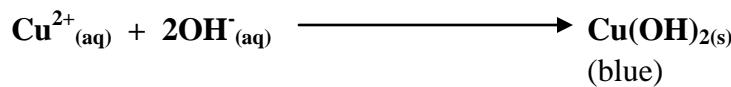
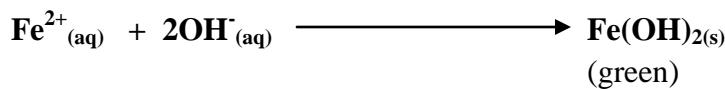


Urea ($\text{NH}_2-\overset{\text{O}}{\underset{\text{||}}{\text{C}}}-\text{NH}_2$) is excreted as a waste product of metabolism in the human body

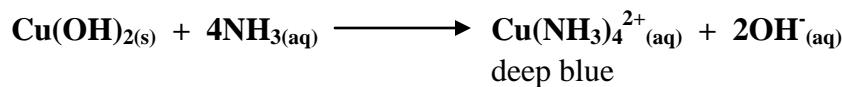
9. Ammonia is thermally decomposed at above 500°C to N_2 and H_2 .



10. Aqueous ammonia is a precipitating agents. It precipitates the insoluble hydroxide of most metals from solution of their salts.



Zn(OH)_2 and Cu(OH)_2 dissolve in excess aqueous ammonia to form colourless and deep blue complex compounds respectively. Formation of these complex compounds is used as a confirmatory test for Zn^{2+} and Cu^{2+} in the school chemistry laboratory.



Uses of ammonia

1. Liquid ammonia is used in some refrigerators as a cooling agent.
2. Aqueous ammonia is used as a precipitating agent in school chemistry laboratory.
3. Aqueous ammonia is used in removing grease and oil stains in laundries.
4. It is used in the making of fertilizers like:-

$(\text{NH}_4)_2\text{SO}_4^-$ Ammonium tetraoxosulphate (vi)

$(\text{NH}_4)_2\text{PO}_4$ Ammonium tetraoxophosphate (v)

NH_4NO_3^- Ammonium trioxonitrate (v)

10.5 Oxides of Nitrogen

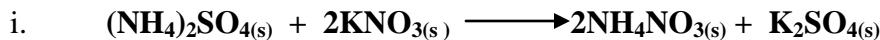
Nitrogen exhibits different oxidation numbers which ranges from -3 to +5. It forms many oxides with oxygen. The oxides of nitrogen are as follows:-

Names of Oxides	Formula	Oxidation Number of nitrogen
Dinitrogen (i) oxide	N_2O	+1
Nitrogen (ii) oxide	NO	+2
Dinitrogen (iii) oxide	N_2O_3	+3
Nitrogen (iv) oxide	NO_2	+4
Dinitrogen (v) oxide	N_2O_5	+5

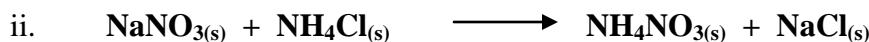
Dinitrogen (i) oxide was discovered by PRIESTLEY. It is called laughing gas because if inhaled, it produces a parculiar intoxication which could result in uncontrollable fits of laughter.

10.6 Laboratory Preparation and Properties of Dinitrogen (I) Oxide

Dinitrogen (i) oxide is prepared by thermal decomposition of ammonium trioxonitrate (v), NH_4NO_3 . The direct heating of this ammonium salt is explosive and dangerous. To avoid this dangers, the gas, N_2O is preferably prepared by heating a mixture of any salt with any ammonium salt which will yield ammonium trioxonitrate (v) by double decomposition reaction



Or



Immediately NH_4NO_3 is formed, it acquires enough heat energy and melts. Effervescence occurs and a colourless gas, N_2O liberated is collected over warm water. The gas is soluble in cold water.



Properties of N_2O

1. It is a colourless gas with pleasant sickly smell.

It is soluble in water and 1.5 times denser than air. It is a neutral oxide.

2. When N_2O is passed over heated iron or copper, it is reduced to nitrogen.





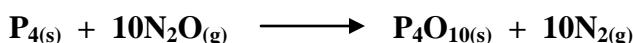
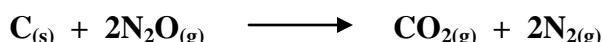
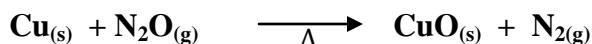
3. N_2O does not burn. At a very high temperature (600°C) it decomposes to nitrogen and oxygen. Oxygen then supports combustion.



4. N_2O rekindles a brightly glowing splinter. This is because the heat produced is so high that it can decompose the gas liberating oxygen which supports the glowing splinter.

N_2O extinguishes a feebly glowing splinter.

N_2O rekindles the burning of these substances; Cu, Mg, S, C and P_4 .



Test for N_2O

It has a faint pleasant, sickly smell and rekindles a bright glowing splinter.

Uses of N_2O

In minor surgical operation like dental surgery, N_2O is used as a mild anesthesia.

10.7 Laboratory preparation of nitrogen (ii) oxide

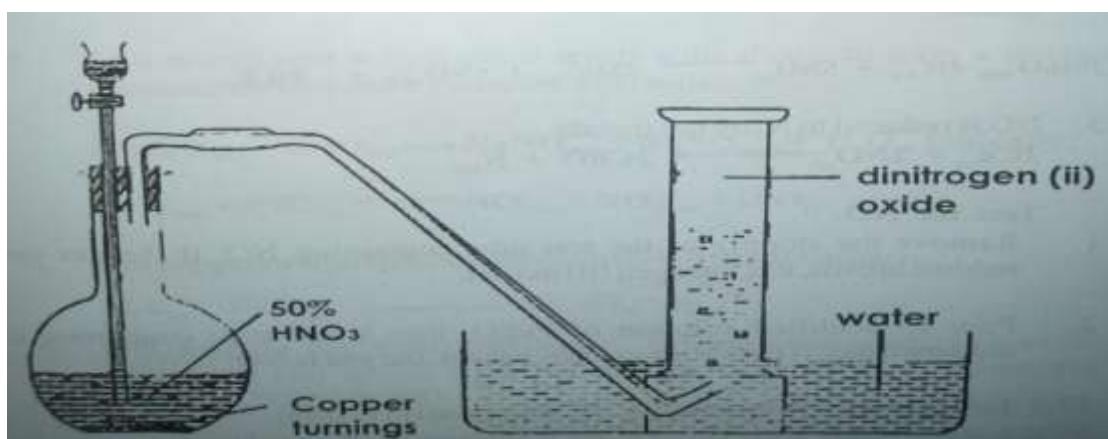


Fig.10.4: Preparation of nitrogen (ii) oxide

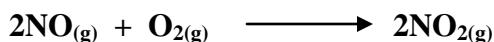
Nitrogen (ii) oxide is produced in the laboratory by the action of trioxonitrate (v) acid on most metals (eg. Cu, Fe²⁺ etc).



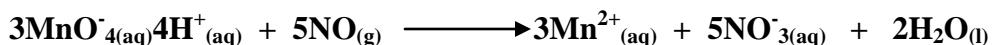
Copper is placed in the flask while conc HNO₃ (equal volumes of conc acid and water) is poured into it through the thistle funnel. When the acid comes in contact with the copper, effervescence occurs, and nitrogen (ii) oxide is produced. Some of this gas is oxidized by the oxygen in the flask to brown fumes of nitrogen (iv) oxide, NO₂. Since NO₂ is very soluble in water it is easily removed when the gas is bubbled over water.

Properties of nitrogen (ii) oxide, NO

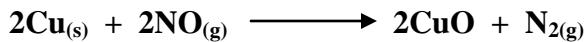
1. It is a colourless and poisonous gas.
2. It is a neutral oxide and almost insoluble in water. It is heavier than air.
3. It combines readily with oxygen to form reddish brown fumes of NO₂



4. It is a weak reducing agent in that it can slowly decolourise acidified KMnO₄.



5. NO is reduced to N₂ by hot metals



Test for NO.

1. Remove the stopper of the test tube containing NO, if the gas turns reddish brown, it is nitrogen (ii) oxide.
2. Pour an acidified solution of FeSO₄ into a test tube containing the unknown gas, if the solution turns brown, the gas is NO.

10.8 Laboratory Preparation of Nitrogen(iv) Oxide:

Nitrogen (iv) oxide is prepared in the laboratory by thermal decomposition of lead (ii) trioxonitrate (v).

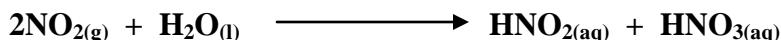


The reddish brown gas, NO₂ is liberated whenever the trioxonitrate (v) of heavy metals (eg. Pb, Cu, Fe etc) are heated. The colourless bench concentrated HNO₃ changes to yellowish colour

after some time. This is due to slight decomposition and subsequent dissolution of NO_2 in the acid.

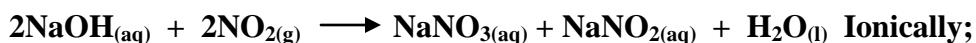
Properties of nitrogen (iv) oxide:

1. Nitrogen (iv) oxide is a reddish - brown gas and it is liquefied at 21°C into a yellow liquid.
2. It has a pungent irritating odour and it is poisonous.
3. The gas is acidic. It turns wet blue litmus paper red.
4. It dissolves in water to form a mixture of dioxonitrate (iii) and trioxonitrate (v) acids. The gas is therefore a mixed acid anhydride

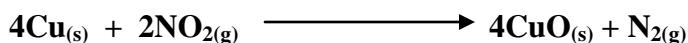


In the presence of excess oxygen, only HNO_3 acid is produced. HNO_2 acid is quickly oxidized to HNO_3 acid.

5. As a mixed acid anhydride, it reacts with alkalis to form a mixture of trioxonitrate (iii) and trioxonitrate (v) salts.



6. Hot copper or iron reduces nitrogen (iv) oxide to nitrogen.



7. NO_2 exist in two molecular forms. At lower temperature, it exists as dinitrogen (iv) oxide, N_2O_4 but at high temperature it dissociates to nitrogen (iv) oxide, NO_2 ,

Uses of NO_2

It is used in the preparation of:

- i. Trioxonitrate (iii) and trioxonitrate(v) salts.
- ii. Trioxonitrate (iii) and trioxonitrate (v) acids.
- iii. Ammonium trioxonitrate (v) used as fertilizer.

10.9 Laboratory Preparation of Trioxonitrate (V) Acid

Trioxonitrate (v) acid is a very important acid used in the laboratory and industries. It is very corrosive to metals and many substances. Because of this, it was previously named aqua fortis' which means strong water.

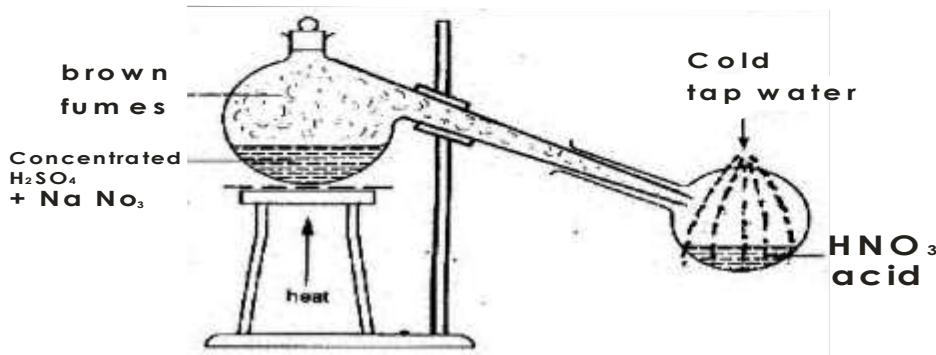
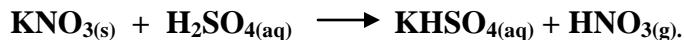


Fig 10.5

Trioxonitrate (v) acid is prepared in the laboratory by addition of concentrated H_2SO_4 acid on potassium or sodium trioxonitrate (v) and then heating the mixture.



On heating, KNO_3 dissolves and effervescence occurs. The trioxonitrate (v) acid formed distills over into the receiving round bottomed flask. Here it is condensed to yellow liquid. The brown fumes observed is due to the decomposition of HNO_3 formed by heat to nitrogen (iv) oxide, NO_2 . NO_2 imparts the brown colour to the flask. When some of the brown fumes of NO_2 dissolves in the acid, HNO_3 , the colour changes to yellow.



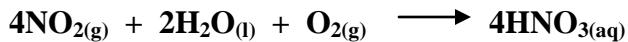
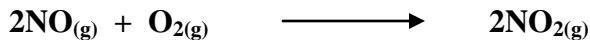
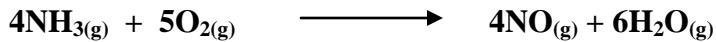
Pure trioxonitrate (v) acid is colourless and the pure form can be obtained by bubbling air (O_2) through the acid solution to remove NO_2 impurity.

In preparation of HNO_3 , the apparatus used must all be glass because the trioxonitrate (v) acid vapour attack rubber and cork rapidly.

Industrial preparation of trioxonitrate (v) acid;

Commercial quantity of trioxonitrate (v) acid is obtained by catalytic oxidation of ammonia. The catalyst used is platinum-rhodium at a temperature of 700°C .

The reactions involved in the industrial process are represented as follows:



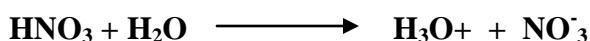
10.10 Properties of Trioxonitrate (V) Acid:

Pure trioxonitrate (v) acid is a colourless, volatile liquid. Its boiling point is 86°C and its specific gravity is 1.5. It is a very good conductor of heat and electricity in dilute form.

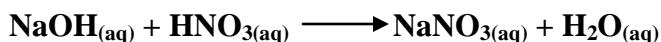
Trioxonitrate (v) acid is monobasic. It behaves as an acid and an oxidising agent.

As an Acid;

- i. Trioxonitrate (v) acid turns moist blue litmus paper red.
- ii. It is a very strong acid in that it ionizes completely in water.



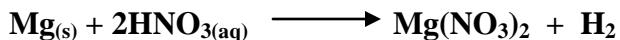
- iii. Trioxonitrate (v) acid is very corrosive to most organic substances eg skin, paper, cloth, wood etc.
- iv. It reacts with bases or alkalis to form salt and water only.



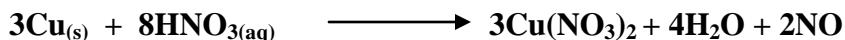
- v. It reacts with trioxocarbonate (iv) to liberate carbon (iv) oxide.



- vi. Unlike other acids, trioxonitrate (v) acid rarely liberates hydrogen with metal except in very dilute condition (about 1%) can react with Ca, Mg and Mn to produce hydrogen.



The reason is that the hydrogen formed is immediately oxidized to water. If the acid is slightly concentrated by diluting it with its own volume of water NO is produced.



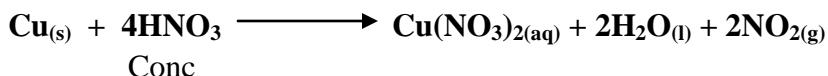
Aluminium and Iron do not react with the concentrated acid probably because of the formation of surface coating of oxides which is passive and stops any further action by the acid. Therefore concentrated HNO_3 is usually stored or transported in containers lined with aluminium or iron. The chemical action of trioxonitrate (v) acid is complicated by the fact that it is both a strong acid and a powerful oxidizing agent. In dilute form, it is almost completely ionized and thus

strongly acidic. On the other hand, pure and concentrated HNO_3 is poorly ionized. Therefore as its concentration increases, its strong oxidizing power tends to predominate.

As an oxidizing agent:

HNO_3 acid is an oxidizing agent because it readily accepts electrons. It can be reduced to a wide variety of reduction products e.g. NO , NO_2 , N_2O , N_2 depending on the concentration of the acid, strength of the reducing agent and temperature.

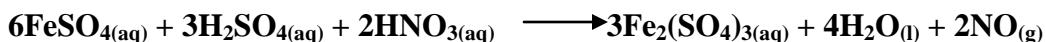
- i. It reacts with Cu, oxidizing it to Cu^{2+} while itself is reduced to NO_2 .



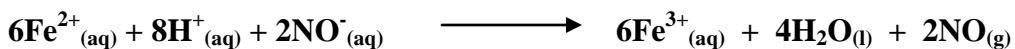
3. $\text{Cu}_{(s)} + 8\text{HNO}_{3 \text{ (50% conc.)}} \longrightarrow 3\text{Cu}(\text{NO}_3)_{2 \text{ (aq)}} + 4\text{H}_2\text{O}_{(l)} + 2\text{NO}_{(g)}$

Mineral acids e.g. H_2SO_4 , HCl cannot react with copper because it is below hydrogen in the electrochemical series.

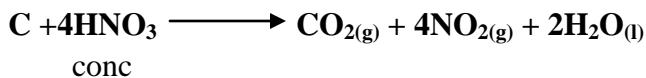
- ii. It oxidizes Fe^{2+} to Fe^{3+} in the presence of H_2SO_4 acid



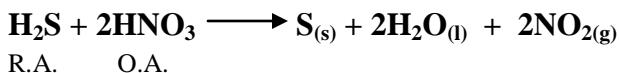
Ionically:



- iii. Carbon is oxidized to carbon (iv) oxide when heated with trioxonitrate (v) acid.



- iv. Hydrogen sulphide which is a reducing agent reacts with concentrated trioxonitrate (v) acid and is oxidized to sulphur. HNO_3 is reduced to NO_2 .



- v. Concentrated HNO_3 oxidizes sulphur to tetraoxosulphate (vi) acid.



Uses of trioxonitrate (v) acid:

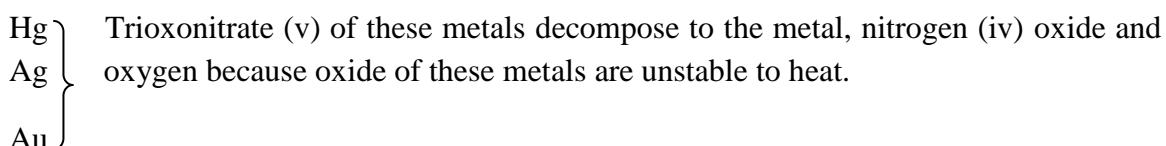
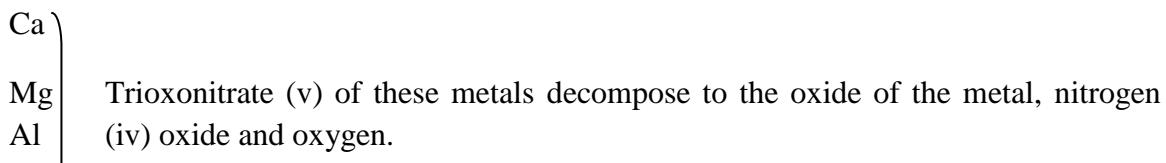
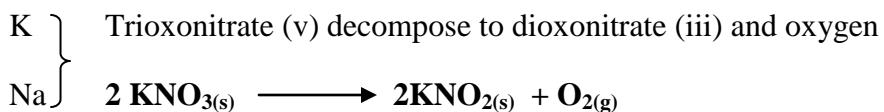
1. It is used as an oxidizing agent in the production of polymeric materials like Nylon and terylene.

2. It is used as a nitrating and oxidizing agents.
3. It is used in the making of fertilizers.
4. It is used in the manufacture of explosive eg methyl 2, 4, 6 - trinitrobenzen.
5. A mixture of $\text{HCl}_{(\text{aq})}$ and $\text{HNO}_3_{(\text{aq})}$ (Aqua regia) in the ratio of 3 : 1 respectively is used as a solvent for sturbon stain removal. It is also a solvent for gold and platinium.

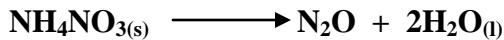
10.11 The Trioxonitrate (V) Salts

These salts should be studied under solubility, Action of heat and their uses.

When a sample of trioxonitrate (v) is added to water, it will dissolve. All trioxonitrate (v) are soluble in water.



Ammonium trioxonitrate (v) decompose to dinitrogen (i) oxide and water.

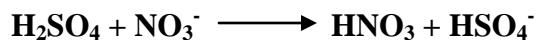


Test for trioxonitrate (v) ion, NO_3^-

i. **The brown ring test:**

To the solution of the sample of an unknown salt, add freshly prepared FeSO_4 solution. Then hold the test tube in slanting position and add through the side of the test tube a conenous stream of conc. H_2SO_4 . The acid will settle at the bottom of the test tube. At junction where the solution met, brown ring occur if the sample contains trioxonitrate (v) ion, NO_3^-

2. The mixture of the unknown solid and copper with some concentrated H_2SO_4 is warmed. The formation of reddish brown fumes of NO_2 will confirm the presence of NO_3^- ion.



Uses of trioxonitrate (v) salts

KNO_3 is used in the making of gun powder.

NaNO_3 is used as a nitrogenous fertilizer. NaNO_3 is used in the making of HNO_3 acid, KNO_3 and NaNO_2 . NaNO_3 cannot be used in the manufacture of gun powder because of its deliquescent nature.

A mixture of $\text{Ca}(\text{NO}_3)_2$ and lime, $\text{Ca}(\text{OH})_2$ is used as fertilizer.

10.12 Nitrogen Cycle: Nitrogen cycle is a continuous process by which the atmospheric nitrogen enters into the soil as the soil Nitrates, followed by continuous decomposition of these nitrates to release nitrogen back into the atmosphere.

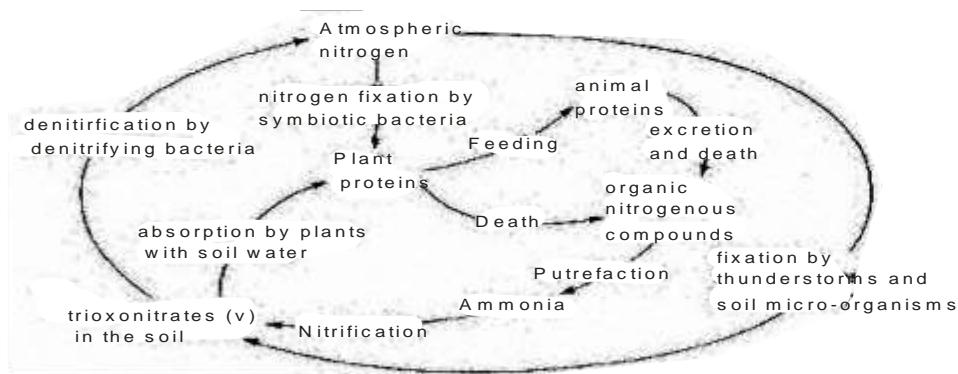
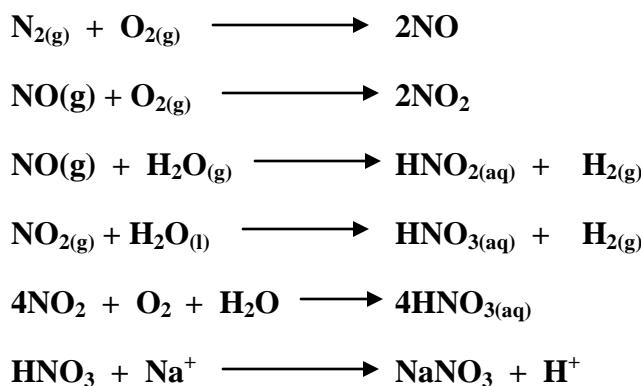


Fig. 10.6: Nitrogen cycle

Living cells are made of protoplasm which is merely protein. Unlike carbohydrate and fat, protein is made up of an element - nitrogen. Nitrogen is therefore the major constituent of

protein. Animals cannot make their proteins but plants can synthesis their own proteins. Animals therefore get their proteins by eating other animals and plants protein. Plants cannot make use of the free atmospheric nitrogen for their synthesis. They use ammonium salt or trioxonitrate (V) which they absorb from the soil. Atmospheric nitrogen is being introduced into the soil through the following processes:

Nitrogen fixation: During lightening, some atmospheric nitrogen is oxidized to nitrogen (ii) oxide, NO. Some of NO is further oxidized to nitrogen (iv) oxide, NO₂. These two oxides of nitrogen dissolve in rain water to form dioxonitrate (iii) and trioxonitrate (v) acids respectively. When these acid enter the soil, they combine with the cations of the mineral salts in the soil to form trioxonitrate (v) salts.



Nitrogen fixing bacteria: The soil micro-organism such as the symbiotic bacteria - Rhizobium living in the root nodules of leguminous plants help to replenish soil trioxonitrate (v). They convert the atmospheric nitrogen into organic nitrogenous compound used by the plant. The plants protein or the organic nitrogenous compound are re-introduced into the soil when the plants die. Azotobacter is another bacteria (living in the soil) that converts atmospheric nitrogen to trioxonitrate (v).

Harber process: During this process atmospheric nitrogen is converted to ammonia which is chiefly used in the making of fertilizers. The fertilizers are applied to the soil and most of them are absorbed by plants.

Decay of organic matter: The nitrogen is taken up by plants in form of nitrate or trioxonitrate (v). The trioxonitrate (v) or nitrates are taken up by animals by eating the plants. The trioxonitrate (v) is returned to the soil during excretion and when the organisms (plants or animals) die and decay. Putrefying bacteria and fungi convert the organic nitrogenous compound in the dead organisms to ammonia and free nitrogen. Some of the ammonia are converted to trioxonitrate (v) by nitrifying bacteria in the soil. The action of denitrifying bacteria covers the nitrogenous compounds in the soil to free nitrogen.

Exercises:

1. What are the components of air?
2. Briefly describe how nitrogen can be obtained in an industrial quantity.
3. Use chemical equation only to show how nitrogen is prepared in the school chemistry laboratory.
4. Write two physical properties of nitrogen.
5. Write two chemical properties of nitrogen.
6. What are the uses of nitrogen?
7. List two oxides of (i) Nitrogen (ii) Phosphorus
8. Briefly describe a laboratory method of preparing ammonia.
9. Write two chemical properties of ammonia.
10. Write three importance of ammonia.
11. How can you test ammonia in the school chemistry laboratory?
12. Name five oxides of nitrogen.
13. What differentiates N_2O from O_2 since both can rekindle the burning of a bright glowing splint?
14. Use an equation only to show a laboratory preparation of nitrogen (ii) oxide.
15. NO_2 gas is a mixed acid anhydride. Explain this statement and support it with a balanced equation.
16. Use equations only to show:
 - (a) Laboratory preparation of trioxonitrate (v) acid.
 - (b) Industrial preparation of trioxonitrate (v) acid.
17. Use two chemical equations each to show that trioxonitrate (v) acid behaves as:
 - (a) an acid (b) an oxidizing agent.
18. Write two uses of trioxonitrate (v) salts.
19. What are the chemical tests for trioxonitrate (v) ion?

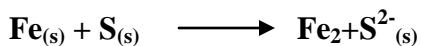
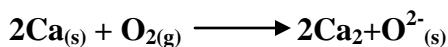
20. Define ‘Nitrogen cycle’

CHAPTER 11

SULPHUR

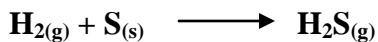
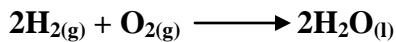
11.1 General Properties of Group (VIA) Elements

Elements in group 6 or (VI) include oxygen and sulphur. They are non-metals and are electron acceptors. They are therefore oxidizing in nature. For example;



The two elements do not react with water. Unlike sulphur, oxygen is slightly soluble in water.

Oxygen and sulphur combine with hydrogen to form water and hydrogen sulphide respectively.



11.2 Electronic Structure of Sulphur

The atomic number of sulphur is 16. Its atomic structure can be represented thus;

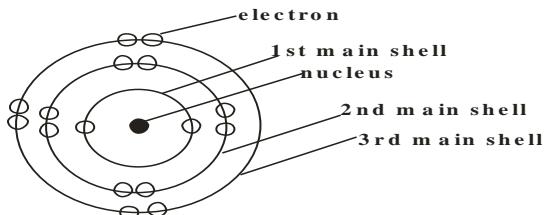


Fig. 11.1 Atomic structure of sulphur

It has six electrons in its outermost main shell and it is therefore classified as group 6 element. Its electronic configuration is; $\text{1S}^2, \text{2S}^2, \text{2P}^6, \text{3S}^2, \text{3P}^4$.

11.3 Allotropes; Preparation, Properties and Uses:

Compounds of Sulphur

Sulphur is a non-metallic element. It is in the group 6 and period 3 in the periodic table. It is a yellow solid and highly medicinal.

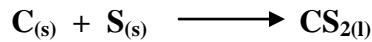
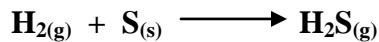
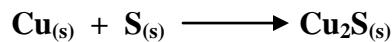
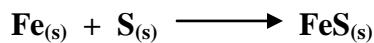
Effects of Heat on Sulphur:

If some powdered sulphur is put in a narrow test tube and warmed gently, shaking well and rotating the test tube to avoid over heating, sulphur undergoes the following changes as the temperature rises.

1. The solid yellow sulphur changes to orange brown. It melts at about 119°C to an ambere-coloured mobile liquid.
2. It becomes much darker in colour as the temperature increases and at about 180°C it suddenly becomes very viscous. Its viscosity is so high that at about 200°C , the test-tube can be inverted without losing its content (sulphur).
3. As the temperature is increased further, the molten sulphur become mobile again and dark reddish in colour.
4. At a temperature of 444°C , the sulphur boils to give a reddish - brown vapour. As the sulphur cools, the above changes occur in a reverse order. These changes are possible when air is excluded.

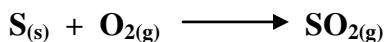
11.4 Physical And Chemical Properties of Sulphur:

1. Sulphur is a yellow solid. It is insoluble in water but dissolves in carbon disulphide. It is a poor conductor of heat and electricity. It melts when heated.
2. Sulphur combines directly with nearly all metals and most non-metals to form sulphides. Sulphide is a binary compound of Sulphur.



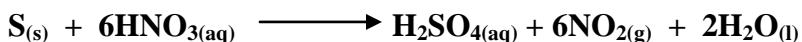
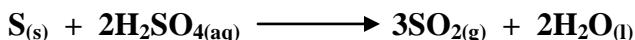
Very reactive metals like sodium and potassium can react spontaneously with sulphur without heating if the two substances (elements) are ground to fine powder.

3. Sulphur reacts with or burns in an excess supply of air to form sulphur (iv) oxide, SO_2 , and a small traces of sulphur (vi) oxide, SO_3



The gas can be tested using wet blue litmus paper, acidified KMnO_4 and $\text{K}_2\text{Cr}_2\text{O}_7$ solution.

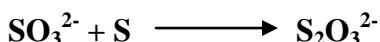
4. Dilute acids, concentrated or dilute hydrochloric acid have no effect on sulphur. Oxidizing agents such as hot concentrated H_2SO_4 and HNO_3 oxidize sulphur.



5. Sulphur reacts, with hot alkalis to form a mixture of sulphides and trioxosulphate (iv)



If sulphur is in an excess supply, a mixture of polysulphide and a trioxothiosulphate (vi) is formed.



11.5 Allotropes of Sulphur

Like phosphorus, carbon and oxygen, sulphur can exist in many allotropic forms. This means that sulphur can exist in many forms in the same physical state. The four allotropic forms of sulphur are: Rhombic, Monoclinic, Plastic and amorphous sulphur. They are all solid.

The Rhombic Sulphur or Alphae Sulphur (μ - Sulphur): Rhombic sulphur is yellow in colour. It is made up of S_8 molecules. It has a crystalline structure and it is octahedral in shape. Sulphur exists naturally as Rhombic sulphur and it is the only allotrope which is stable at a temperature below 96°C .

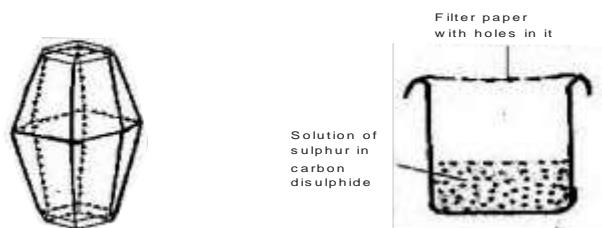


Fig. 11.2

Preparation: Solution of sulphur in carbon (iv) sulphide, CS_2 is made by adding some powdered sulphur in CS_2 . It is then shaken to dissolve. It is done at a temperature below 96°C . Filter the solution and cover the filtered solution with a filter paper (which is perforated with a needle to make many holes on it). Leave the beaker in a fume cupboard so that the solvent will evaporate. Rhombic sulphur crystals will then be formed.

NOTE: This preparation is carried out in a fume cupboard because CS_2 is poisonous and highly inflammable.

Monoclinic sulphur or prismatic sulphur (b sulphur): The crystals of monoclinic sulphur are thin, long and needle shaped. It consists of S_8 molecules. It is the only stable allotropic form at temperature between 96°C and 119°C

Monoclinic sulphur is obtained by cooling molten sulphur. Some powdered sulphur is heated in a large crucible till it melts. More sulphur is added and heating is continued till the crucible is almost filled with molten sulphur. It is then removed from the flame and allowed to cool. A hard crust is formed at the top. Pierce two holes through the crust and pour off the remaining molten sulphur. Needle shaped crystals of monoclinic sulphur is then found deposited on the sides of the crucible.

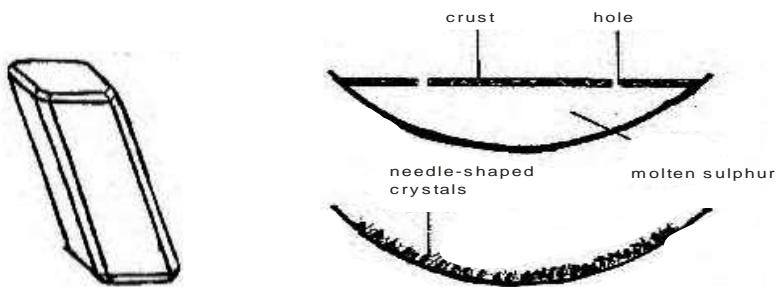
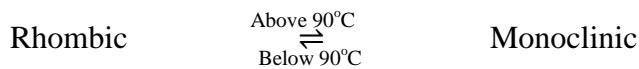


Fig. 11.3

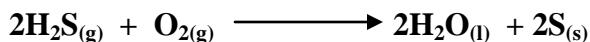
The existence of sulphur as either Rhombic or monoclinic allotropes is dependent on temperature. Rhombic changes to monoclinic above 96°C and vice versa. 96°C is therefore called the transition temperature of the two allotropes.



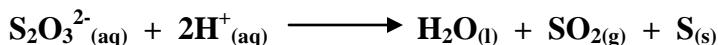
	Rhombic	Monoclinic
Melting point	114°C	119°C
Colour	Bright yellow	Amber
Density	2.08gcm ⁻³	1.98g cm ⁻³
Shape	Octahedral	Needle shape
Transmission of light.	Translucent	Transparent

The two main crystalline allotropic forms of sulphur are rhombic and monoclinic. The two non-crystalline forms are plastic and amorphous sulphur. Plastic sulphur is a super cooled form of sulphur. It is not a true allotrope because it is unstable and can change back to rhombic sulphur on standing.

Amorphous sulphur (pale yellow) is formed by exposing an aqueous saturated solution of hydrogen sulphide to air.



It can also be formed by the reaction between dilute hydrochloric acid and trioxothiosulphate (vi) soluiton.



11.6 Laboratory Preparation of Hydrogen Sulphide

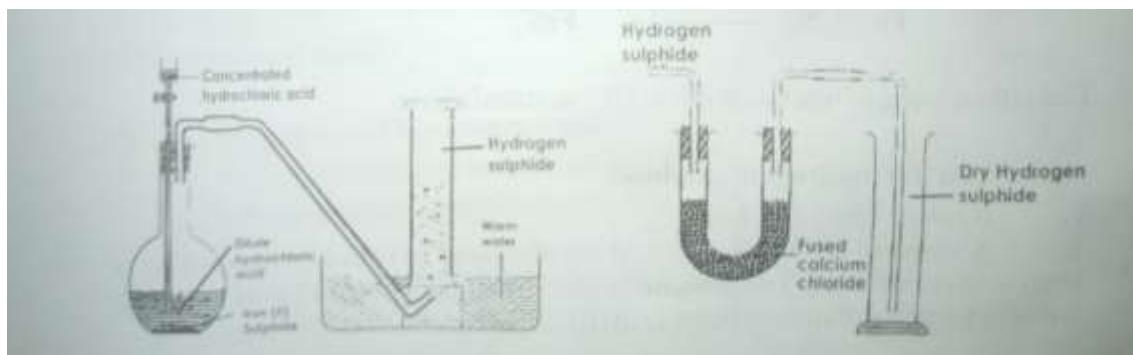
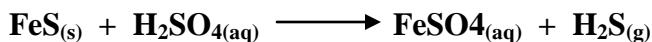
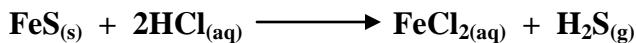


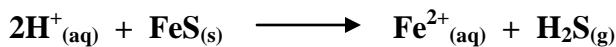
Fig. 11.4(a) Preparation of hydrogen sulphide

Fig. 11.4(b): Collection of dry hydrogen sulphide

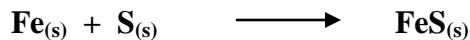
Hydrogen sulphide is prepared in the laboratory (and in a commercial quantity) by the action of dilute hydrochloric acid on iron (ii) sulphide. Some iron (ii) sulphide is put into the flat bottomed flask. Dilute hydrochloric acid or tetraoxosulphate (vi) acid is poured into the flask through the dropping funnel. As the acid comes in contact with the metallic sulphide, there is effervescence and hydrogen sulphide is produced. It is collected over warm water (because hydrogen sulphide is soluble in cold water).



Ionically,



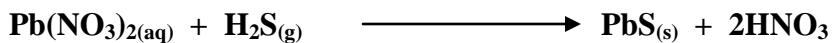
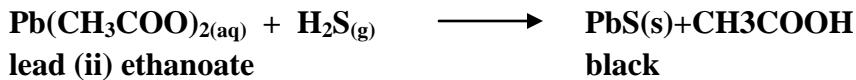
H_2S gas can easily be liquefied and stored for sale in a steel cylinder. If the gas is required dry, it is passed through fused calcium chloride, CaCl_2 . It is then collected by downward delivery in fume chamber as the gas is denser than air. Concentrated tetraoxosulphate (vi) acid cannot act as a drying agent for H_2S gas because it oxidizes the gas to sulphur. If the gas is to be prepared starting from sulphur, then iron (ii) sulphide is first of all produced by heating a mixture of iron filings and sulphur in the ratio of 56 : 32 by mass. The heating is continued till a black or dark grey residue of FeS is left.



FeS is then made to react with dilute HCl as stated above.

Test for hydrogen sulphide:

1. It smells like a rotten egg.
2. A stripe of filter paper soaked in either lead (ii) ethanoate solution or lead (ii) trioxonitrate (v) solution turns black if dropped into a jar of H_2S . This is due to the precipitation of black lead (ii) sulphide.



Kipp's apparatus: This is a device for obtaining intermittent supplies of a frequently used gas such as hydrogen, carbon (iv) oxide or hydrogen sulphide.

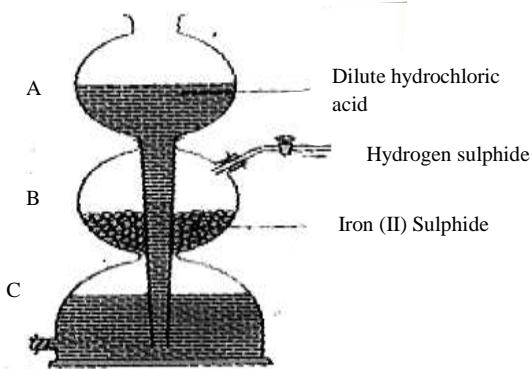
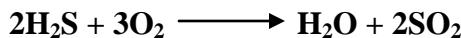


Fig. 11.5: Kipp's apparatus

The apparatus can be used in the production of H_2S . When the tap T is opened, the pressure in B decreases. The acid from C then rises into B and attacks the FeS inside it producing H_2S . The gas is delivered through T. When T is turned off, the gas, H_2S which is still inside B increases or raises the pressure in B. The increased pressure in B forces the acid back into C. This prevents the acid from reacting once more with the sulphide. The generation of H_2S is then stopped till tap T is opened again.

11.7 Properties of Hydrogen Sulphide

1. Hydrogen sulphide is a colourless gas that smells like a rotten egg.
 2. It burns in excess supply of air to form sulphur (iv) oxide.



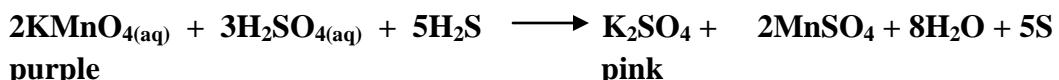
Here H_2S is oxidized to SO_2 .

In a limited supply of air, H₂S is oxidized to sulphur.

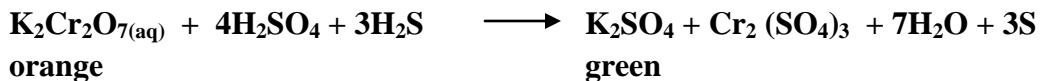
3. H_2S is a powerful reducing agent. It acts as electron donor.



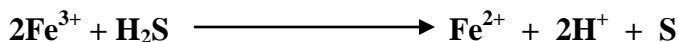
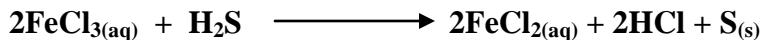
- a. H_2S reacts with acidified potassium tetraoxomaganate (vii) and decolourizes the purple colour of KMnO_4 to pale pink with deposit of sulphur.



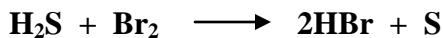
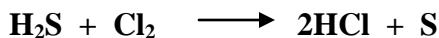
- b. H₂S reacts with acidified potassium heptaoxodichromate (vi), K₂Cr₂O₇ and changes the orange colour of K₂Cr₂O₇ to green with deposit of sulphur.



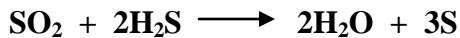
- c. H_2S reacts with brownish yellow iron (iii) chloride and reduces it to a green solution of iron (ii) chloride with deposit of sulphur.



- d. H₂S reacts with halogens in the presence of moisture. H₂S is then oxidized to S.



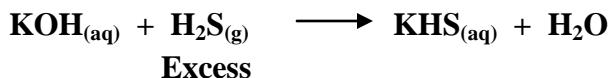
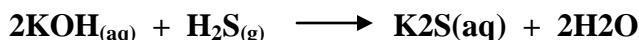
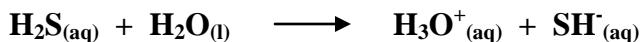
- e. H₂S acts as a stronger reducing agent when it reacts with sulphur (iv) oxide. In the presence of moisture it reduces SO₂ to S.



- f. Oxidizing acids react with and oxidize H_2S to sulphur.



H_2S dissolves in water to form dibasic acid. It therefore forms both normal and acid salts when it reacts with an alkali.

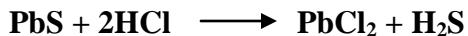


Almost all sulphides except Na^+ , K^+ and NH_4^+ sulphides are insoluble. Pb, Cu, Ag and Fe form black sulphides. ZnS is white.

Most sulphides when heated in air form SO_2 and the metallic oxide.

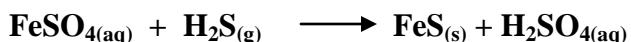
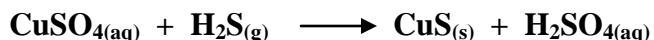


Mercury, Hg yields the metal and most sulphides react with hydrochloric acid on heating to yield H_2S .



Uses of hydrogen sulphide

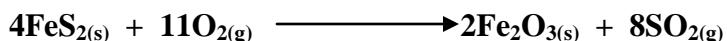
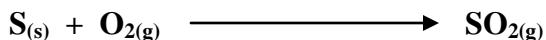
- It is used in qualitative analysis as a precipitating agent. It precipitates the insoluble coloured sulphides from their ores.



- Sodium sulphide is used for the preparation of ‘fast’ dyes, ie dyes which do not fade.

11.8 Oxides of Sulphur: Trioxosulphate (IV) Acid and Trioxosulphate (IV) Salts.

Preparation of Sulphur (iv) Oxide: Sulphur (iv) oxide is prepared in an industrial quantity by burning sulphur or metallic sulphide in air or oxygen.



In the laboratory, it is readily given off when a mixture of sodium or potassium trioxosulphate (iv) and conc. tetraoxosulphate (vi) or hydrochloric acid is heated. The trioxosulphate (iv) acid formed decompose immediately to yield water and sulphur (iv) oxide.

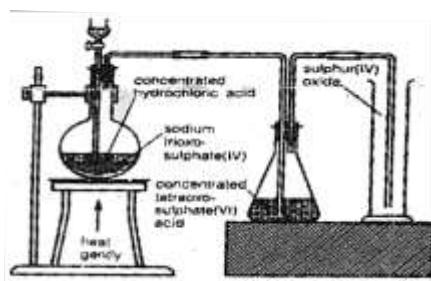
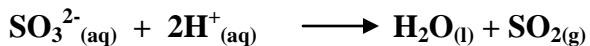


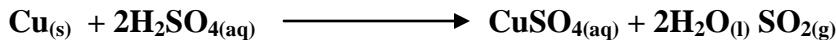
Fig. 11.6: Preparation of Sulphur (iv) Oxide



Ionomically,



The gas can also be generated by heating conc tetraoxosulphate (vi) acid with copper.

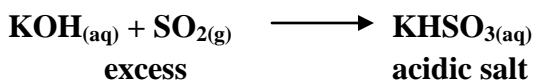
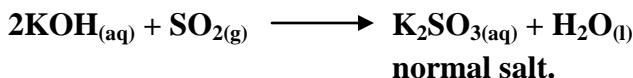


Dry the gas evolved by passing it through concentrated H_2SO_4 acid and collect it by downward delivery. SO_2 is denser than air. The gas jar is covered with moist potassium dichromate (vi) paper which turns from orange to green colour when the gas jar is full.

NOTE: This experiment is carried out in a fume cardboard.

Properties of sulphur (iv) oxide:

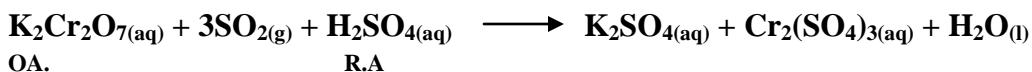
1. It is heavier than air. It is colourless and has a very irritating smell.
2. It is a poisonous gas and dissolves in water to form trioxosulphate (iv) acid.
3. As an acidic gas, it reacts with alkali to form salt.



4. SO_2 is a strong reducing agent. It is an electron donour.



- a. SO_2 reacts with potassium heptaoxo- dichromate (vi) solution and turns its colour from orange to green. Chromium (vi) is reduced to chromium (iii).



- b. SO_2 reduces brown FeCl_3 solution to green FeCl_2 solution.

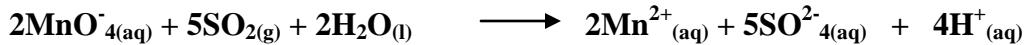


SO_2 is then oxidized to H_2SO_4

- c. SO_2 decolourizes an acidified potassium tetraoxomanganate (vii) solution.

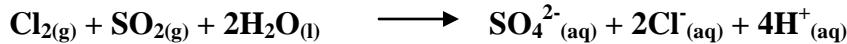


Ionically

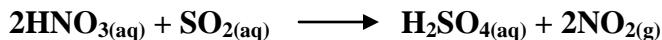


SO_2 reduced Mn (vii) to Mn (ii) while it is oxidized to H_2SO_4 acid.

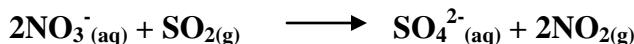
- d. SO_2 reduces the coloured solutions of halogens (Br, Cl, I) to their colourless solutions of hydrogen halides.



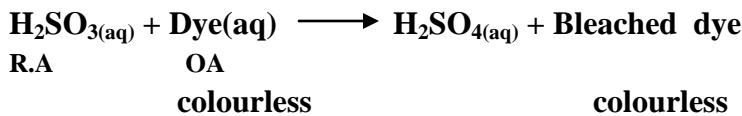
- e. It reduces concentrated trioxonitrate (v) acid to reddish brown nitrogen (iv) oxide.



Ionomically

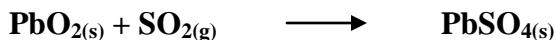


5. SO_2 acts as a bleaching agent because of its reducing powers. It can bleach both natural and artificial dyes.



Simple comparison of the bleaching action of SO_2 and Cl_2 shows that as Cl_2 bleaches by oxidizing the dye, SO_2 bleaches by reducing the dye. Thus the dye bleached by SO_2 cannot last as long as the one bleached by Cl_2 because the bleached dye may be reoxidized by atmospheric oxygen to form the original coloured compound. This is the reason why white straw hat turns yellow after some time.

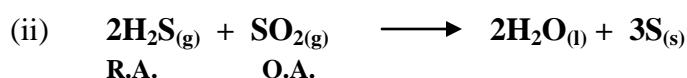
6. SO_2 reacts with heated lead (iv) oxide and reduces it to lead (ii) tetraoxosulphate (vi).



7. SO_2 reacts with O_2 in the presence of a catalyst to form SO_3



8. In the presence of a stronger reducing agent, SO_2 can act as an oxidizing agent.



The (ii) reaction occur in the presence of moisture.

Test for SO_2 :

- (1) It has a very irritating smell
- (2) It bleaches moist coloured flowers.
- (3) It changes the colour of both acidified potassium tetraoxomangate (vii) and potassium heptaoxodichromate (vi) (see the reaction above)

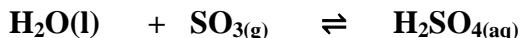
SO_2 is used in the manufacture of the following: H_2SO_3 acid, H_2SO_4 acid, germicide, bleaching agent and other important chemicals used in the industry.

Sulphur (vi) oxide, SO_3 :

It is prepared by combining SO_2 and O_2 in the presence of any of these catalysts: Platinized asbestos or vanadium (v) oxide. A slight pressure and a temperature of $400 - 450^\circ\text{C}$ are other conditions required.

Properties of SO_3 :

It has a low boiling point, 45°C . It is an acidic oxide. It reacts with water to form tetraoxosulphate (vi) acid. It is therefore an acid anhydride of H_2SO_4 acid.



It dissolves in concentrated tetraoxosulphate (vi) acid to form oleum.



11.9 Trioxosulphate (IV) Acid. It is an unstable acid formed when SO_2 reacts with water. It decomposes with ease at room temperature.



H_2SO_3 acid is prepared using the apparatus shown below.

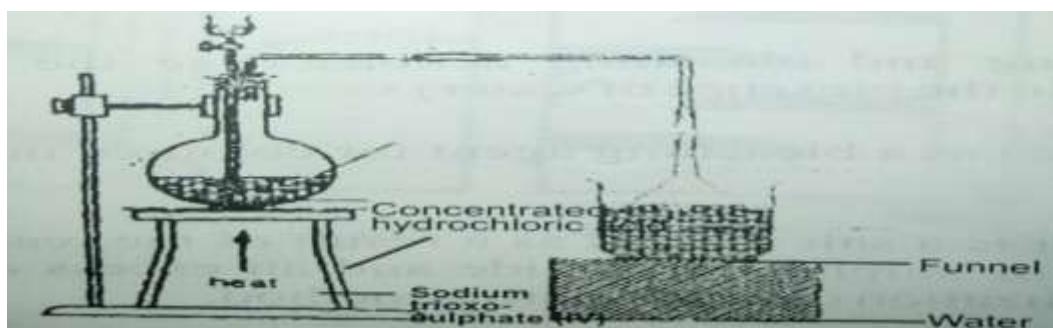
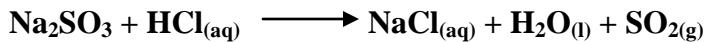


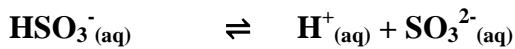
Fig. 11.7

A mixture of HCl acid and Na₂SO₃ is heated in the round bottomed flask. The gas, SO₂ evolved dissolves in water to form H₂SO₃ acid.

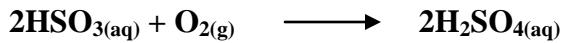


Properties and uses of H₂SO₃

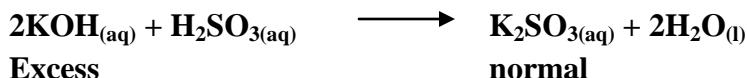
- It is a dibasic acid. It ionizes thus:



- It turns blue litmus paper red.
- H₂SO₃ combines with atmospheric oxygen to form H₂SO₄ acid.



- As a dibasic acid, it forms both normal and acid salts.



- It has both reducing and bleaching actions (See the reducing and bleaching actions of SO₂ in the presence of water).

Trioxosulphate (iv) is used as a bleaching agent for coloured materials and in the making of germicide.

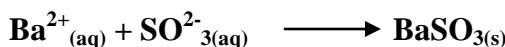
Trioxosulphate (iv) salt is a salt formed as a result of the reaction between trioxosulphate (iv) acid or sulphur (iv) oxide and an excess of alkali. It is prepared by either neutralization or precipitation method.

Test for trioxosulphate (iv)

- If it is warmed with a dilute acids SO₂ gas is liberated.



2. If some barium trioxonitrate (v) solution is added to the unknown salt, a white precipitate of barium trioxosulphate (iv) will be formed in the presence of trioxosulphate (iv). If dilute trioxonitrate (v) is then added to the white precipitate, it dissolves.



11.10 Tetraoxosulphate (VI) Acid and its Salts

Tetraoxosulphate (vi) acid is manufactured in a commercial quantity by a more recent and popular method known as the contact process. Lead chamber process which is an older method is no longer very popular.

Industrial Preparation of Tetraoxosulphate (VI) Acid.

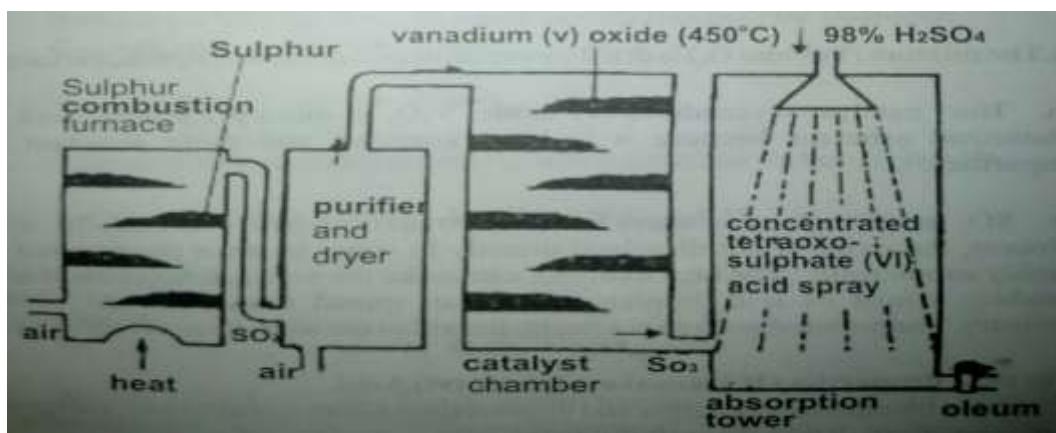
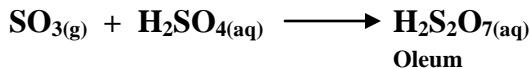
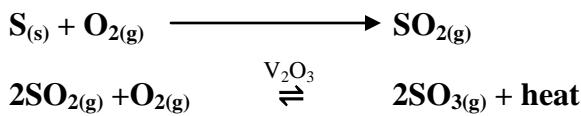
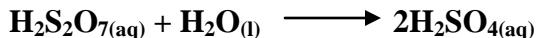


Fig. 11.8: The Contact Process.

In the contact process, sulphur is heated with the dry air in the sulphur combustion furnace. The sulphur (iv) oxide formed mixes with the air (O_2) and enters into an electric chamber where the mixture is purified and dried. The mixture (SO_2 and O_2) then enters into a catalytic chamber where the two gasses combine catalytically to form sulphur (vi) oxide, SO_3 . The SO_3 gas is cooled and made to enter into the absorption tower. In this tower, SO_3 dissolves in concentrated tetraoxosulphate (vi) acid to form a very thick liquid or a fuming liquid called Oleum. The oleum is then diluted with a calculated amount of water to produce 98% tetraoxosulphate (vi) acid.

The reactions that occur in the contact process are as follows:





NOTE:

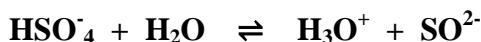
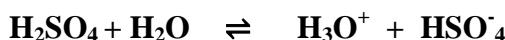
- i. In the electric chamber, impurities like dust particles and arsenic compounds are removed from the mixture of gases. These impurities can make the catalyst inactive.
- ii. The mixture (SO_2 and O_2) is dried by concentrated tetraoxosulphate (vi) acid.
- iii. The catalyst - vanadium (v) oxide, V_2O_5 is more preferably used to platinized asbestos because it is less expensive and more resistant to impurities.
- iv. SO_3 gas dissolves in water to form tetraoxosulphate (vi) acid. In this process, the oxide is not dissolved directly in water because the reaction is highly exothermic. The heat evolved, can make the acid solution to boil and produce a mist of acid droplets which can spread throughout the whole industry. The polluted air then is a health hazard to the workers in the industry.

Physical Properties of Tetraoxosulphate (vi) Acid.

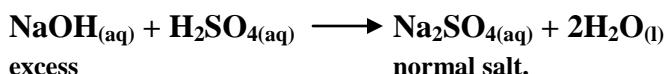
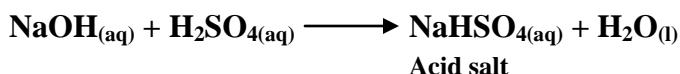
1. It is a colourless, odourless and viscous liquid with a density of 1.84gcm^{-3} . In dilute form, it turns blue litmus paper red.
2. It is hydroscopic, and it is therefore used as a drying agent for many gases.
3. It is corrosive and causes severe burns to human skin.

Chemical Properties of Tetraoxosulphate (vi) Acid

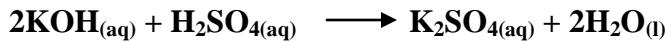
1. Dilute tetraoxosulphate (vi) is a strong acid. It is dibasic and ionizes in two stages as follows.



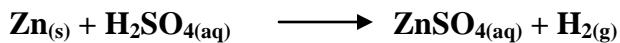
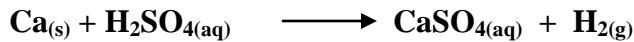
- a. H_2SO_4 acid can therefore form both acidic and normal salts.



- b. Dilute H_2SO_4 acid neutralises bases or alkalis to form salt and water only.



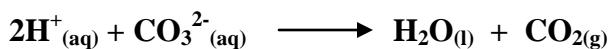
- c. It reacts with the more electropositive metals to liberate hydrogen.



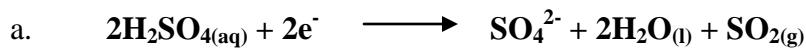
- d. It reacts with traoxocarbonate (iv) salts to generate carbon (iv) oxide.



Ionically,



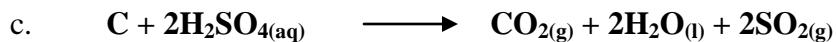
2. H_2SO_4 acid is an oxidizing agent because it can accept electrons from a reducing agent and it is reduced to sulphur (iv) oxide.



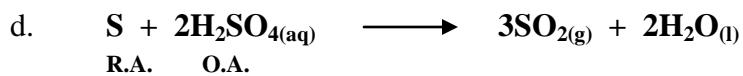
It oxidizes many metals and non-metals:



Ionically,

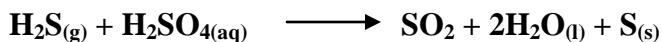


Carbon is oxidized to carbon (iv) oxide while H_2SO_4 is reduced to SO_2 .

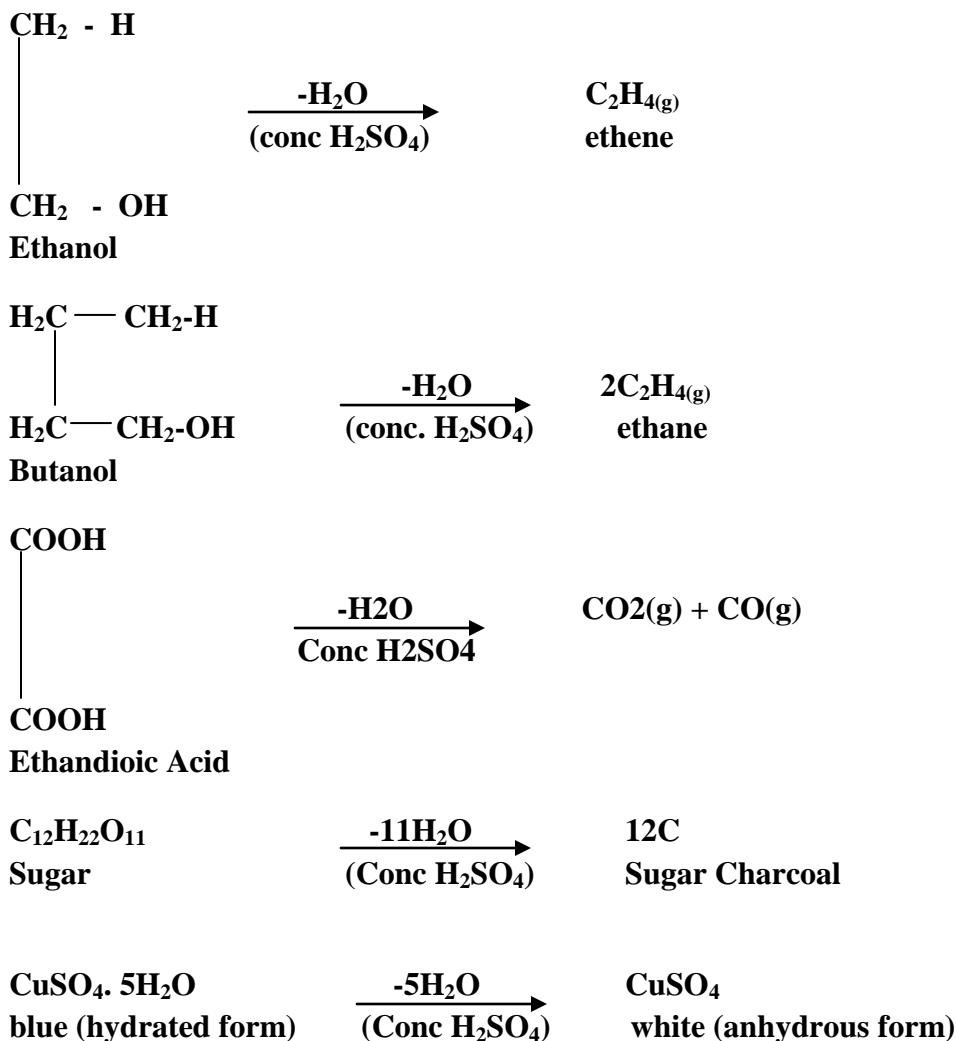


Sulphur is oxidized to sulphur (iv) oxide.

- e. H_2SO_4 acid also oxidizes H_2S gas to Sulphur and it is reduced to sulphur (iv) oxide.



3. Tetraoxosulphate (vi) acid is a powerful dehydrating agent because it can remove water molecule or molecules from compounds like ethanol, ethandioic acid, sugar etc. As a result of the dehydration, the compound is decomposed. The formation of anhydrous salt when the water of crystallization is removed from a hydrated salt is another form of dehydration. The corrosive action of the acid on paper, wood, human skin and cloth is due to its dehydration property.



Uses Of Tetraoxosulphate (VI) Acid

1. It is used in the manufacture of these fertilizers.

$(\text{NH}_4)_2\text{SO}_4$ - Ammonium tetraoxosulphate (vi)

$\text{Ca}(\text{H}_2\text{PO}_4)_2$ - Calcium hydrogen tetraoxophosphate (v).

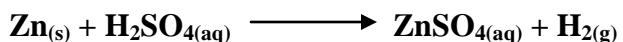
2. It is used in the making of pigments such as: TiO_2 - Titanium (iv) oxide and BaSO_4 - Barium tetraoxosulphate (vi). These pigments are used in the manufacture of paints and dyes.
3. It is an electrolyte used in batteries and lead accumulators.
4. It is used as a drying agent for some gases and in nitration of compounds used in making explosives.
5. It is used in the production of metallic tetraoxosulphate (vi), trixonitrate (v) acid and hydrochloric acid.

The Tetraoxosulphate (vi) Salts

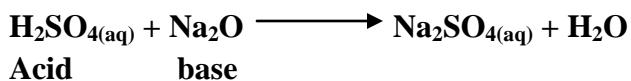
They are either normal or acid salts. Most of them are crystalline and dissolve in water with ease. PbSO_4 , CaSO_4 , Hg_2SO_4 and BaSO_4 are the anhydrous tetraoxosulphate (vi). They cannot dissolve in water.

Preparation of Soluble Tetraoxosulphate (vi)

1. By the action of dilute H_2SO_4 acid on metal



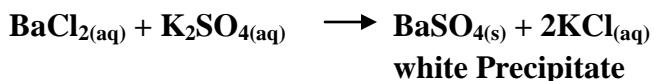
2. By neutralization reaction



3. By the action of dilute H_2SO_4 acid on trioxocarbonate (iv)



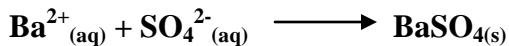
Insoluble Tetraoxosulphate (vi) is prepared by double decomposition method.



Metals	Method of preparation
Potassium	By the action of dilute tetraoxosulphate (vi) acid on the oxides, hydro xides or the trioxocarbonate (iv) of the metals.
Sodium Calcium	
Magnesium	By the action of dilute tetraoxosulphate (vi) acid on the metals, their oxides or their trioxocarbonate (iv)
Aluminium	Aluminium, however requires the action of concentrated tetraoxosulphate (vi) acid.
Zinc Iron	
Lead	By double decomposition
Copper	By the action of hot concentrated tetraoxosulphate (vi) acid on the metals.
Mercury Silver	

Test For Tetraoxosulphate (vi) Ion.

Pour some of the solution of the unknown salt into a test tube. Acidify it with dil HCl acid and add a few drops of BaCl₂ solution. A white precipitate of BaSO₄ salt will be formed if the unknown solution contains SO₄²⁻. The white precipitate will not dissolve in an excess of the acid.



The acidification of the unknown solution is of vital importance in that it will prevent the precipitation of the barium compounds like barium trioxocarbonate (iv) or barium trioxosulphate (iv) which are soluble in an excess of the acid unlike the tetraoxosulphate (vi)

It is important to know that dilute HNO₃ acid and Ba(NO₃)₂ can be used instead of dilute HCl and BaCl₂ in testing SO₄²⁻ ion.

Exercises:

1. Write two chemical properties of sulphur.
2. Name the four allotropic forms of sulphur.
3. Write a balanced chemical equation only to show the laboratory preparation of H₂S gas.

4. Of what importance is kipp's apparatus to a laboratory chemist?
5. What are the two chemical tests of hydrogen sulphide?
- 6a. Name two oxides of sulphur.
- b. Use one chemical equation only to show how each of them can be prepared.
7. Use one chemical equation only to show:
 - a. The acidic property of sulphur (iv) oxide.
 - b. The reducing property of sulphur (iv) oxide
 - c. The oxidizing property of sulphur (iv) oxide.
8. Trioxosulphate (iv) acid is dibasic.
 - a. Write two possible ways this acid can ionize.
 - b. Use equations only to show two types of salt it can form.
 - c. Name the IUPAC nomenclature of the two salts.
9. Why is it that sulphur (vi) oxide is not dissolved directly in water to form H_2SO_4 acid in the contact process?
10. Write two physical properties of tetraoxosulphate (vi) acid.
11. Use two chemical equations only to show the oxidizing property of tetraoxosulphate (vi) acid.
12. Tetraoxosulphate (vi) acid is a dehydrating agent. Use two chemical equations only to illustrate this.
13. Write a method of preparing:
 - (a) A named soluble tetraoxosulphate (vi) salt.
 - (b) A named insoluble tetraoxosulphate (vi) salt.
14. Why is the acidification process very necessary while testing for tetraoxosulphate (vi) ion?

CHAPTER 12

OXIDATION REDUCTION REACTIONS

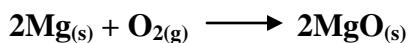
Oxidation Reduction reaction is also known as “REDOX” reaction. It comprises of both oxidation and reduction reactions. Read various definitions of oxidation and reduction reactions in Part 2, Chapter 2 of this book.

12.1 Definitions of Oxidation and Reduction

a. Oxidation and Reduction in terms of Oxygen.

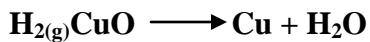
Oxidation is the addition of oxygen and reduction is the removal of oxygen.

i. Burning of magnesium in oxygen



Conversion of magnesium metal to its oxide by the addition of oxygen is Oxidation. Magnesium is oxidized.

ii. Passing hydrogen gas over heated copper(ii) oxide: $\text{H}_{2(\text{g})} + \text{CuO}$ to Cu , by the removal of oxygen is reduction. CuO is reduced to copper metal by hydrogen gas.



(b) Oxidation and Reduction in terms of hydrogen: Oxidation is the removal of hydrogen from a substance, while reduction is the addition of hydrogen to a substance e.g.

i. Decomposition of ammonia:

$2\text{NH}_{3(\text{g})} \longrightarrow \text{N}_{2(\text{g})} + 3\text{H}_{2(\text{g})}$. The conversion of NH_3 to N_2 by the removal of hydrogen is oxidation. Ammonia is oxidized to nitrogen.

ii. Formation of ammonia:

$\text{N}_{2(\text{g})} + 3\text{H}_{2(\text{g})} \longrightarrow 2\text{NH}_{3(\text{g})}$. The conversion of N_2 to NH_3 by the addition of H_2 is reduction. Nitrogen is reduced to ammonia by the addition of H_2 gas.

c. Oxidation and Reduction in terms of Electronegativity of elements.

Oxidation is the addition of an electronegative element to another element, while reduction is the removal of an electronegative element from a compound. Oxidation and reduction involve all the elements, either as being electro-negative or electro-positive. In the periodic table, electropositive elements are on the right, while those on the left are electronegative.

In non-metals: P, C, S, I, Br, N, Cl, O, F

Increasing electronegativity →

In metals and hydrogen:

← Increasing electropositivity

K, Na, Ca, Mg, Al, Zn, Fe, Sn, Pb, H, Cu, Hg, Ag, Agu

Increasing electronegativity →

i. Formation of Sodium Chloride:

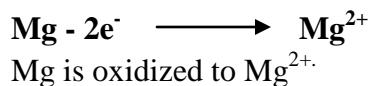
$2\text{Na}_{(\text{s})} + \text{Cl}_{2(\text{g})} \longrightarrow 2\text{NaCl}_{(\text{s})}$. The conversion of sodium to sodium chloride by the addition of chlorine is oxidation, because chlorine is more electronegative than sodium. Sodium is oxidized to sodium chloride.

ii. Action of a reactive metal on an acid:

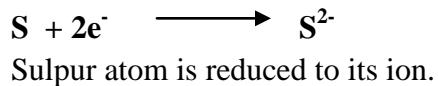
$\text{Fe}_{(\text{s})} + 2\text{HCl}_{(\text{aq})} \longrightarrow \text{FeCl}_{2(\text{aq})} + \text{H}_{2(\text{g})}$. The conversion of HCl to H_2 gas by the removal of chlorine is reduction, because in HCl, chlorine is more electronegative than hydrogen. HCl is reduced to H_2 gas.

b. Oxidation and Reduction in terms of Electron transfer

Oxidation is the process of electron loss; Reduction is the process of electron-gain. This concept of oxidation and reduction is limited to reactions between metals and non-metals, where electrons are transferred from the valence shell of the metal atom to the valence shell of the non-metal atom, leading to the formation of electrovalent or ionic compounds. Magnesium atom loses its two valence electrons to form magnesium ion.

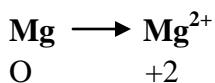


ii. When sulphur (non-metal) gains two electrons in order to complete its valence shell, it forms sulphur ion.

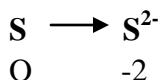


e. Oxidation and Reduction in terms of Oxidation Number (O.N.)

Oxidation is an increase in oxidation number. Reduction is a decrease in oxidation number. The oxidation number of an element is the charge on its atom or ion. The oxidation number of a neutral atom is zero, because the number of protons in its nucleus is equal to the number of electrons in its shell.



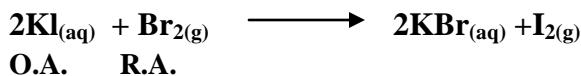
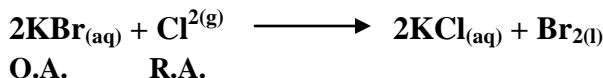
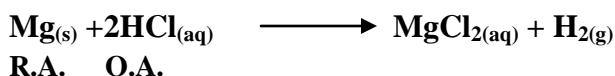
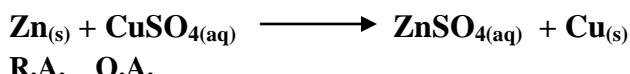
The oxidation number of Mg increases from 0 to +2. When sulphur atom ionizes by gain of two electrons; it forms sulphur ion.



The oxidation number of sulphur ion, S^{2-} is -2. Hence the O.N. of sulphur decreases from 0 to -2. It is reduction.

12.2 Redox Reactions

Some of the reactions that involve oxidation and reduction reactions are as follows:



12.3 Oxidation No of Central Elements in Some Compounds

Rules for Assigning Oxidation Numbers (O.N.)

1. The O.N. of an atom in an uncombined free element, whether monatomic or polyatomic is zero; e.g. Noble gases (He, Ne etc.), metals (Na, Cu, etc.), solid non-metals (C, S, etc.), and molecules such as (H_2 , O_2 , O_3 , Cl_2 , P_4 , S_8 etc.).
2. The O.N. of one oxygen atom is -2 in its compounds, (except in peroxides: H_2 , O_2 , Na_2O or O_2^{2-} , where it is -1 and in its binary compound with fluorine, OF_2 where it is +2).
3. The oxidation number of one hydrogen atom is +1 in its compounds (except in hydrides of metals, e.g. NaH , where it is -1).
4. Group I alkali metals (Li, Na, K, etc.) have O.N. of +1, group II alkaline - earth metals (Mg, Ca, etc.) have O.N. of +2, while group III metals (Al, etc.) have O.N. of +3 in all their compounds.

5. The halogens (F, Cl, Br, I, etc.) in group VII have O.N. of -1 per atom in their binary compounds only (except when bonded to a more electronegative atom - as in IF_7 , where the O.N. of iodine atom is +7).
6. The O.N. of the atom in a monatomic ion is equal to the charge on it; e.g. the O.N. of sodium atom in Na^+ is +1; that of Al in Al^{3+} is +3; and that of S in S^{2-} is -2.
7. In any complex ion, the sum of the O.N.'s of all the atoms is equal to the charge on the ion.
8. In any molecule or compound, the sum of the oxidation numbers of all the atoms adds to zero.

Worked examples

Example 1:

Calculate the oxidation number of sulphur in tetraoxosulphate (vi) ion.

The oxidation number of SO_4^{2-} = -2

$$\therefore (\text{oxidation number of S}) + 4(\text{oxidation number of O}) = -2$$

$$\text{oxidation number of S} = 6$$

$$\text{S} + [4 \times (-2)] = -2$$

$$\text{S} - 8 = -2$$

$$\text{S} = -2 + 8 = +6$$

Example 2:

Calculate the oxidation number of sulphur in K_2SO_4 . Write the IUPAC nomenclature of the compound.



$$(+1 \times 2) + \text{S} + (-2 \times 4) = 0$$

$$2 + \text{S} + (-8) = 0$$

$$2 + \text{S} - 8 = 0$$

$$\text{S} = 0 + 8 - 2$$

$$\text{S} = +6$$

$\therefore \text{K}_2\text{SO}_4$ = Dipotassium tetraoxosulphate (vi).

Example 3:

Calculate the oxidation number of carbon in calcium trioxocarbonate (iv).



$$+2 +\text{C} + (-2 \times 3) = 0$$

$$+2 +\text{C} + (-6) = 0$$

$$+2 +\text{C} - 6 = 0$$

$$\text{C} = 0 + 6 - 2$$

$$\text{C} = +4$$

Example 4:

Calculate the oxidation number of manganese in KMnO_4 . Then write the IUPAC nomenclature of the compound.



$$+1 +\text{Mn} + (-2 \times 4) = 0$$

$$+1 +\text{Mn} + (-8) = 0$$

$$+1 +\text{Mn} - 8 = 0$$

$$\text{Mn} = 0 + 8 - 2$$

$$\text{Mn} = +7$$

$\therefore \text{KMnO}_4$ = Potassium tetraoxomanganate (vii).

Example 5:

Find the oxidation number of Bromine in HBrO_3 and write the IUPAC nomenclature of the acid.



$$+1 +\text{Br} + (-2 \times 3) = 0$$

$$+1 +\text{Br} + (-6) = 0$$

$$+1 +\text{Br} - 6 = 0$$

$$\text{Br} = 0+6-1$$

$$\text{Br} = +5$$

$\therefore \text{HBrO}_3$ = Trioxobromate (v) acid.

Example 6:

Find the oxidation number of Chromium in $\text{Cr}_2\text{O}_7^{2-}$. Write the IUPAC nomenclature of the radical.



$$2\text{Cr} + (-2 \times 7) = -2$$

$$2\text{Cr} + (-14) = -2$$

$$2\text{Cr} - 14 = -2$$

$$2\text{Cr} = -2+14$$

$$2\text{Cr} = +12$$

$$\text{Cr} = 12 = +6$$

2

$\therefore \text{Cr}_2\text{O}_7^{2-}$ = Heptaoxodichromate (vi) ion.

Compound	Conventional Name	IUPAC Name
K_2SO_4	Potassium sulphate	Dipotassium tetraoxosulphate(vi)
CaCO_3	Calcium carbonate	Calcium trioxocarbonate(iv)
KMnO_4	Potassium permanganate	Potassium tetraoxomanganate(vii)
H_2SO_4	Sulphuric acid	Tetraoxosulphate (vi) acid
HNO_3	Nitric acid	Trioxonitrate(v) acid
NaBrO_3	Sodium bromate	Sodium trioxobromate(v)

12.4 Oxidizing and Reducing Agents

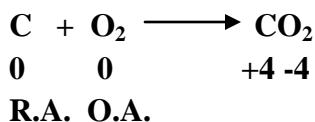
A reducing agent (O.A.), (a reductant or reducer) is the substance that loses or donates electrons, or gets oxidized; or whose oxidation number increases.

An oxidizing agent (R.A.), (an oxidant or oxidizer) is the substance that gains or accepts electrons, or get reduced; or whose oxidation number decreases.

Types of Redox Reactions

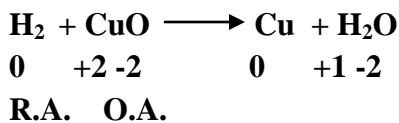
1. All addition and combustion reactions are redox.

E.g. Combustion of carbon



The oxidation number of Carbon has increased from 0 to +4 in CO₂. This is oxidation. Carbon is a reducing agent; it has been oxidized to CO₂. The O.N. of oxygen atom has decreased from 0 in O₂ to - 2 in CO₂; this is reduction. Oxygen gas is an oxidizing agent; it has been reduced to an oxide, CO₂.

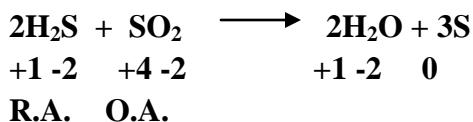
2. All displacement reactions are redox E.g.



When hydrogen gas is passed over heated copper (ii) oxide, hydrogen gas is oxidized to water by an increase in O.N. from 0 to +1, while copper(ii) oxide is reduced to copper by a decrease in O.N. from +2 to 0. Hydrogen gas is the reducing agent, while copper (ii) oxide is the oxidizing agent.

Note: The oxidation number of oxygen atom has not changed in the reaction above.

3. Some double decomposition reactions are redox. E.g.



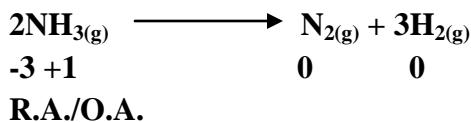
In this reaction, O.N. of sulphur atom in H₂S has increased from -2 to 0; it has been oxidized to sulphur; H₂S is the reducing agent. Sulphur atom in SO₂ has been reduced to sulphur from +4 to 0; SO₂ is the oxidizing agent.

Note: The oxidation number of hydrogen and oxygen did not change.

Generally, metals, hydrogen and carbon are examples of reducing agents, while oxygen and the halogens are oxidizing agents. In fact, the activity series arranges the metals and hydrogen in decreasing order of their reducing ability. Hence, potassium is the strongest reducing agent, while flourine is the strongest oxidizing agent.

4. Some decomposition reactions are redox

a. Decomposition of ammonia gas:



In this reaction, the O.N. of nitrogen atom has increased from -3 in NH₃ to 0 in N₂ gas. This is oxidizing; hence NH₃ is a reducing agent. At the same time, the O.N. of hydrogen atom has decreased from +1 in NH₂ to 0 in H₂ gas. This is reduction; hence, NH₃ is an oxidizing agent.

When a chemical substance acts as an oxidizing agent and as a reducing agent in the same reaction, it is said to disproportionate; the redox reaction is called disproportionate reaction.

Exercises

1. Define ‘Oxidation’ and ‘reduction’ in terms of;
 - (a) oxygen (b) hydrogen (c) electron transfer
2. Calculate the oxidation numbers of these elements in the given compound. Then write the IUPAC nomenclature of the compound.
 - (a) Sulphur in H₂SO₄; H₂S₂O₇. (b) Chromium in K₂Cr₂O₇
 - (c) Sulphur Al₂S₃ (d) Chromium in Cr₂O₇²⁻
 - (e) Nitrogen in Pb(NO₃)₂; Al(NO₃)₃
 - (f) Sulphur in FeSO_{4.7}H₂O (g) Copper in [Cu(NH₃)₄]²⁺
- 3a Write a balanced equation on a redox reaction.
- b. Identify the reducing and an oxidizing agents by writing under the substance R.A. (for reducing agent) and O.A. (for oxidizing agent).

CHAPTER 13

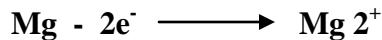
IONIC THEORY

13.1 Electrovalent and Covalent Compounds: (Read Part I, Chapter 7 of this book)

Arrhenius was the first man who presented the ionic theory in 1887 to explain the phenomenon of electrolysis. The theory pointed out that when an electrolyte is dissolved in water or melted, some or all the electrolyte dissociate or break up into freely moving charge particles called ions. The process of dissociation into ions is known as ionization.



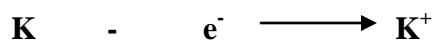
Ionization can also be defined as the process of electron loss by the metals.



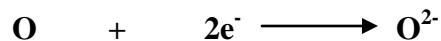
Hydrogen ion (H^+), ammonium ion (NH_4^+) and metallic ions have positive charges. The non-metallic ions and hydroxyl ions have negative charges. The valency of an atom or group of atoms is the number of electrical charges carried by the atom or group.

Eg.	Atoms or Group of atoms	Valency
	Ca^{2+}	+2
	NH_4^+	+1
	K^+	+1
	SO_4^{2-}	-2

Metallic ion is formed by the corresponding atom by the loss of electrons equivalent to its valency



Non-metallic ion is formed from the corresponding atom by the gain of a number of electrons equal to the valency of the atom:



An ion is a charged atom or a group of atoms (radical). Radical is a charged group of atoms:

Eg. SO_4^{2-} , HSO_4^- , NH_4^+ , CO_3^{2-} , HCO_3^- etc.

The current passing through the electrolyte is carried by the moving ions to the electrodes. Electric current or electricity is the moving electrons. It has great effect on matters. If it is passed through an electrolyte, the freely random moving ions are halted. The positive ions migrate to the cathode because they are attracted to it while the negative ions migrate to the anode.

13.2 Some common terms used in the study of electrolysis are:

Electrolysis: is the chemical decomposition of a compound by the passage of electric current through a molten or solution of the compound.

Electrolyte: is a compound in molten or in solution form which can conduct electricity and is decomposed by the electric current. Eg. Salts, alkalis and acids.

Concentrated H_2SO_4 (free from water) is not an electrolyte. It is an electrolyte in its diluted form.

Non-electrolyte: is a compound which cannot conduct electricity in solution or in molten form. It cannot allow electricity to pass through it. It cannot therefore be decomposed by electric current. They are covalent compounds and contain no ions. Eg. Ether, benzene etc.

Electrodes: are the two poles of carbon or metal at which electric current enters or leaves the electrolyte.

Anode: is the positive electrode through which the electrons leave the electrolyte.

Cathode: is the negative electrode through which the electrons enter the electrolyte.

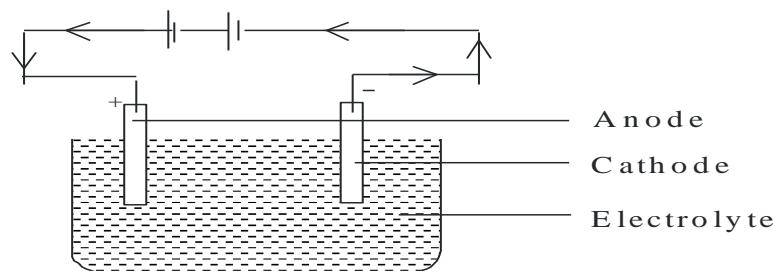


Fig. 13.1 A simple electrolytic cell.

EXAMPLES OF IONIZATION

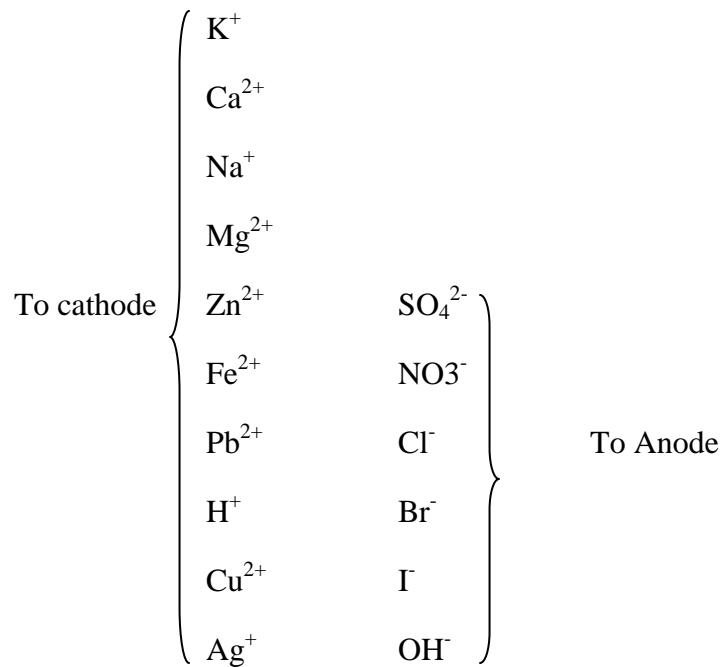
COMPOUNDS	IONS	
Tetraoxosulphate (vi) acid	2H^+	SO_4^{2-}
Sodium Chloride	Na^+	Cl^-
Sodium hydroxide	Na^+	OH^-
Copper (ii) tetraoxosulphate (vi)	Cu^{2+}	SO_4^{2-}
Ammonium Chloride	NH_4^+	Cl^-

Ionization is complete in strong electrolyte while in weak electrolyte ionization is only slight. Most of the electrolyte, exist in solution in form of unionized molecules eg Ethanoic acid and most organic acids.

Electrolytic reactions are redox reactions. Anodic half reaction (ie reaction that occur at the anode) is oxidation reaction while the cathodic half reaction is a reduction reaction.

13.3 Selective Discharge of Ions: When two or more ions of similar charge are present in a solution under similar conditions, one is selected for discharge at each of the electrodes. The preferential selection is dependent on the following factors.

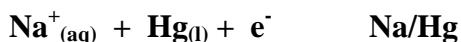
1. Position of the ions in the electrochemical series (ECS)



The ions lower in the electrochemical series (ECS) will be discharged in preference to the ion higher up in the ECS if all factors are constant. Thus, in Na_2SO_4 solution, containing H^+ (from H_2O) and Na^+ , H^+ is discharged in preference to Na^+ because it is lower in the ECS and will require less energy for it to be discharged. The anions present are OH^- and SO_4^{2-} . OH^- will be preferentially selected for discharge because of same reason.

2. **Concentration:** Increase in the concentration of an ion effects its discharge, Eg. In concentrated HCl , containing OH^- (from H_2O) and Cl^- as anions, Cl^- is preferentially selected for discharge due to concentration effect. It is important to note that it is only in the electrolysis of chlorides that concentration effect overtakes the position in E.C.S.
3. **Nature of the electrodes:** Electrode may be inert or active. Inert electrodes (eg Platinum, carbon electrodes etc) take no part in the electrolytic reaction. Active electrodes take part in the electrolytic reaction and they influence the ion that will be discharged due to their strong affinity for certain ion.

Eg. In the electrolysis of brine (NaCl solution) using Pt electrodes, H^+ is selected for discharge in preference to Na^+ . If Mercury cathode is used, Na^+ is discharged and it associates with the mercury to form sodium amalgam



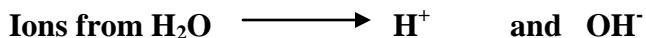
Due to the fact that Hg has much affinity to sodium, Na^+ will require less energy for its discharge than H^+ .

In the electrolysis of CuSO_4 solution using Cu electrodes (active electrodes) the anions SO_4^{2-} and OH^- migrate to the anode (Cu) but none of them is discharged. The Cu anode goes into solution and is deposited at the cathode.

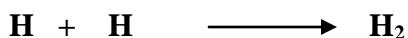
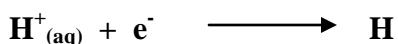
Electrolysis of Molten / Fused Salts and Solutions of Electrolytes:

- 13.4 **Electrolysis of acidified H_2O or dil. H_2SO_4 :** is carried out in the apparatus popularly known as Hoffman voltmeter.

Electrodes are made of Platinum foil.



At the Cathode H^+ migrates to the cathode where it gains an electron and is reduced to hydrogen atom(H). The latter pairs up to form hydrogen molecules.



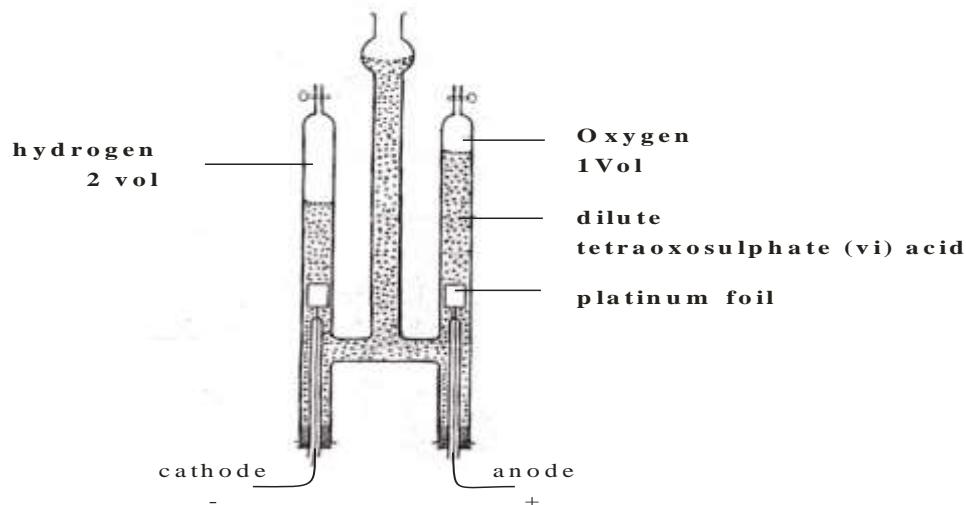
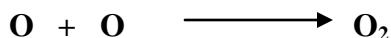
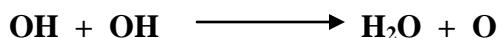
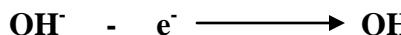


Fig. 13.2 Hoffman Voltmeter

At the Anode: SO_4^{2-} and OH^- both migrate to the anode. OH^- is discharged in preference to SO_4^{2-} because it is lower in the E.C.S.



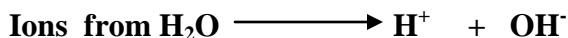
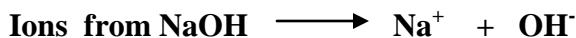
After electrolysis, the ratio of hydrogen to oxygen produced is 2:1. The discharge of OH^- disturbs the ionic equilibrium of water. More water therefore ionizes to restore it.



The total amount of H_2SO_4 acid in the electrolytic cell remains constant but its concentration increases due to the removal of water.

13.5. Electrolysis Of Sodium Hydroxide Solution: The apparatus employed in this electrolysis is Hoffman voltmeter.

Electrodes: Platinum foil



Cathode	Anode
<p>Both Na^+ and H^+ migrates to the cathode. H^+ is discharged in preference to Na^+ because it is lower in the ECS</p> $\text{H}_{(\text{aq})}^+ + \text{e}^- \longrightarrow \text{H}$ $\text{H}_{(\text{g})} + \text{H}_{(\text{g})} \longrightarrow \text{H}_2$ <p>More H_2O ionizes to restore H^+ since its discharge disturbs the ionic equilibrium of H_2O</p> $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$ <p>Exercise OH^- produced combines with the incoming Na^+ to cause an increase in the concentration of NaOH.</p> <p>Conc of NaOH increases</p>	<p>OH^- ion migrates to the anode where it is oxidized by loss of election</p> $\text{OH}_{(\text{aq})}^- - \text{e}^- \longrightarrow \text{OH}_{(\text{g})}$ $\text{OH}_{(\text{g})}^- + \text{OH}_{(\text{g})}^- \longrightarrow \text{H}_2\text{O}_{(\text{l})} + \text{O}_{(\text{g})}$ $\text{O}_{(\text{g})}^- + \text{O}_{(\text{g})}^- \longrightarrow \text{O}_2_{(\text{g})}$ <p>The discharge of OH^- and the migration of Na^+ to the cathode, causes the decrease in the concentration of NaOH</p> <p>Conc of NaOH decreases</p>

The products of electrolysis of NaOH solution are 2 volumes of hydrogen and 1 volume of oxygen. After the electrolysis, the total amount of NaOH in the electrolytic cell is unchanged but its concentration is increased due to the removal of water molecules.

13.6 Electrolysis of Sodium Chloride Solution or (Brine)

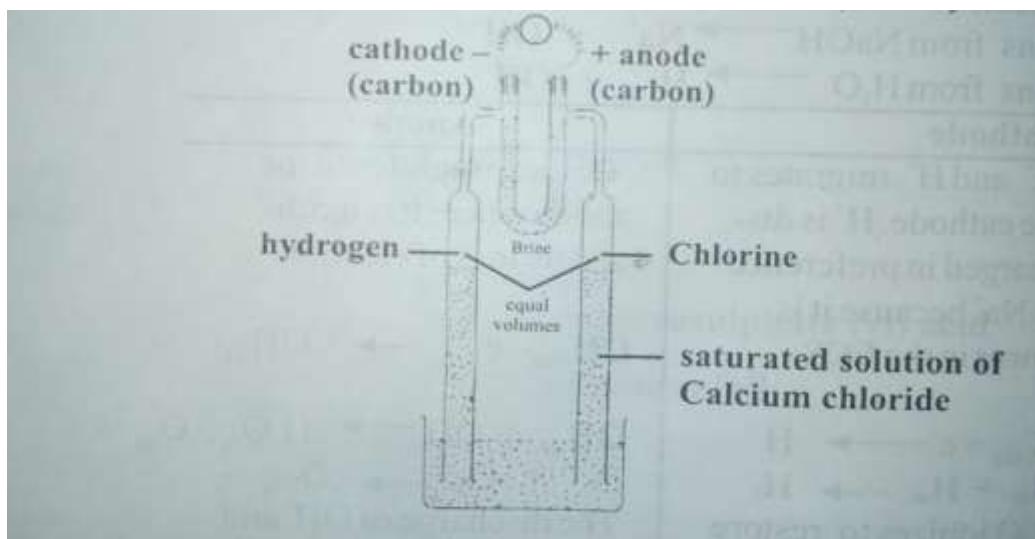


Fig. 13.3

Electrodes: The anode must be carbon to resist the attack of Cl_2 but the cathode may be Pt or carbon

Ions present from



Cathode	Anode
<p>Both Na^+ and H^+ migrate to the cathode. H^+ is discharged due to its lower position in E.C.S.</p> $\text{H}^+ + \text{e}^- \longrightarrow \text{H}$ $\text{H} + \text{H} \longrightarrow \text{H}_2$ <p>Or $2\text{H}^+ + 2\text{e}^- \longrightarrow \text{H}_2$</p> <p>More water ionizes to restore H^+ since its discharge disturbs the ionic equilibrium of H_2O.</p> $\text{H}_2\text{O} \longrightarrow \text{H}^+ + \text{OH}^-$ <p>Excess OH^- so produced react with the incoming Na to cause the increase in the conc of NaOH</p>	<p>Both Cl^- and OH^- migrate to the anode. Cl^- is preferentially discharged because of its higher concentration.</p> $\text{Cl}^- - \text{e}^- \longrightarrow \text{Cl}$ $\text{Cl}^- + \text{Cl}^- \longrightarrow \text{Cl}_2 \text{ or}$ $2\text{Cl}^- - 2\text{e}^- \longrightarrow \text{Cl}_2$

The products of the electrolysis of brine are equal volumes of hydrogen and chlorine.

NOTE: If molten NaCl is electrolysed, the only ions present are Na^+ and Cl^- . The electrolytic products will be Na at the cathode and Cl_2 at the anode.

13.7 Electrolysis Of Copper (ii) Tetraoxosulphate (vi) Solution (i.e. CuSO_4 Soln):

Electrodes: Carbon or Platinium



Cathode	Anode
<p>Both Cu^{2+} and H^+ migrate to the cathode. Due to relative lower position of Cu^{2+} in the E.C.S., It is discharged in preference to H^+</p> $\text{Cu}^{2+} + 2\text{e}^- \longrightarrow \text{Cu}_{(s)}$	<p>Both SO_4^{2-} and OH^- migrate to the anode. OH^- is discharged because it is lower in the E.C.S.</p> $\text{OH}_{(\text{aq})} - \text{e}^- \longrightarrow \text{OH}_{(\text{g})}$ $\text{OH}_{(\text{g})} + \text{OH}_{(\text{g})} \longrightarrow \text{H}_2\text{O}_{(\text{g})} + \text{O}_{(\text{g})}$ $\text{O}_{(\text{g})} + \text{O}_{(\text{g})} \longrightarrow \text{O}_2$

After the electrolysis, Cu is deposited at the cathode while O₂ is liberated at the anode. The solution becomes acidic with H₂SO₄ acid:

Note: If copper anode is used:



Cathode	Anode
<p>Cu is deposited</p> <p>$\text{Cu}^{2+} + 2\text{e}^- \longrightarrow \text{Cu}$</p>	<p>If Cu anode is used instead of C or Pb anode, there will be anode oxidation. The Cu (an anode) passes into Solution from the anode to form</p> Cu^{2+} $\text{Cu} - 2\text{e}^- \longrightarrow \text{Cu}^{2+}$ <p>The Cu²⁺ migrates to the cathode where it is deposited.</p>

It is important to note that with Pt or C electrodes, the blue colour of solution (CuSO₄ soln) will fade because the Cu²⁺ from the solution is deposited on the cathode and the remaining solution becomes acidic. With Cu electrodes, the blue colour of the solution will not change because the electrolysis merely transferred copper from anode to cathode.

Exercises:

1. Define ionization in two ways:
(b) Write an equation each to explain each of the two definitions.
2. Define these terms:
(a) Electrolysis (b) Electrolyte (c) Non-electrolyte
(d) Electrode (e) Cathode (f) Anode
3. What are the three factors on which the selective discharge of ions depend?
4. What is the difference between

(a) A molten or fused salt and (b) Salt solution?

5. Briefly describe the electrolysis of

a. Molten NaCl b. NaCl solution

c. Are the products of the electrolysis of these two electrolytes the same?

d. Name the electrolytic products of each.

6. Why is it that in the electrolysis of CuSO₄ solution using copper anode, the blue colour of the electrolyte does not change after the electrolysis?

CHAPTER 14

FARADAY'S LAW OF ELECTROLYSIS: USES OF ELECTROLYSIS

In 1833, Michael Faraday put forward two laws to describe the quantitative aspects of electrolysis. From his experimental findings, he pointed out that the quantity of products liberated at the electrodes during electrolysis is dependent on three factors. These factors are:-

- * The magnitude of the steady current passing.
 - * The time of passing the steady current.
 - * The charge on the ion of the element liberated.

14.1 Faraday's First Law: Faraday first law of electrolysis states that the mass (m) of an element discharged during an electrolysis is directly proportional to the quantity of electricity (Q) passed.

M α Q

The unit of quantity of electricity (Q) is the coulomb.

Quantity of electricity - current x time
 (Coulomb) (ampere) (second)

$$Q = I \times t$$

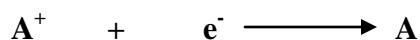
Thus $m \alpha$ I x t

or $m = Eit$ (where E is a constant)

When m is expressed in grams, the constant, E is the mass in gram of the substance liberated by the passage of I coulomb (C) of electricity (ie a current of I ampere passed for I second and is called the electrochemical equivalent of the substance.

Faraday's Second Law: Faraday's second law of electrolysis states that when the same quantity of electricity is passed through different electrolytes, the relative number of moles of the element discharged are inversely proportional to the charges on the ions of each of the elements respectively.

His experiments showed that the minimum quantity of electricity required of a single-charged ion is 96,500 coulombs. 96,500 Coulombs will liberate 1 mole of hydrogen ion (1.008g) or 1 mole of chloride ions (35.5g). This quantity of electricity is called the faraday and it is denoted by F.





The charge of one mole of a single charged ion involves the transfer of one mole of electrons. Faraday = 1 mole of electrons (6.02×10^{23} electrons). If one Faraday discharges one mole of a univalent element such as hydrogen, potassium, sodium, silver, chlorine, it follows from the second law of electrolysis that 2 Faradays will discharge 1 mole of a divalent element such as zinc, calcium and 3 Faradays will liberate 1 mole of trivalent element such as aluminium.

Problems and Solutions

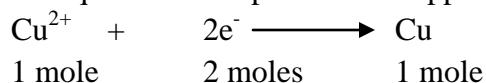
Question 1:

During the electrolysis of a copper(II) salt solution, 0.4 mole of copper was deposited. What quantity of electricity was passed during this process?

Answer:

Copper has a valency of 2

the equation for deposition of copper is:



From the equation:

1 mole of Cu^{2+} requires 2 mole of electrons \equiv 2 Faradays

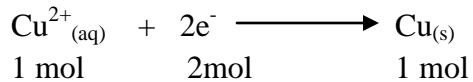
$$\begin{aligned} \text{1 mole of Cu} &= 2\text{F} \\ \text{0.4 mol of Cu} &= 2\text{F} \times \frac{0.4 \text{ mol}}{1 \text{ mol}} \\ &= 0.08 \text{ F} \end{aligned}$$

Question 2:

0.1 Faraday of electricity was passed through a solution of copper(II) tetraoxosulphate (VI). What maximum mass of copper is deposited on the cathode? ($\text{Cu} = 63.5$).

Answer:

Equation for deposition of copper is



From the equation

$$2\text{e}^- \circ 2\text{F} \circ 1 \text{ mol of Cu} = 63.5 \text{ g of Cu}$$

2 Faraday of electricity ° 63.5g of Cu (divalent)

$$\therefore \text{0.1 Faraday will deposit } 63.5 \text{ g} \times \frac{0.1 \text{ F}}{2 \text{ F}} \\ = 3.2 \text{ g of Cu}$$

Question 3:

Calculate the mass of silver deposited when a current of 2.6A is passed through a solution of a silver salt for 70 minutes.

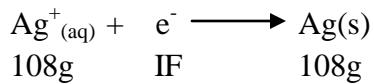
(Ag = 108; 1 Faraday = 96500C)

Answer:

Quantity of electricity used = 1 x t

$$Q = 1 \times t$$

$$Q = 2.6 \times 70 \times 60 = 10920 \text{ C}$$



Thus, 96,500 C (IF) liberated 108g of Ag ∴ 10920C will liberate?

96,500 C (IF) liberated 108g of Ag ∴ 10920C will

$$\begin{aligned} \text{Liberated } 108 \times \frac{10920}{96,500} \\ = 12.22 \text{ g} \end{aligned}$$

Question 4:

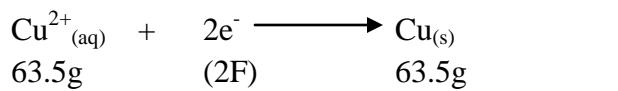
Calculate the mass of copper deposited when a current of a 0.9A is passed through a solution of a copper salt for 1 hr 20 mins.

(Cu = 63.5, / Faraday = 96500 C)

Answer:

$$Q = 1 \times t$$

$$Q = (0.9 \times 80 \times 60 = 4320 \text{ C})$$



(2 x 96500) C liberates 63.5g of Cu

$$\therefore 4320\text{C will liberate } \frac{63.5 \times 4320}{2 \times 96,500}$$

$$= 1.42\text{g}$$

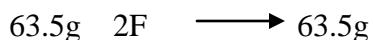
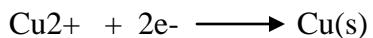
Question 5:

Three electrolytic cells containing solutions of copper(ii) tetraoxosulphate(vi), silver trioxonitrate (v) and tetraoxosulphate (vi) acid respectively are connected in series. A steady current is then passed through these solutions. If 6.55g of copper are deposited at the cathode of the copper electrolytic cell, calculate:

- The mass of silver deposited at the cathode of the silver electrolytic cell.
- The volume of hydrogen gas liberated at 25°C and 730 mm Hg in the third electrolytic cell. (Cu = 63.5, Ag = 108, 1 Faraday = 96,500 C; molar volume of H₂ at ST.P is 22.4 dm³)

Answer:

In the copper electrolytic cell,



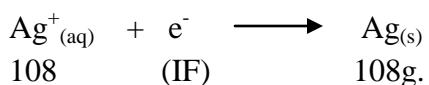
63.5g of Cu is discharged by (2 x 96,500) C

$$\therefore 6.55\text{g of Cu is discharged by } \frac{2 \times 96500 \times 6.55}{63.5}$$

$$= 19908 \text{ C}$$

\therefore Quantity of electricity used = 19908C

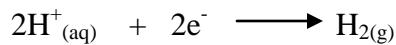
In the silver electrolytic cell;



96500C produced 108g of Ag

$$19908 \text{ C will produce } 108 \times \frac{19908}{96,500} = 22.28 \text{ g.}$$

In the third electrolytic cell



2F 22.4 dm³ at s.t.p

2 x 96500 C liberate 22400 cm³ of H₂

$$\therefore 19908 \text{ C will liberate } 22400 \times \frac{19908}{2 \times 96500} = 2311 \text{ cm}^3 \text{ of H}_2 \text{ at s.t.p.}$$

Convert this to its volume at 25°C and 730 mm Hg

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \dots \text{Equation 1}$$

Where P₁ = 760mm

V₁ = 2311cm³

T₁ = 273 K

P₂ = 730mm

T₂ = (25 + 273) K = 298 K

V₂ ?

From Equation 1

$$T_1 P_2 V_2 = P_1 V_1 T_2$$

$$V_2 = \frac{P_1 V_1 T_2}{T_1 P_2}$$

$$V_2 = \frac{760 \times 2311 \times 298}{730 \times 273} = 2626 \text{ cm}^3$$

The volume of hydrogen liberated at 730 mm Hg and 25°C is 2626 cm³.

14.2 Use of Electrolysis

Electrolysis is very important to the chemists or industrialists. It is widely applied in carrying out the following operations in the industry.

1. It is used in the preparation of some compounds like:
 - a. Sodium trioxochlorate (v)
 - b. NaOH, Cl₂ and H₂ from brine.
 - c. O₂ from the electrolysis of acidified water.
2. It is used in the purification of metals such as Cu, Hg, Ag, Au etc.
3. It is used in the extraction of many metals and non-metals. Some of them are; Na, Ca, Zn, K, Al, Cl₂, H₂, F₂ etc.
4. It is employed in the electroplating of one metal by another. This is to improve the appearance of the metal and to prevent it from rusting.

14.3 Electrochemical Cells: If a metal rod is placed in a solution of one of its salts, some of the atoms from the metal rod, will ionize and go into solution as positively charged ions leaving behind their valence electrons on the surface of the metal rod. At the same time, some of the metallic ions already in the solution will take up electrons from the metal rod and deposit themselves as neutral atoms on the rod, leaving behind an excess of anions on the salt solution. Depending on which reaction is more favoured the metal rod (electrode) becomes either negatively or positively charged, with respect to the solution. Thus a potential difference (differences in electric potential) known as the electrode potential of the metal is set up between the metallic electrode and the solution (electrolyte).

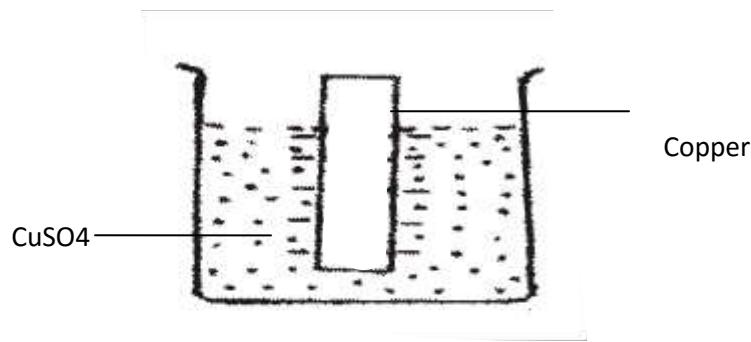


Fig. 14.1: Electrode Potential of Copper

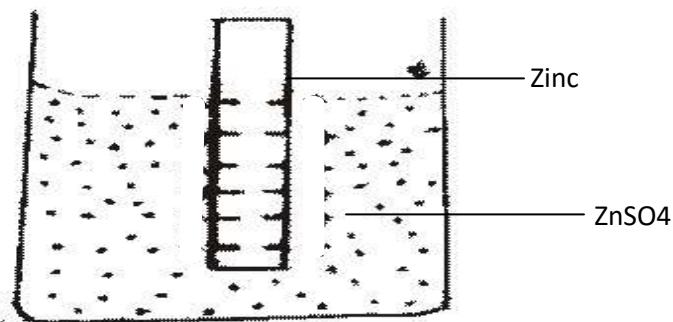


Fig. 14.2: Electrode Potential of Zinc

In Fig. 14.1, when copper plate is dipped into a copper (ii) tetraoxosulphate (vi) solution, the net reaction is that the positive Cu²⁺ ions from the solution acquire two electrons each from the copper plate to become deposited as neutral metallic copper atom on the plate. The copper plate or rod becomes positively charged due to deficiency in electrons relative to the solution which is now negatively charged due to excess SO₄²⁻.

A potential difference known as the electrode potential of copper is thus set up between the copper metal and the solution of its salt. In Fig. 14.2, the Zinc atoms on the plate will loss electrons to the plate and go into solution as Zn²⁺. The zinc plate becomes negatively charged due to the excess electrons while the solution becomes positively charged due to the excess Zn²⁺ going into the solution. There is a potential difference known as electrode potential of zinc set up between the zinc metal and the solution containing its ions. ***The electrode potential of an element is the potential difference between the element (electrode) and the solution containing its ions (electrolyte) in which it is immersed.*** The electrode potential varies from element to element depending on concentration and temperatures. Therefore for comparing electrode potential values of different elements, the standard electrode potential of the element is used.

Standard electrode potential is the potential difference set up when an electrode is in contact with one molar solution of its ion at 25°C. For an electrode potential of an element to be measured, it must be compared with other elements. Hydrogen electrode has been chosen as an arbitrary standard with an electrode potential of zero (0) at all temperature upon which other elements are determined. The + or - sign of the electrode potential of an element will depend on whether electron flows from or to the hydrogen electrode when the half cell of the element is connected to the hydrogen half cell.

Standard electrode potential (E°) of some elements are as follows:

Element	Its ion	Electrode potential (E°)
K	K^+	- 2.92 V
Ca	Ca^{2+}	- 2.87 V
Na	Na^+	- 2.71 V
Zn	Zn^{2+}	-0.76 V
$\frac{1}{2} H_2$	H^+	0.00 V
Cu	Cu^{2+}	+0.34 V
Ag	Ag^+	+0.80 V
Pt	Pt^{2+}	+1.12V
Au	Au^+	+1.68 V

Electrochemical Cells:

If the zinc electrode (Fig.14.2) is connected with wire to copper electrode (Fig. 14.1) electron flows from zinc (cathode) electrode to copper (Anode). This flow of electron or electric current is due to the difference in electrode potential of zinc and copper.

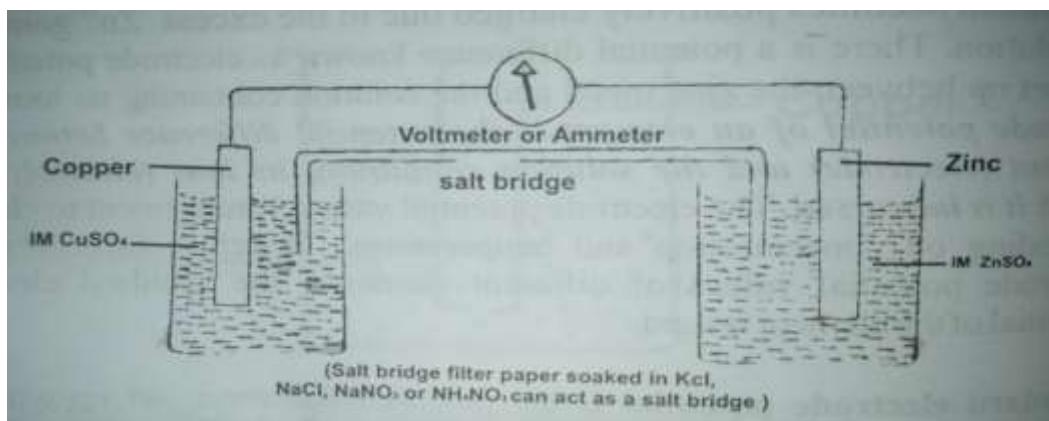


Fig. 14.3

As soon as the connection between the two half cells is made, the voltmeter needle moves or the bulb lights indicating that electricity has been generated. This device which converts chemical energy to electrical energy is called electrochemical cell or the galvanic cell. The salt bridge serves to prevent a charge built up by allowing ions to be transferred between the two solutions with an inverted U tube containing a salt such as $NaNO_3$ or KCl .

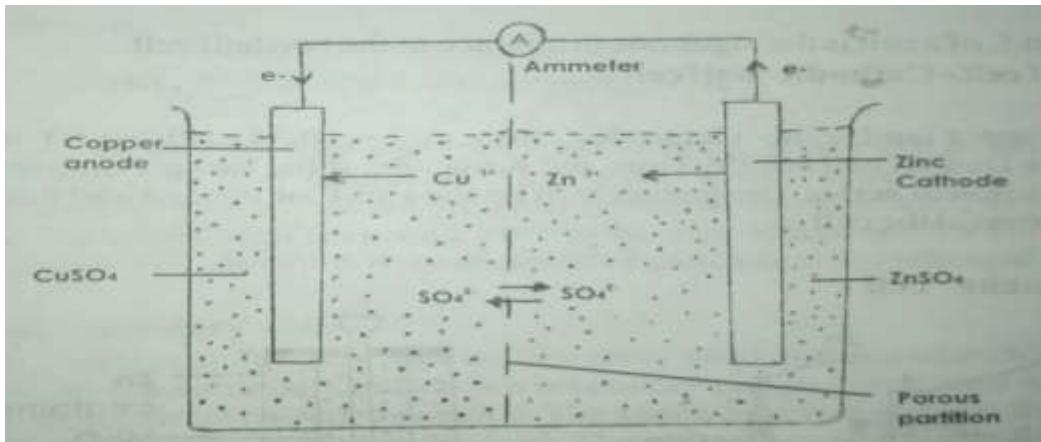
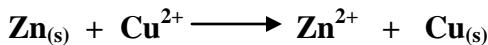
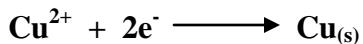
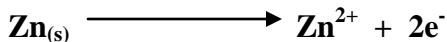


Fig. 14.4

In the Daniell cell, instead of the salt bridge, a porous partition is used for the exchange of ions between the two half cells. The direction in which electrons flow through the circuit depends on the position of the metal in the electrochemical series. The metal higher in the series will form the negative electrode of the cell (cathode) and the lower will form the negative electrode of the cell (cathode) and the lower one will form the positive electrode (anode). The overall cell reaction can be attained by combining the two half reactions.



A short hand notation is often used to describe cells. The substances involved are denoted by their formulae, with their concentration and pressure written in parenthesis following the particular formula. A single vertical line indicates contact between solid and liquid or liquid and gas. Two or more substances in the same solution are separated by commas. The salt bridge or porous partition or other electrolytic contact is indicated by a part of vertical lines.

Thus the cell in Fig. 8.3 or Fig. 8.4 is represented as:

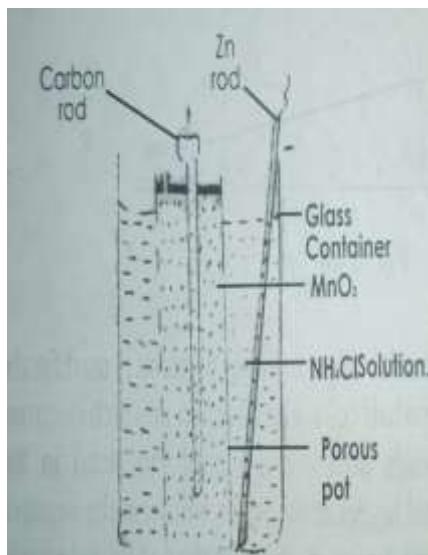


Thus the e.m.f. of a cell is the algebraic difference of the two half cells:

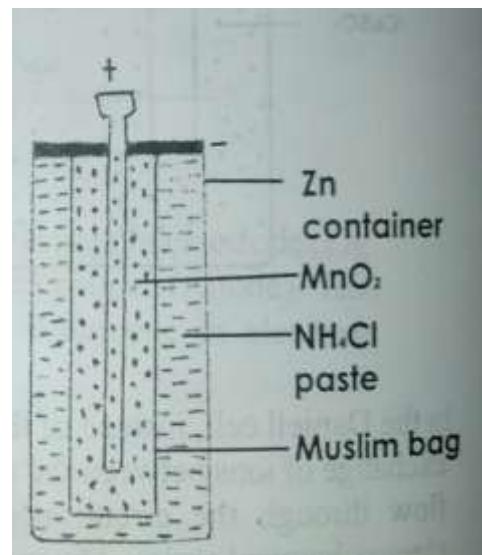
Anodic half cell - Cathodic half cell

Thus for copper and Zinc electrodes, the e.m.f. = +0.34 - (-0.76) = 1.10V The emf of a Daniell cell is 1.1V. Thus, the further apart the two metals are on the electrochemical series, the greater will be the e.m.f. of the cell and hence the more powerful the cell is.

The leclanche cell



(a) The wet cell



(b) The dry cell

Fig. 14.5

Another type of primary cell is the leclanche cell. Here the cathode is zinc, the anode is carbon rod and the electrolyte is ammonium chloride. They are two types: The wet cell and the dry cell.

In the dry cell, the electrolyte is NH_4Cl paste

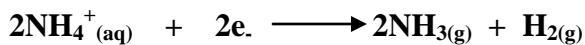
At the cathode:



Two electrons released pass through an external circuit performing useful work such as the lighting up of a touch bulb, the ringing of a door bell or the functioning of a transistor radio before arriving at the anode of the cell.

At the anode:

The ammonium ions accept the electrons to become reduced to ammonia and hydrogen.



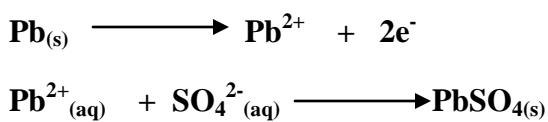
The hydrogen produced will accumulate at the anode and polarizes the anode and stops the flow of current. This is prevented by manganese (iv) oxide a depolarizer which oxidizes the hydrogen produced to water. The leclanche cell like the Daniell cell is a primary cell in that it produces electric current by using up the chemicals which it is composed of and cannot be recharged.

The secondary cell:

The lead accumulator which is commonly used in cars and motor cycle as car battery is a secondary cell. It is a secondary cell in that it must be charged by passing a direct current through it. The anode of accumulator is lead (iv) oxide, PbO_2 . The cathode is metallic lead. The electrolyte is dilute H_2SO_4 . When the two electrodes are connected to an external circuit, it produces electricity by discharging.

At the anode:

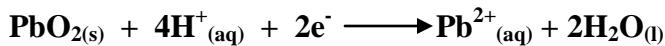
During the process of discharging the lead cathode releases two electrons and form Pb^{2+} and then combines with the SO_4^{2-} of the electrolyte to form PbSO_4 .



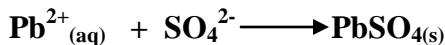
The electrons released pass round an external circuit as an electric current to be used for starting the car engine, car lamps etc. before arriving at the cathode.

At the cathode:

The electrons from the anode are accepted at the cathode where the lead (iv) oxide and the hydrogen ions from the electrolyte undergo reduction reaction to form Pb^{2+} and water.

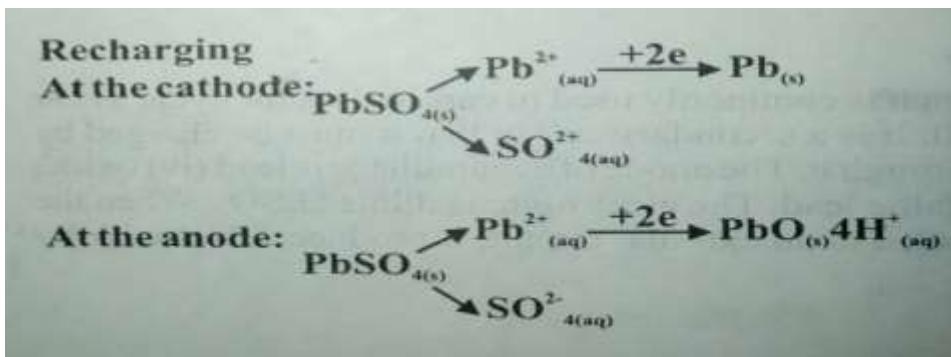


The Pb^{2+} then combines with the SO_4^{2-} of the electrolyte to form $\text{PbSO}_{4(s)}$



When both electrodes are covered with lead (ii) tetraoxosulphate (vi) deposits, the cell ceases to discharge any more current until it is recharged. Charging the battery removes the deposited PbSO_4 on the cathode and anode and restores the cathode to $\text{Pb}_{(s)}$ and anode to $\text{PbO}_{2(s)}$ by chemical reactions. The concentration and the density of the acid (electrolyte) also decrease due to absorption of H^+ and SO_4^{2-} during discharging and drops to about 1.15 g cm^{-3} . After charging, the acid density returns to 1.25 g cm^{-3} . The e.m.f. also returns from 1.8V to 2.2V after charging.

Recharging



Exercises:

1. State: (a) Faraday's first law of electrolysis
(b) Faraday's second law of electrolysis.
2. If 6.4g of Cu is deposited by a given quantity of electricity, how many grams of aluminium will be deposited at the same time?
(Cu = 63.5, Al = 27, 1 Faraday = 96500)
3. Calculate the mass of silver deposited when a current of 2.4A is passed through a solution of a silver salt for 70 minutes. (Ag = 108; 1 Faraday = 96500 C)
1. What are the uses of electrolysis?
2. What is the difference between an electrochemical cell and electrolytic cell in term of the charges on the cathode and anode?
3. Define standard electrode potential.
- 4(i) Name two types of Leclanche cell.
(ii) What is the difference between them?
- 5(a) What do you understand by:
(i) Primary cell? (ii) Secondary cell?
(b) Write an example of each.
(c) What is the difference between the two?

CHAPTER 15

ORGANIC CHEMISTRY

Organic chemistry is the study of the chemistry of all carbon compounds with the exception of the: metallic carbides, sulphides, oxides of carbon and trioxocarbonates (iv). These three classes of carbon compounds are regarded as inorganic compounds and are studied under the inorganic chemistry. Some common characteristics of organic compounds are as follows:

1. They are covalent compounds.
2. They have lower melting and boiling points than the inorganic compounds.
3. They decompose at low temperature.
4. Most organic compounds burn in excess supply of air to form carbon (iv) oxide and water. The reaction is exothermic.
5. They are insoluble in water.
6. Most of them are soluble in non-polar solvents e.g Benzene.



Fig. 15.1: Structure and valency of Carbon

15.1 SP³, SP² and SP Hybridization in Carbon Atom

The electronic configuration of carbon is $1s^2, 2s^2, 2px^1, 2py^1$ (Ground state configuration). In the excited form, one of the 2s electron is promoted to the vacant 2pz orbital of higher energy. i.e $1s^2, 2s^1, 2px^1, 2py^1, 2pz^1$ (Excited configuration).

This shows that the four valencies of carbon are not equivalent in every way. To overcome this anomaly one 2s and three 2p orbitals of carbon in the excited state undergo a process of mixing up and redistribution to form four equivalent new orbitals (four equivalent new valencies). This process is known as hybridization.

Hybridization can therefore be defined as the mixing up and redistribution of orbitals of nearly the same energy to form the same number of orbitals of equal energy. The new orbitals of equal energy formed is called the hybrid orbitals. Thus in an excited carbon, one 2s and three 2p orbitals are available for hybridization and depending on the actual number of orbitals taking part in it, the hybridization can be of three types.

Sp³ Hybridization is formed if one 2s and three 2p orbitals take part in hybridization. It results in the formation of four new equivalent hybrid orbitals known as Sp³ orbitals.

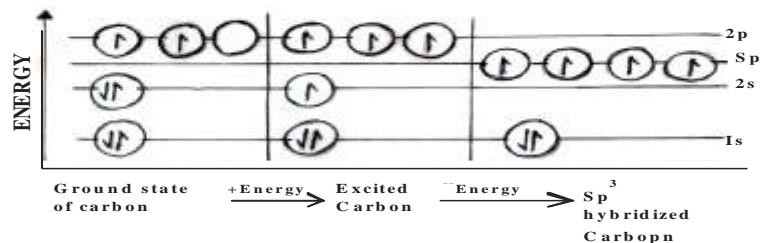


Fig. 15.2 Electronic configuration of Sp³ hybridized carbon

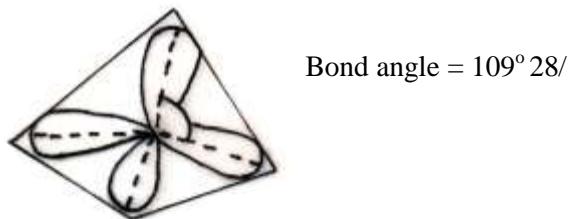


Fig 15.3 Shape of Sp³ Orbitals

Sp² Hybridization is formed if one 2s and two 2p orbitals take part in hybridization. It results in the formation of three new equivalent hybrid orbital known as Sp² orbitals.

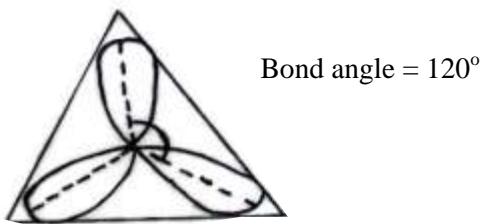


Fig 15.4 Shape of Sp² orbitals

Sp Hybridization is formed if the only one 2s and one 2p orbital takes part in hybridization. It results in the formation of two new equivalent hybrid orbitals known as Sp orbitals.

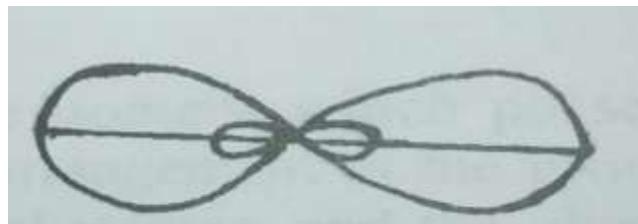


Fig 15.5 Shape of Sp orbitals

Homologous series: A homologous series is a family of organic compounds with the same molecular formula.

The homologous series include the alkanes, C_nH_{2n+2} , the alkanols, $C_nH_{2n+1}OH$ and carboxylic acids, $C_nH_{2n+1}COOH$. Some common characteristics of homologous series are:

1. All members conform to a general molecular formula.
2. Each member differs in molecular formula from the next neighbour by CH_2 and in its relative molecular mass by 14.
3. They exhibit similar chemical reaction (though varying in vigour).
4. They have general method of preparation e.g alkanes, CH_4 , C_2H_6 and C_3H_8 can be prepared using the same general method.
5. They exhibit a gradation of physical properties. For example, the alkanes, $CH_4-C_4H_{10}$ are gases, $C_5H_{12}-C_{17}H_{36}$ are liquids while the higher members are solids. The same is applicable to other physical properties (see Fig 8.5 below).

Properties Name	Molecular Formular (C_nH_{2n+2})	Relative Molecular Mass	Ordinary Physical State	Boiling Point (°C)	Melting Point (°C)
Methane	CH_4	16		-162	-183
Ethane	C_2H_6	30		-89	-172
Propane	C_3H_8	44		-42	-190
Butane	C_4H_{10}	58		-1	-135
Pentane	C_5H_{12}	72		36	-130
Hexane	C_6H_{14}	86		69	-94
Heptane	C_7H_{16}	100		98	-91
Octane	C_8H_{18}	114		125	-57
Nonane	C_9H_{20}	128		151	-54
Decane	$C_{10}H_{22}$	142	Liquides	174	-30

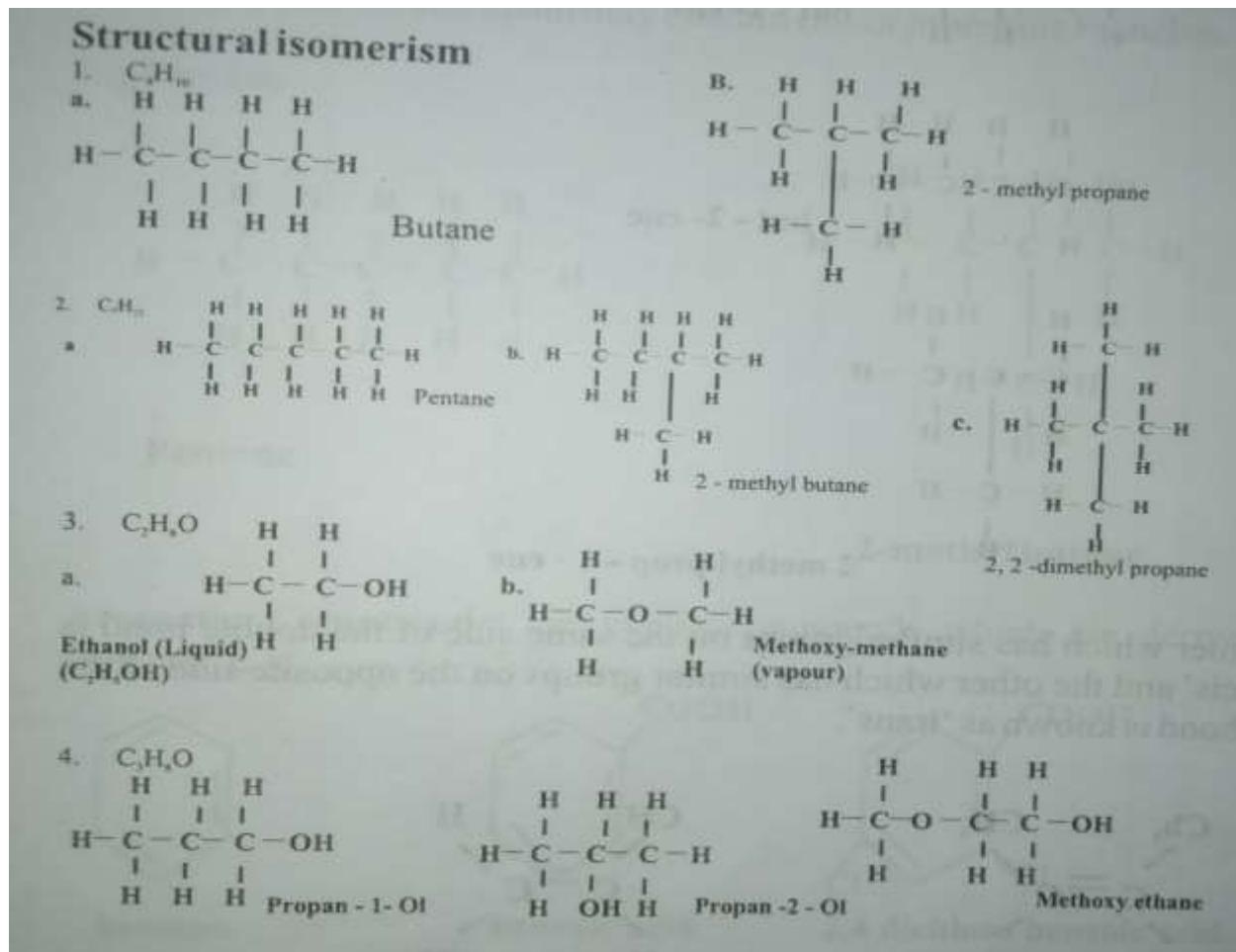
Fig 15.6

15.2 Isomerism: Isomerism is the occurrence of two or more compounds (known as isomers) with the same molecular formula but different molecular structure. This different molecular structure makes isomers with the same molecular formula to have different physical and chemical properties.

Two types of isomerism are:

1. Structural Isomerism

2. Stereo Isomerism - Geometrical and Optical Isomerism.

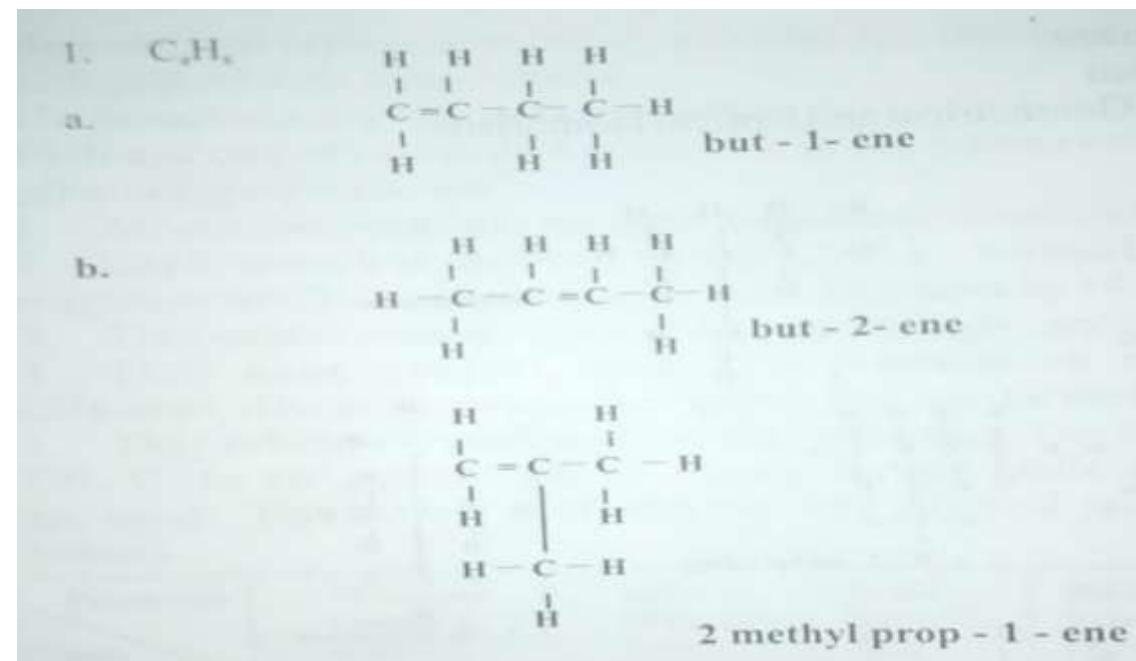


Stereo isomerism: Compounds having the same atoms or groups with a different spatial arrangement are called stereo isomers. Stereo isomerism is a phenomenon that occurs when compounds having the same atoms or groups posses different spatial arrangement.

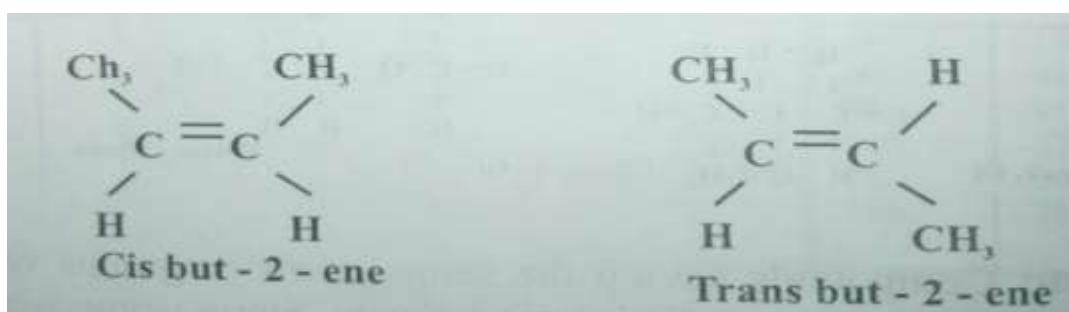
Two types of stereo isomerism are:

- (i) Geometrical Isomerism and
- (ii) Optical Isomerism.

Geometrical Isomerism: The isomers, which posses the same structural formula but differ in spatial arrangement of the groups around the double bond are known as geometrical isomers and the phenomenon is known as geometrical isomerism. Geometrical isomerism arises due to rigidity of double bond which does not allow free rotation.

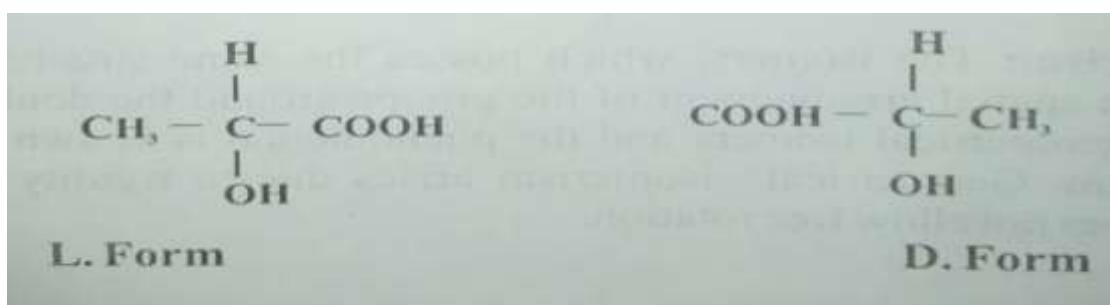


The isomer which has similar groups on the same side of the double bond is called ‘cis’ and the other which has similar groups on the opposite side of the double bond is known as ‘trans’.



Optical Isomerism: This occurs where two isomers are identical in their chemical and physical properties and have same structural formula except in the fact that they rotate the plane of plane-polarised light to an equal extent but in opposite direction e.g

2 - hydroxyl propanoic acid

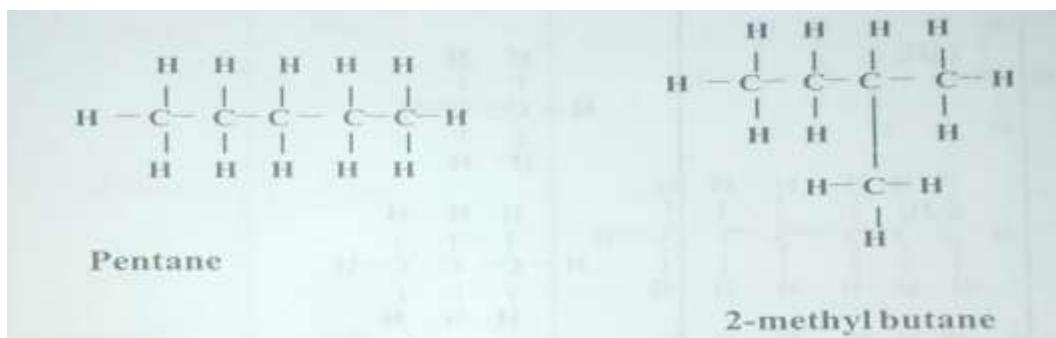


Classification of organic compounds

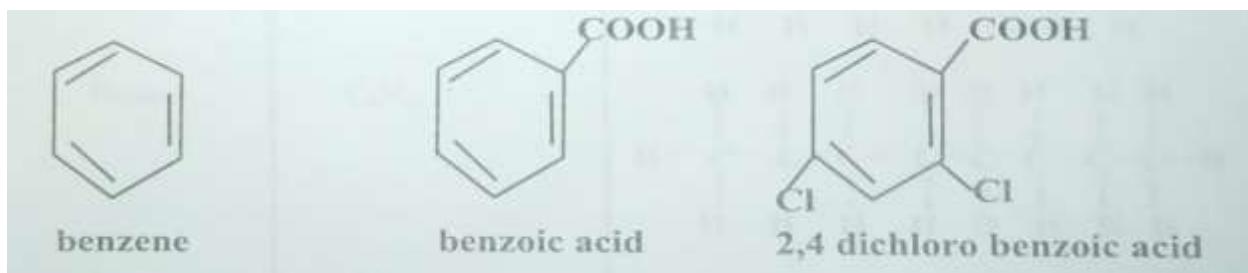
Organic compound is classified into aliphatic and aromatic compounds.

Aliphatic Compounds are molecules that posses chains of carbon atoms. Sometimes, the carbon chain may contain one or more side branches.

E.g Pentane



Aromatic Compounds: are cyclic compounds which are derived from benzene. E.g



15.3 Saturated Hydrocarbon-the Alkanes

Alkanes are the saturated hydrocarbon commonly known as the paraffins. Hydrocarbon which is a binary compound of carbon and hydrogen can be classified into: (1) The aliphatic and (2) The aromatic hydrocarbons. (See chapter 8 of this book). They can also be divided into the saturated and unsaturated hydrocarbons. Alkanes are saturated in the sense that they are chemically inert and have little affinity for other elements or compounds. Alkanes have the general molecular formula of C_nH_{2n+2} where 'n' is a whole number showing the number of carbon atoms present in a molecule. The first and second members are methane and ethane respectively. Organic compounds can be represented with molecular formulae and molecular structure. The order in which atoms are joined together in a molecule and the orientation of the atoms in space is known as the molecular structure.

ALKANE	MOLECULAR FORMULAR	MOLECULAR STRUCTURE
Methane	CH_4	$ \begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{H} \\ \\ \text{H} \end{array} $
Ethane	C_2H_6	$ \begin{array}{cc} \text{H} & \text{H} \\ & \\ \text{H}-\text{C} & -\text{C}-\text{H} \\ & \\ \text{H} & \text{H} \end{array} $
Propane	C_3H_8	$ \begin{array}{ccc} \text{H} & \text{H} & \text{H} \\ & & \\ \text{H}-\text{C} & -\text{C}-\text{C}-\text{H} \\ & & \\ \text{H} & \text{H} & \text{H} \end{array} $

ALKANE	MOLECULAR FORMULAR	MOLECULAR STRUCTURE
Butane	C ₄ H ₁₀	<pre> H H H H H - C - C - C - C - H H H H </pre>
Pentane	C ₅ H ₁₂	<pre> H H H H H H - C - C - C - C - C - H H H H </pre>
Hexane	C ₆ H ₁₄	<pre> H H H H H H H - C - C - C - C - C - C - H H H H </pre>
Heptane	C ₇ H ₁₆	<pre> H H H H H H H H - C - C - C - C - C - C - C - H H H H </pre>
Octane	C ₈ H ₁₈	<pre> H H H H H H H H H - C - C - C - C - C - C - C - C - H H H H </pre>
Nonane	C ₉ H ₂₀	<pre> H H H H H H H H H H - C - C - C - C - C - C - C - C - C - H H H H </pre>
Decane	C ₁₀ H ₂₂	<pre> H H H H H H H H H H H - C - C - C - C - C - C - C - C - C - C - H H H H </pre>

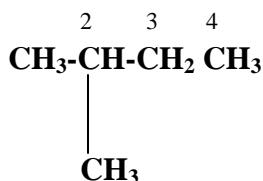
15.4 The Rules Observed in IUPAC Nomenclature of Alkanes

1. Alkanes have the suffix or end in — ane.
2. The name of the alkane corresponds to the longest chain of the carbon atom in the molecule
3. Assign number to each carbon atom of the chain in order to locate branches. Begin at the end of the chain that will result to assigning the smallest number to the carbon atoms holding the branches.

An alkyl group is an alkane in which one hydrogen is removed. They have the general molecular formula:

C_nH_{2n+1} . Some examples of alkyl group are:

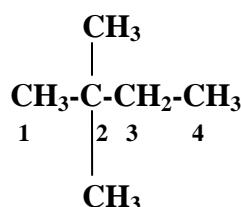
- CH_3 methyl group
- C_2H_5 ethyl group
- C_3H_7 Propyl group
- C_4H_9 Butyl group



2-methyl butane

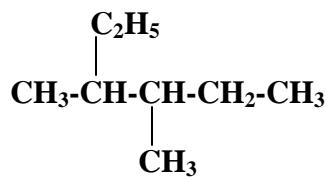
4. If two identical groups are attached at the same carbon atom, the same number are given to each of the groups.

Eg.



2,2 - dimethyl butane

5. If two or more different groups are attached to a chain, arrange the groups in order of complexity or in alphabetical order. e.g.



2 - Ethyl - 3-methyl pentane

6. The following substances are used as substituents

Flourine - Floro

Chlorine - Chloro

Bromine - Bromo

Iodine - Iodo

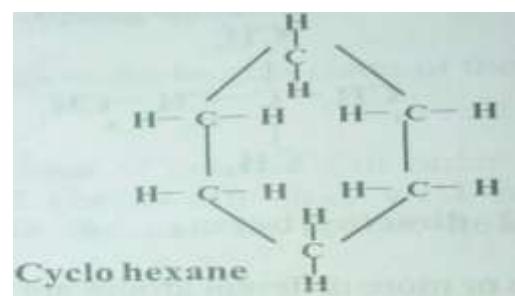
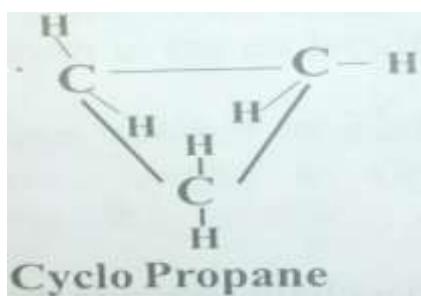
NO_2 - Nitro

NH_2 - Amino

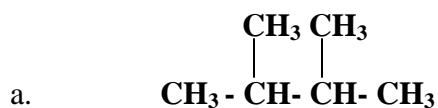
OH - Hydroxyl

COOH - Carboxyl

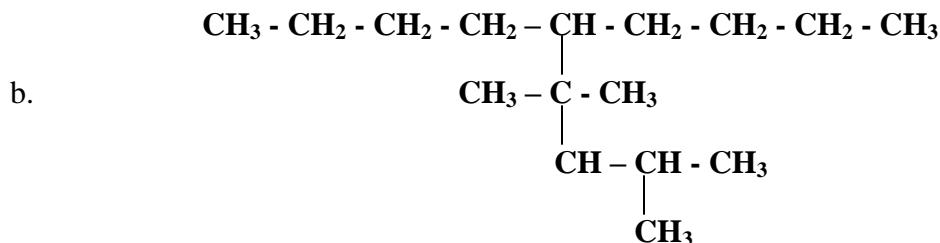
7. Cyclic compounds are named by adding the prefix cyclo to the name of the corresponding alkane. E.g Cyclopropane, cyclo hexane.



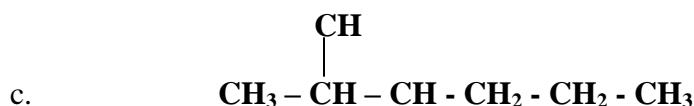
The IUPAC name of some alkanes are as follows:



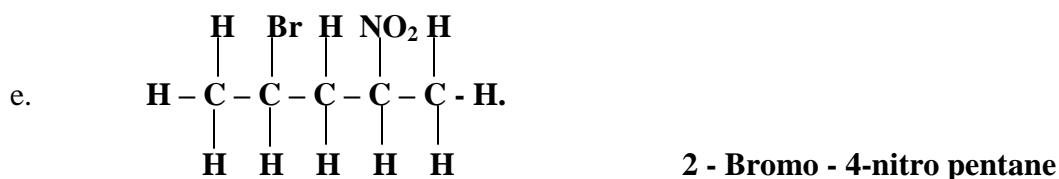
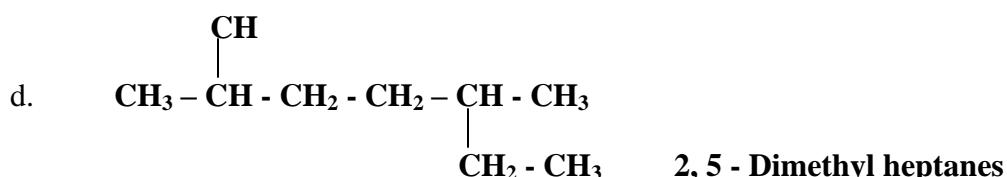
2, 3 - Dimethyl butane



5 - Butyl - 2,4,4 - trimethyl nonane



2 - methyl hexane

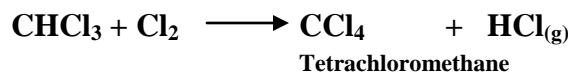
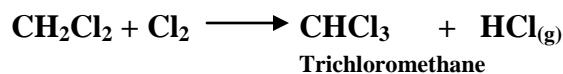
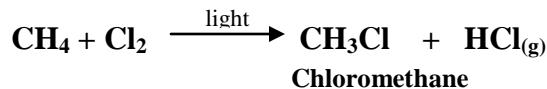


Chemical reactions of alkanes

Alkanes are chemically unreactive. They are the main components of petroleum (crude oil) and natural gas.

They undergo two main types of reaction at standard temperature and pressures (s.t.p.)

1. **Substitution reaction:** Alkanes undergo substitution reaction with halogen in the presence of light. Hydrogen is successively being replaced by the halogen. Eg.

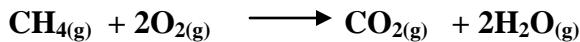


In the above reaction, the light acts as a catalyst in the splitting of chlorine molecules into free chlorine radicals.

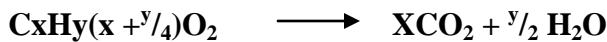


The Cl^+ reacts with CH_4 and displaces one of the hydrogen from it forming a halogenated product of alkane. Then the Cl^- combines with the displaced free hydrogen to form HCl gas.

2. **Combustion reaction:** Alkanes react with or burn in air to form carbon (iv) oxide and steam.



The general formula for complete combustion of hydrocarbons is as follows:



Uses of halogenated products of alkanes

1. They are used as solvent
2. Tetrachloromethane is used as antiseptic and as a solvent on removing grease stain from cloth.
3. Dichloromethane is used as solvent in dissolving paints.
4. Trichloromethane (chloroform) is used in anaesthesia.

General method of preparing alkanes

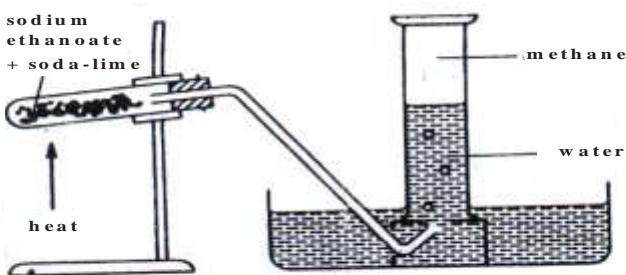
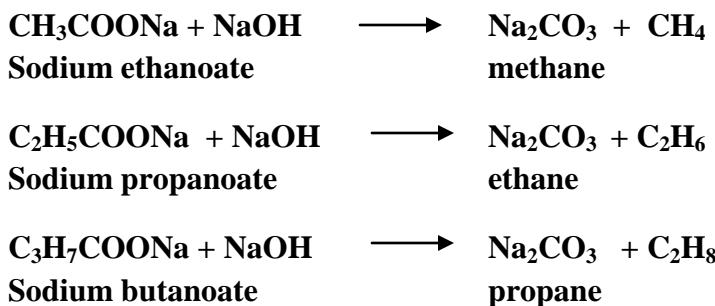


Fig 15.7 Preparation of methane

Alkanes are prepared in the laboratory by heating a mixture of an appropriate organic sodium salt and soda lime.

Sodalime is a mixture of sodium hydroxide and calcium hydroxide.

Sodalime is more preferably used to caustic soda. This is because of the fact that sodalime is not deliquescent and it does not easily attack glass.



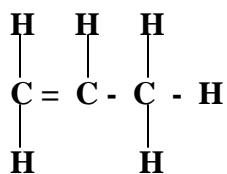
15.5 Unsaturated Hydrocarbon - The Alkenes

Alkenes are the unsaturated hydrocarbons that contain one or more carbon-carbon double bond per molecule. They belong to a homologous series with general molecular formula C_nH_{2n} . They are called the olefines and they have two hydrogen atoms less than the corresponding alkane. The first member with $n = 2$ is ethene. The molecular and structural formulae of the first five members are as follows:

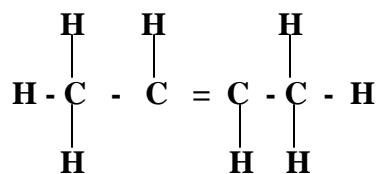
COMMON NAME	MOLECULAR FORMULAR	STRUCTURAL FORMULAR	IUPAC NAME
Ethene	C ₂ H ₄	$ \begin{array}{c} \text{H} & \text{H} \\ & \\ \text{C} = \text{C} \\ & \\ \text{H} & \text{H} \end{array} $	Ethene
Propane	C ₃ H ₈	$ \begin{array}{ccccc} \text{H} & & \text{H} & & \\ & & & & \\ \text{C} = \text{C} - \text{C} - \text{H} & & & & \\ & & & & \\ \text{H} & \text{H} & \text{H} & & \end{array} $	Prop-1-ene
Butene	C ₄ H ₈	$ \begin{array}{ccccc} \text{H} & \text{H} & \text{H} & & \\ & & & & \\ \text{C} = \text{C} - \text{C} - \text{C} - \text{H} & & & & \\ & & & & \\ \text{H} & \text{H} & \text{H} & & \end{array} $	But-1-ene
Pentene	C ₅ H ₁₀	$ \begin{array}{ccccc} \text{H} & \text{H} & \text{H} & \text{H} & \\ & & & & \\ \text{C} = \text{C} - \text{C} - \text{C} - \text{C} - \text{H} & & & & \\ & & & & \\ \text{H} & \text{H} & \text{H} & \text{H} & \end{array} $	Pent-1-ene
Hexene	C ₆ H ₁₂	$ \begin{array}{ccccc} \text{H} & & \text{H} & \text{H} & \\ & & & & \\ \text{H} - \text{C} - \text{C} = \text{C} - \text{C} - \text{C} - \text{C} - \text{H} & & & & \\ & & & & \\ \text{H} & \text{H} & \text{H} & \text{H} & \end{array} $	Hex-2-ene

Rules observed in IUPAC Nomenclature of Alkene

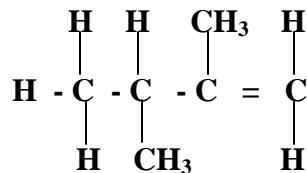
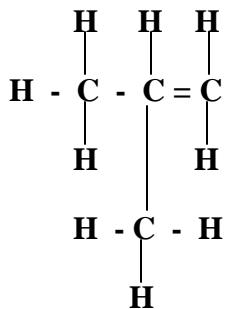
1. The class suffix of the alkenes or olefins is -ene.
2. The longest carbon chain bearing the double bond is chosen as the parent alkene.
3. The position of the double bond and side chains are indicated by numbers.
4. The lowest possible number is given to the carbon atom bearing the double bond. The number is written before the suffix -ene-.



Prop-I-ene

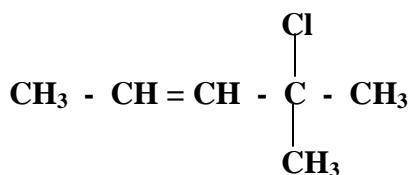


But-2-ene

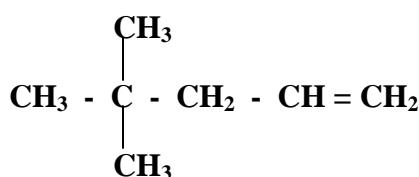


2,3-dimethyl but-I-ene

2-methyl prop-I-ene

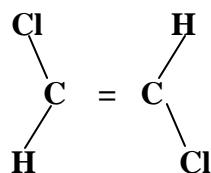


4-Chloro-4-methyl pent-2-ene

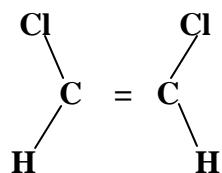


4,4-dimethyl pent-I-ene

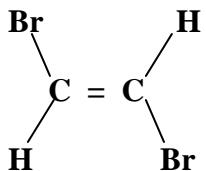
The IUPAC name of some geometric isomers of alkene are as follows:



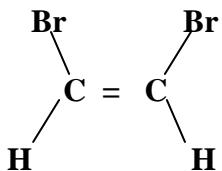
Trans-i, 2-dichloro



Cis-1, 2-dichloroethene



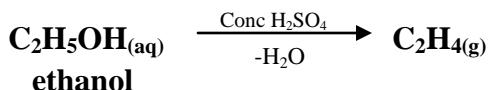
Trans-1, 2-dibromoethene



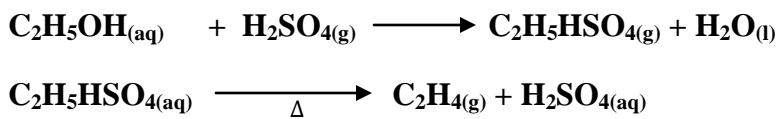
Cis-1, 2-dibromoethene

Preparation of ethene

Ethene is produced in the laboratory by the dehydration of ethanol by hot conc. H_2SO_4 .



In practice, two steps are involved in this reaction. On mixing, $\text{C}_2\text{H}_5\text{OH}$ reacts with the drying agent (Conc. H_2SO_4) to form an intermediate salt known as: ethyl hydrogentetraoxosulphate(vi). The latter, readily decompose on heating to form C_2H_4 and H_2SO_4 .



NOTE: The ethene gas produced is passed over NaOH or KOH to remove some gaseous impurities like CO_2 and SO_2 . These acidic oxides may be formed by certain oxidation by the hot conc H_2SO_4 .

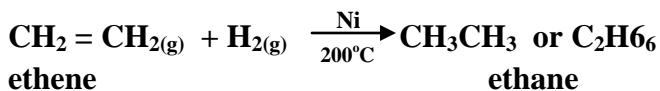
Industrially, ethene is prepared by cracking the gas oil.



Reactions of alkene (ethene)

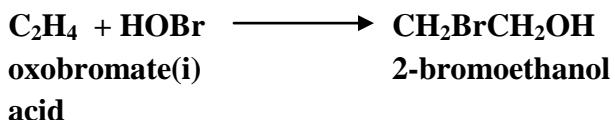
Alkenes are more reactive than alkanes due to the presence of double bond. Typical reactions undergone by alkenes are addition and polymerization reactions. Ethene will be used to show these reactions.

1. **Addition reaction:** All unsaturated hydrocarbon undergo addition reaction. It is the type of reaction where two substances combine to form a single substance. Examples of addition reactions are as follows:
 - a. **Addition of hydrogen:** This type of reaction is also called hydrogenation reaction. Hydrogen adds to ethene to form ethane at a very high temperature and in the presence of catalysts like finely divided nickel, palladium and platinum.

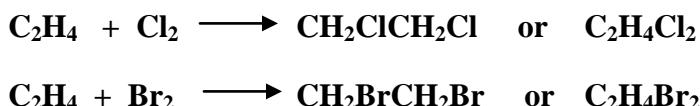


b. **Addition of bromine and chlorine water:**

Ethene reacts with the above named two substances to form 2-bromoethanol and 2-chloroethanol respectively.



- c. Addition of halogens: Ethene reacts with chlorine or bromine to form 1,2-dichloroethane or 1,2-dibromoethane respectively.



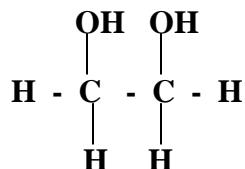
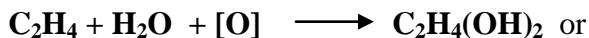
Ethene decolourises the reddish brown colour of bromine and greenish yellow colour of chlorine. This reaction distinguishes the saturated gaseous alkane from under from unsaturated alkenes. Ethene destroy the reddish brown colour of bromine vapour.

d. Addition of potassium tetraoxomanganate (vii):

Ethene reacts rapidly if shaken with potassium tetraoxomanganate (vii) solution.

- i. If KMnO_4 solution is acidic, ethene decolourizes it from purple to colourless solution of a manganese (ii) salt.
 - ii. If KMnO_4 solution is alkaline, ethene changes it from purple to green solution of potassium tetraoxomanganate (vi).

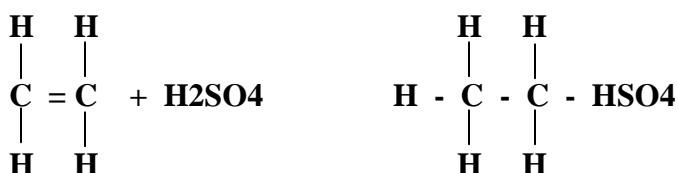
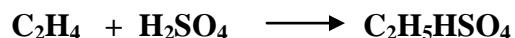
In both reactions, ethene is converted to ethane -1, 2-diol.



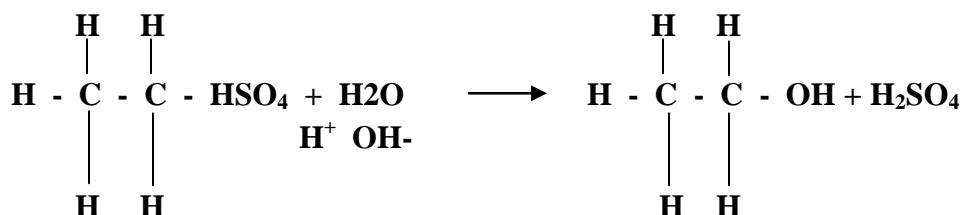
This reaction again distinguishes ethene (unsaturated hydrocarbons) from ethane (or the saturated hydrocarbons).

e. **Addition of Conc. H₂SO₄:**

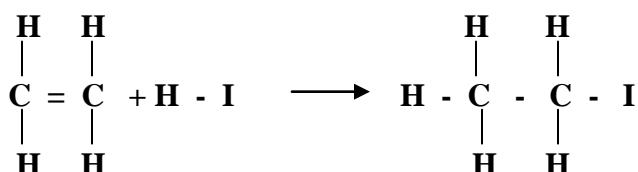
Ethene reacts with Conc. H₂SO₄ to form ethyl hydrogen tetraoxosulphate (vi) which on hydrolysis by boiling it with H₂O yields ethanol and H₂SO₄ is recovered.



Ethylhydrogen tetraoxosulphate(vi)

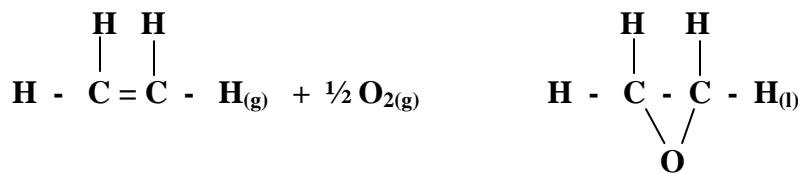


f. **Addition of hydrogen halide:** Ethene reacts with hydrogen halides (vapour) like HI, HCl and HBr at room temperature to form iodoethane, chloroethane and bromoethane respectively.



iodoethane

- g. **Addition of Oxygen:** Alkenes react with air or oxygen at temperature of about 250 and in the presence of silver catalyst to form an epoxide.



Epoxyethane

2. Polymerization

Polymerization is a process whereby two or more simple molecules known as monomers link together to form a much longer and complex molecule called the polymer.



Where n is the number of monomers and Y is the monomer

If ethene is pressurised to about 100 atmosphere with a trace of oxygen and heated to start the reaction, it polymerises to polyethene.



where 'n' is about 800.

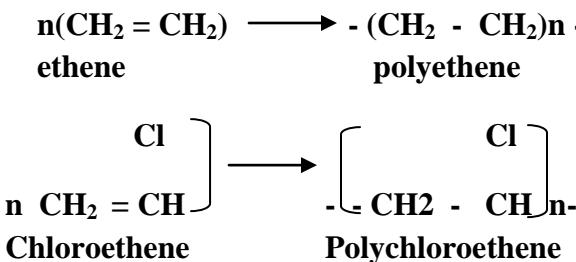
Like alkanes, polyethene is inert and resistant to common chemical reactions. When hot, it can be moulded into a great variety of domestic and scientific articles like funnels, buckets, tables, stools, wash bottles, bags, spoons etc.

Types of polymerization: Three types of polymerization that will be discussed here are:

- i. Addition polymerization
- ii. Condensation polymerization
- iii. Cross linking polymerization

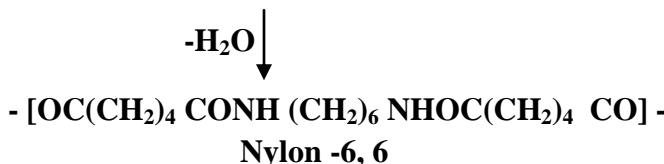
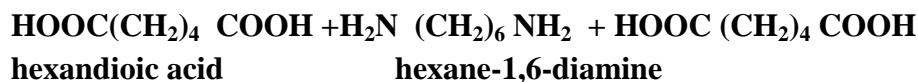
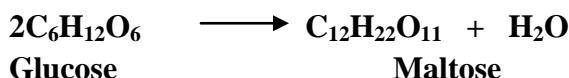
Addition polymerization: Addition polymerization is a process of polymerization where the same kind of monomers combine together to form a large complex addition product without any loss or gain of material.

Polyethene is an addition polymer. It has only one type of monomer molecule and one polymer product e.g.

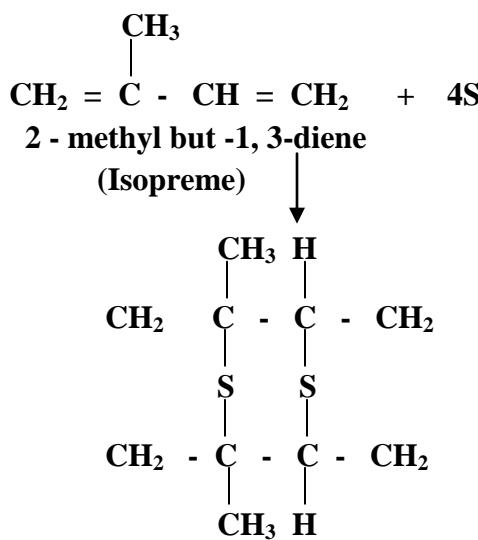


ii. Condensation Polymerization:

Condensation polymerization is the polymerization where a new long chain polymer is formed from the combination of two or more monomers with the elimination of a simple molecules such as ammonia, water or hydrogen chloride. E.g.



Cross linking polymerization: This is the type of polymerization where the added atom forms bridges between the adjacent chain molecules of the polymer at the double bond position. An example of this is seen in the vulcanization of rubber, where the added sulphur forms bridges of sulphur atoms between the adjacent chain molecules of the polymer (rubber) at the double bond positions.



Natural rubber:

Rubber is prepared from the milky liquid (latex) obtained by making incision into the back of rubber tree. Natural rubber is a polymer of 2-methyl but-1, 3-diene (isoprene)



The emulsion is coagulated by heating gently and the addition of ethanoic acid to form crude rubber. The elasticity of pure rubber is retained over a very limited temperature range. On cooling to 0°C , it becomes brittle while gentle heat render it sticky. The temperature range over which rubber retains its elasticity and durability are enhanced by vulcanization. Vulcanization is the heating of raw rubber with about 3% by mass of sulphur. The added sulphur forms a cross linked polymer with the adjacent polymers (rubber). See the cross linking polymerization). Cross-linking makes rubber to be more elastic, less soft and less soluble in organic solvents.

Uses of Ethene:

1. It is used in the making of polyethene which can be used in the manufacturing of a variety of products like toys, plastic buckets, tables, chairs etc.
2. It is used in the making of detergents.
3. It is used in hastening the ripening of fruits.

15.6 Unsaturated Hydrocarbon - The Alkynes

Alkynes are another group of unsaturated hydrocarbons that contains a carbon-carbon triple bond per molecule. They are members of a homologous series with general molecular formula $\text{C}_n\text{H}_{2n-2}$.

They have four and two hydrogens less than their corresponding alkanes and alkenes respectively. Ethyne (Acetylene) is the first member where 'n' is equal to 2. It is also the only important member. The molecular and structural formula of the first four members are as follows:

MOLECULAR FORMULA	STRUCTURAL FORMULA	IUPAC NOMENCLATURE
C ₂ H ₂	H - C ≡ C - H	Ethyne
C ₃ H ₄	H - C ≡ C - CH ₃	Prop-1-yne
C ₄ H ₆	CH ₃ - C ≡ C - CH ₃	But-2-yne
C ₅ H ₈	CH ≡ C - CH ₂ - CH ₂ - CH ₃	Pen-1-yne

PREPARATION OF ALKYNE (Ethyne)



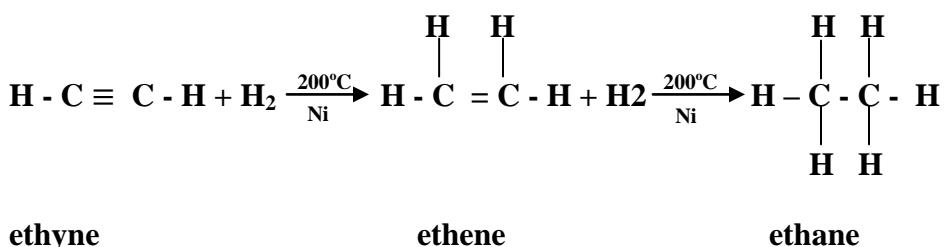
Ethyne is produced when water is added to calcium dicarbide. The reaction is highly exothermic. It is therefore performed on a heap of sand inside the flask. Ethyne is stored in industry in steel cylinder in a solution of propanone at about 12 atmospheric pressure. An impurity like phosphine, PH₃, is removed from ethyne by passing it through a solution of CuSO₄.

Physical and chemical properties of ethyne

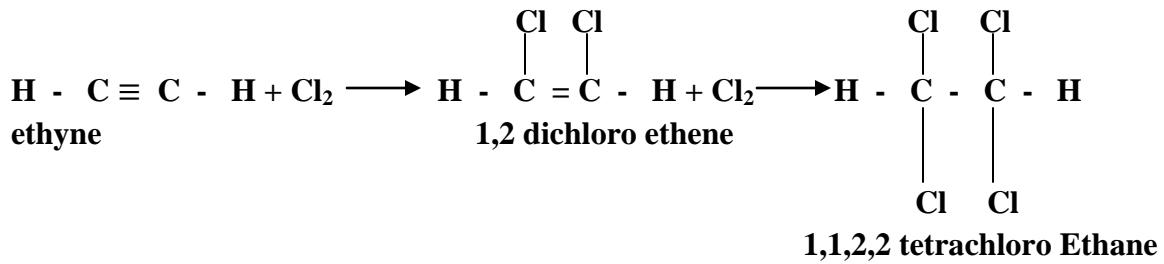
1. It is a colourless gas with a sweet odour when pure.
2. It is slightly soluble in water and lighter than air.

Ethyne is highly unsaturated. It is therefore very unstable and very reactive. It undergoes these type of reactions: (i) Addition reaction (ii) Combustion reaction and (iii) polymerization reaction.

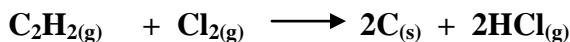
- i. **Addition Reactions:** Addition reactions with four univalent atoms occurs in two stages.
 - a. Addition of hydrogen:



- b. **Addition of halogen:** Halogens adds to ethyne in the presence of a metallic halide to form tetrahalogenated compounds.

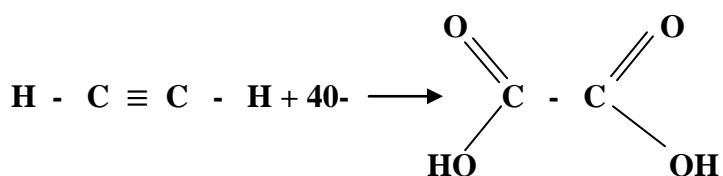


If bromine reacts with ethyne under the same condition tetrabromo ethane is formed. Like the alkenes, it decolourizes the reddish brown colour of bromine vapour. If pure ethyne and pure chlorine reacts in the absence of a catalyst, carbon and hydrogen chloride is formed



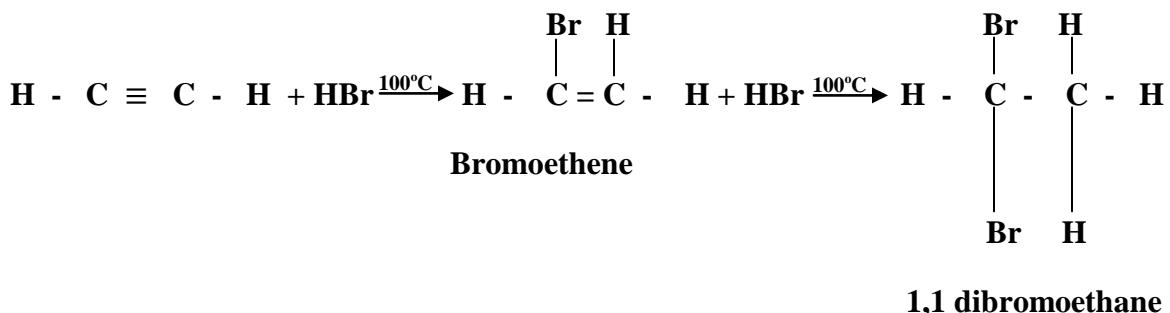
c. **Addition of acidified potassium tetraoxomanganate (vii) KMnO_4 :**

At room temperature, ethyne decolourizes acidified KMnO_4 solution from purple to colourless. If the solution is alkaline, ethyne turns it green and in each case, ethyne is oxidized to ethandioic acid.

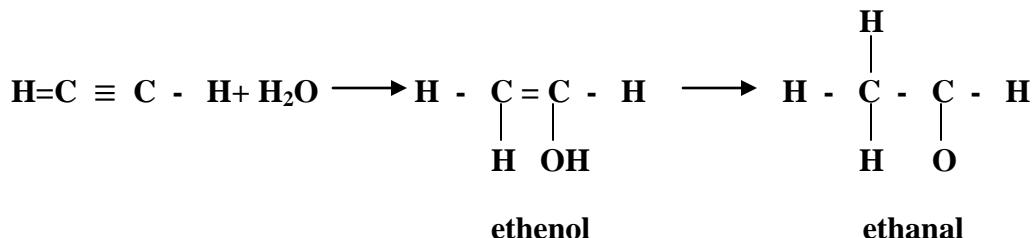


d. Addition of halides (gases)

With the exception of hydrogen chloride which is slow in reaction, hydrogen bromide and hydrogen iodide combine rapidly with ethyne. Reaction with hydrogen iodide occurs at room temperature.

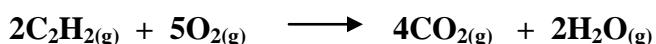


e. **Addition of Water:** Ethyne adds to water if passed through dilute tetraoxosulphate (vi) acid containing mercury(ii) tetraoxosulphate(vi) as a catalyst at about 65°C . An aldehyde - ethanal is formed.



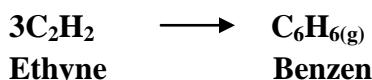
Ethenol cannot be isolated because it is too unstable. It rearranges to form ethanal.

- ii. **Combustion reaction:** Ethyne burns in air with sooty flame. This is due to high proportion of carbon in ethyne which is about 90% by mass. In an unlimited supply of air carbon (iv) oxide and water are formed.

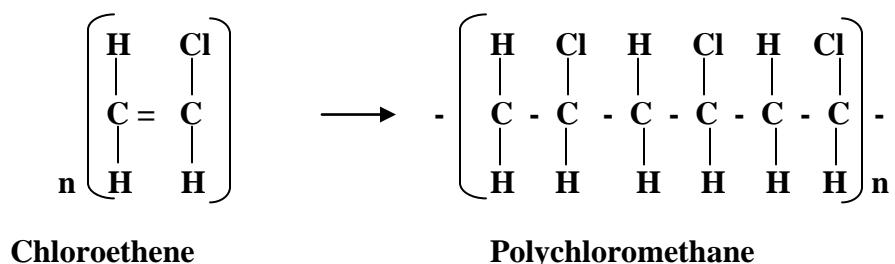


- ### iii. Polymerization reaction:

- a. Ethyne polymerizes to benzene - a cyclic aromatic hydrocarbon if passed through a hot tube.



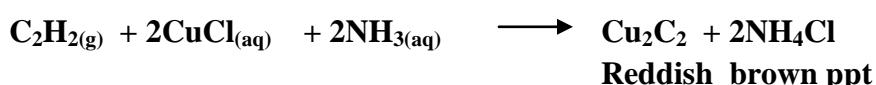
- b. Chloroethene polymerises to polychloroethane. Polychloroethane is composed of very long chain of carbon atom. Chlorine is attached to the alternate carbon atoms.



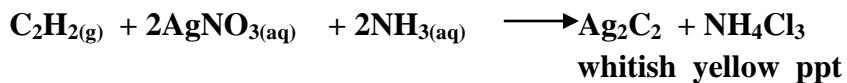
Distinguishing tests for alkenes and alkynes

Unlike ethene, ethyne can form its metallic derivatives.

- a. If ethyne is passed into an ammoniacal solution of copper (i) chloride at room temperature, a reddish brown precipitate of copper (i) dicarbide is formed.



- b. If ethyne is passed into an ammoniacal solution of silver trioxonitrate (v) at room temperature, a whitish yellow ppt of silver dicarbide is formed

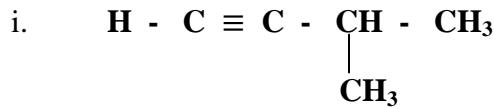


Uses of ethyne

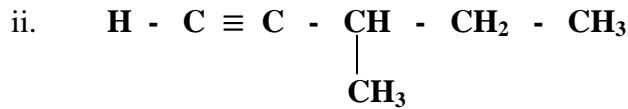
- It is used in making chloroethene which polymerised to polychloroethane. Polychloroethane is used in the manufacture of many domestic and household products such as plastic trays, spoons, tables, stools, chairs and toys.
- It is used in producing ethene. Ethene polymerizes to polythene. Polythene is used in the manufacture of polythene bags and other products.
- A flame of ethyne burning in pure oxygen is called oxyethyne flame. It is very hot and has a temperature of about 3000°C . It is used in welding and cutting of metals.

12. Rules observed in iupac nomenclature of alkynes

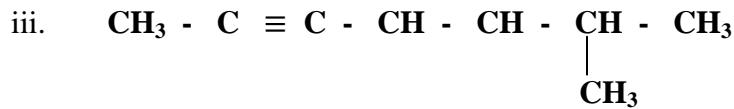
- The class suffix of the alkynes is -yne.
- The longest carbon chain bearing the triple bond is chosen as the parent alkyne.
- The position of the triple bond and side chains are indicated by numbers.
- The lowest possible number is given to the carbon atom bearing the triple bond. The number is written before the suffix-yne-



3 - methyl but -1-yne

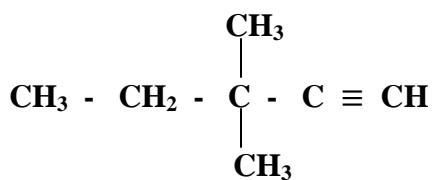


3-methyl pent-1-yne



6-methyl hept -2-yne

iv.



3,3 dimethyl pent-1-yne.

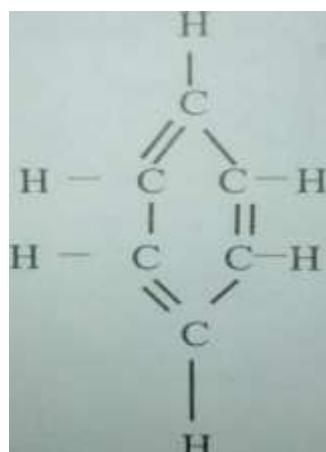
15.7 Aromatic Hydrocarbons - Benzene

Benzene was discovered by Faraday in 1825. The molecular formula of benzene is C_6H_6 . It is a liquid. Benzene has eight (8) hydrogen atoms less when compared with its corresponding alkane - hexane, C_6H_{14} . This shows that benzene must be an unsaturated hydrocarbon. Benzene like unsaturated hydrocarbon such as alkenes and alkynes undergo addition reactions with hydrogen, halogens and ozone. These addition reactions of benzene indicate that its structure consists of three double bonds.

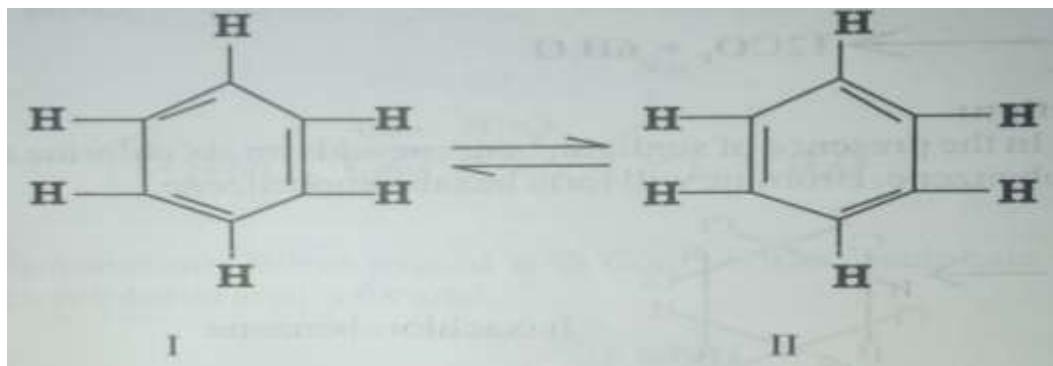
However, another set of empirical evidences cast doubts on the type of double bonds in benzene. Benzene is assumed to posses the above listed evidence including:

1. Unlike the alkene and the alkynes benzene does not add on to acidic halides such as hydrochloric acid, HCl , hydrogen bromide; HBr .
2. Benzen neither decolourizes pink or purple colour of acidified KMnO_4 nor turns the alkaline solution green.
3. Benzen like saturated hydrocarbon undergoes substitution reactions with halogens.

These facts clearly indicate that there is a difference in the nature of the double bonds in benzen as against the nature of the double bonds in alkenes. This problem motivated a lot of chemists who investigated more on the nature of benzen. It was in 1865 that Kekule first suggested that benzen must have a cyclic ring structure with three single and double bonds alternating among its carbon atoms.



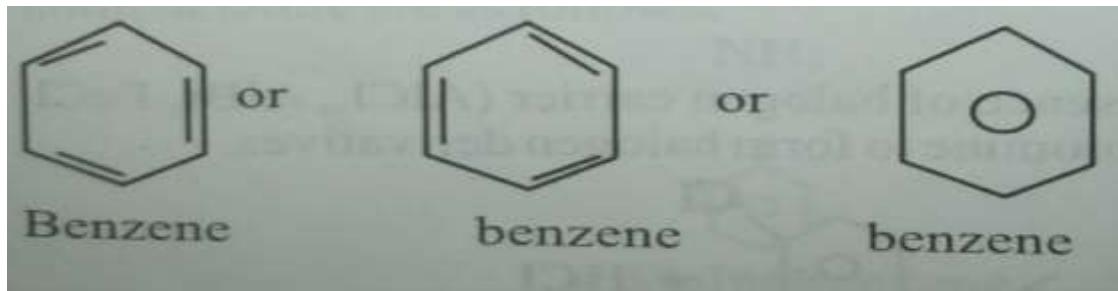
Today, we now say that benzene structure does not strictly consist of alternating single and double bonds. The suggestion is that the benzene structure is a resonance hybrid. That is, the structure is one that is between two alternate structures.



The carbon-carbon bonds in benzene molecule are identical and the structures could be represented as



One can therefore use any of these three structures to represent benzene molecule.



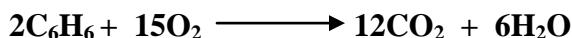
Physical properties of benzene

1. It is a colourless liquid with a characteristic odour.
2. Its boiling point is 80.1°C .
3. It is lighter than water and is insoluble in it.
4. It is a good solvent for fats, resins, sulphur, phosphorus, iodine etc.
5. It is miscible with organic solvents such as alcohol and ether.
6. Benzene vapour is inflammable and toxic in nature.

Chemical properties of benzene

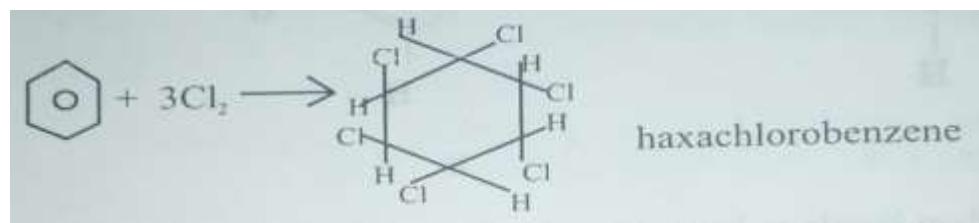
Benzene undergoes both addition, combustion and substitution reactions.

1. **Combustion reaction:** Benzene burns in air with a sooty flame forming carbon (iv) oxide and water vapour.

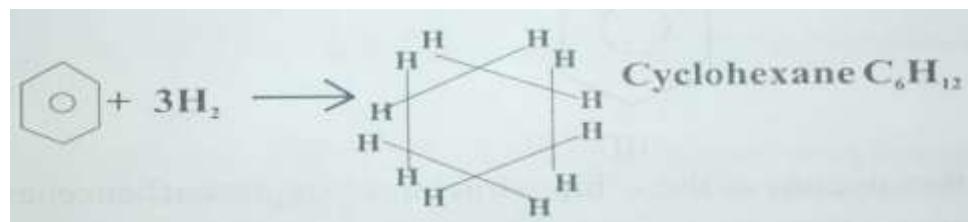


2. **Addition reaction:**

- a. **Halogenation:** In the presence of sunlight, benzene adds on six chlorine atoms to form hexachlorobenzene. Bromine will form hexabromobenzene.

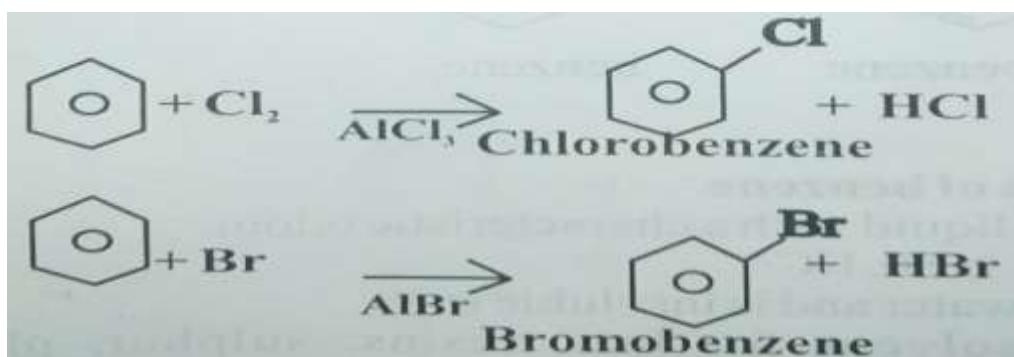


- b. **Hydrogenation:** When a mixture of hydrogen and benzene vapour is heated up to a temperature of 100°C over finely divided platinum or Nickel, cyclohexane is formed.



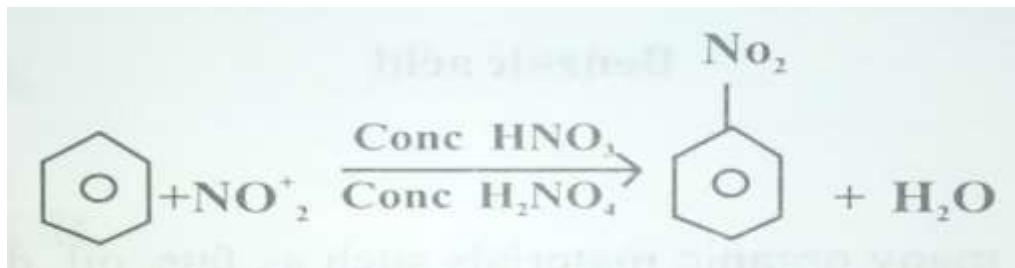
3. **Substitution reaction:**

- a. Halogenation: In dark and in presence of halogen carrier (AlCl_3 , AlBr , FeCl_3 etc) benzene reacts with chlorine or bromine to form halogen derivatives.

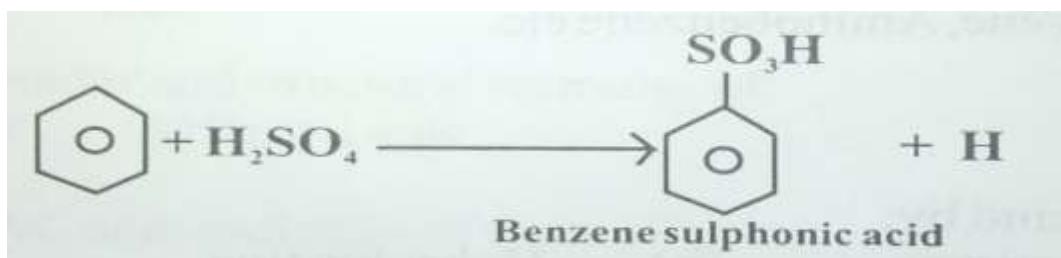


Saturated aliphatic hydrocarbons undergo substitution reaction in the presence of sunlight.

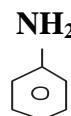
- b. **Nitration:** When benzene is heated with a mixture of conc. Trioxonitrate (v) acid and tetraoxosulphate(vi) acid, nitrobenzene is obtained and the reaction is called nitration.



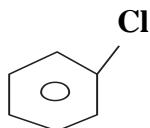
- c. Sulphonation: When treated with Conc. tetraoxosulphate (iv) acid, benzen sulphonic acid is formed.



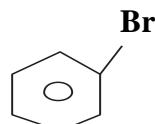
Some of the substituted benzene compounds and their IUPAC nomenclature are as follows:



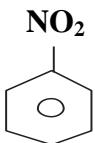
Aminobenzene



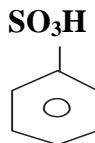
Chlorobenzene



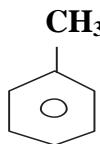
Bromobenzene



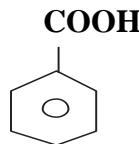
Nitrobenzene



Benzene sulphonic acid



Methylbenzene



Benzoic acid

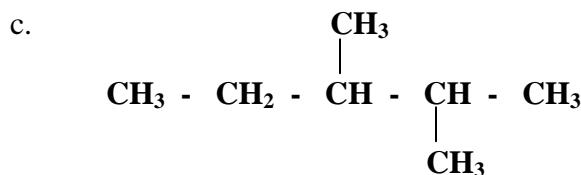
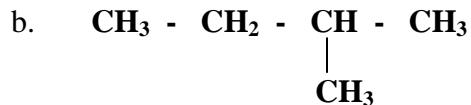
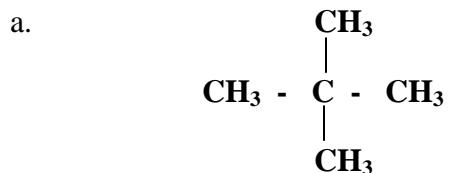
Uses of benzene

1. It is used as solvent for many organic materials such as fats, oil, drugs and dyes. It is also a solvent for some inorganic substances such as iodine, sulphur and phosphorus.
2. It is used in the manufacture of important chemicals like phenol, methylbenzene, Nitrobenzene, Aminobenzene etc.

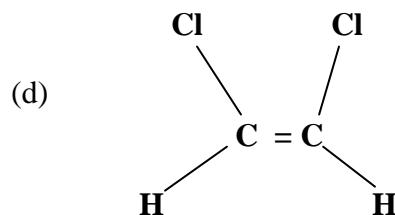
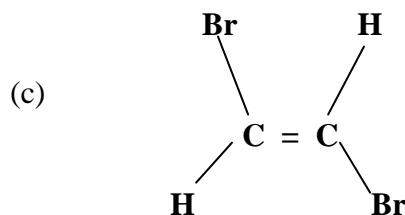
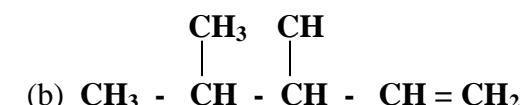
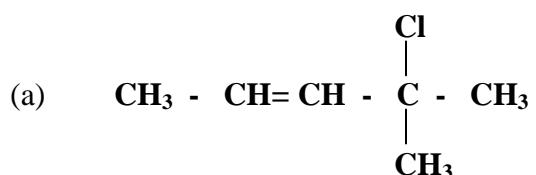
Exercises

1. What do you understand by:
 - (a) Organic chemistry
 - (b) Hybridization
 - (c) Homologous series
 - (d) Isomerism
2. Differentiate among these:
 - (a) sp³ hybridization
 - (b) sp² hybridization
 - (c) sp hybridization
3. Draw the shapes of :
 - (a) sp³ orbitals
 - (b) sp orbitals
4. Write two:
 - (a) Structural isomers

- (b) Geometrical isomers
 (c) Optical Isomer
5. Write the molecular and structural formular of (a) Butane (b) Heptane.
6. Using chemical equations only, show two chemical reactions undergone by the alkanes.
7. Write the IUPAC nomenclature of these alkanes.



8. Write the molecular and structural formular of:
 (a) But-1-ene (b) Pent-1-ene
9. Write the IUPAC nomenclature of these alkenes:

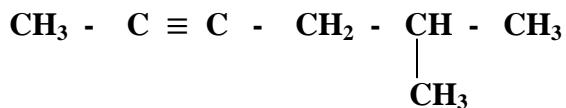


10. Use one equation only to show hydrogenation of ethene.
11. Define these terms:
 (a) Polymerization (b) Addition polymerization
 (c) Condensation polymerization

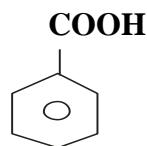
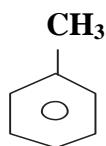
12. Differentiate between Natural rubber and vulcanized rubber.
13. Use a balanced equation only to show how ethyne can be prepared.
14. It was said that ethyne undergoes
 - (i) addition reaction
 - (ii) combustion reaction and
 - (iii) polymerization reaction

Use an equation only to represent each of these reactions.

15. How can you distinguish ethyne from ethene?
16. Draw the structural formulae of these alkynes
 - a. 2,5-dimethyl - hex-3-yne
 - b. 3,4 -dimethyl pent-1-yne
17. Write the IUPAC nomenclature of this alkyne.



18. Why are the double bonds in benzene said to be different in nature from that of the alkenes?
19. Draw three structures that is used to represent benzene ring.
20. Write two physical properties of benzene.
21. Write two chemical properties of benzene.
22. Write the IUPAC name of these benzene derivatives:



CHAPTER 16

AKANOLS

16.1 Alkanols (Alcohols)

Alkanols form a homologous series with general molecular formula C_nH_{2n+1}OH, where n = 1, 2, 3 etc. C_nH_{2n+1} is the general molecular formula for the alkyl groups usually represented with 'R'. Akanol can therefore be represented with the general molecular formula ROH. Examples of alkanols are:

Methanol

- CH₃OH

Ethanol

- C₂H₅OH or CH₃CH₂OH

Propanol

- C₃H₇OH or CH₃CH₂CH₂OH

Butanol

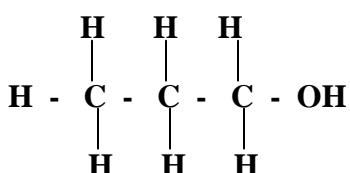
- C₄H₉OH or CH₃CH₂CH₂CH₂OH or CH₃(CH₂)₃ OH

Pentanol

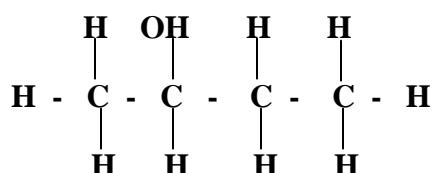
- C₅H₁₁OH or CH₃CH₂CH₂CH₂CH₂OH or CH₃(CH₂)₄OH

IUPAC Nomenclature of alkanols

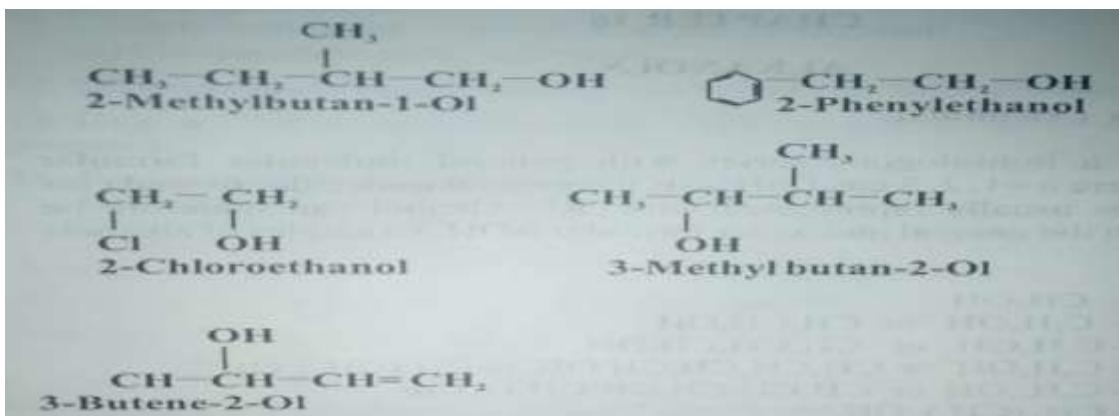
1. The suffix is Ol. Alkanols end inOl.
2. The root name is the longest continuous carbon chain bearing the OH group.
3. The carbon atoms in the longest continuous chain are numbered in such a way that the carbon atom bearing the OH group has the lowest possible number. This number is placed before the suffix e.g. -1-Ol.



Propan-1-Ol



Butan-2-Ol



Functional groups:

Functional group is an atom or group of atoms (radical) common to a homologous series which determines the chemical properties of the series. **Alkyl groups** are all groups derived from the alkanes by the loss of a hydrogen atom (see part III, Chapter 10.2). Most organic compounds is made up of two parts.

- i. The alkyl groups and
- ii. The functional group.

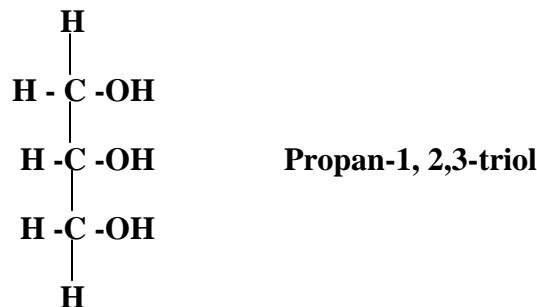
Since the alkyl group is inert, the chemistry of most organic compounds is determined by the functional group. The alkyl groups help in determining the physical properties.

Homologous series Haloalkanes	Formular $\text{R}-\text{x}$ (where $\text{X}=\text{F}, \text{Cl}, \text{Br}$ or I)	Functional group -X
Ethers	$\text{R}-\text{O}-\text{R}$	-O-R
Ketones	$\begin{array}{c} \text{R} \\ \\ \text{C}=\text{O} \\ \\ \text{R} \end{array}$	>C=O
Alkanols	$\text{R}-\text{OH}$	-OH
Aldehydes	$\begin{array}{c} \text{H} \\ \\ \text{C}=\text{O} \\ \\ \text{R} \end{array}$	H>C=O
Carboxylic acids	$\text{R}-\text{C}(=\text{O})-\text{OH}$	-C(=O)-OH
Esters (Alkanoates)	$\text{R}-\text{C}(=\text{O})-\text{O}-\text{R}$	-C(=O)-O-R
Primary amines	$\text{R}-\text{NH}_2$	-NH ₂

Types of alkanols: Alkanols can be classified according to the number of hydroxyl (OH) group present per molecule.

- Monohydric alkanols contain one hydroxyl group per molecule e.g. C₂H₅OH
- Dihydric alkanols contain two hydroxyl group per molecule e.g. Ethan-1,2-diol

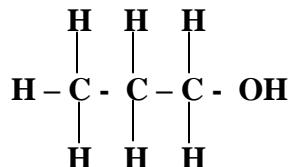
$$\begin{array}{c} \text{CH}_2 - \text{CH}_2 \\ | \quad | \\ \text{OH} \quad \text{OH} \end{array}$$
 or C₂H₄(OH)₂
- Trihydric alkanols contain three hydroxyl group per molecule e.g. Propan-1, 2, 3-triol



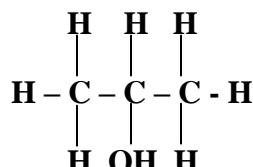
Classification of monohydric alkanols

Monohydric alcohols or alkanols are classified into: Primary alkanol, secondary alkanol and tertiary alkanol.

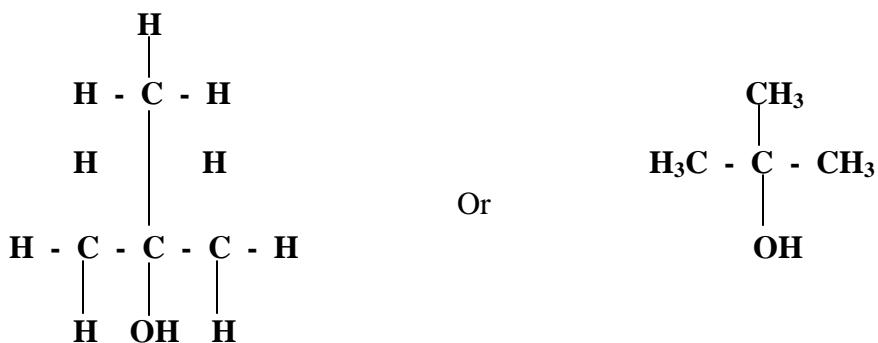
In primary alcohols, the hydroxyl group is attached to the primary carbon atom. A primary carbon atom is the carbon atom that is joined to only one other carbon atom.



Secondary alkanol is one in which the hydroxyl group is attached to a secondary carbon atom. A secondary carbon atom is the one that is joined to two other carbon atoms.



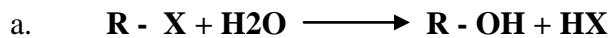
A tertiary alcohol is one in which the hydroxyl group is attached to a tertiary carbon atom. A tertiary carbon atom is one that is joined to three other carbon atoms by (CH₃)₃COH



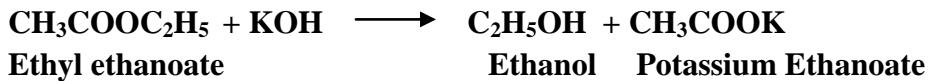
2-Methyl propan-2-Ol

16.2 General Methods of Preparation of Alkanols

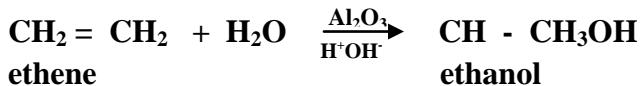
1. By the hydrolysis of alkyl halides with aqueous alkalis



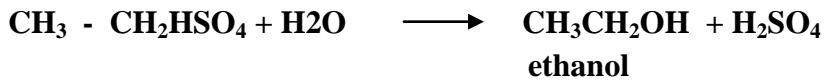
- c. By the hydrolysis of ester with aqueous alkali



- d. By the hydration of alkene in the presence of catalyst like Aluminium oxide or magnesium phosphate

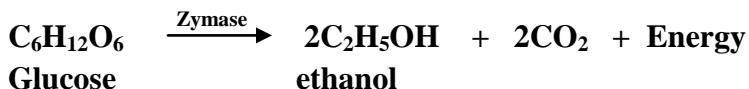


- e. By the sulphonation of alkene, followed by the hydrolysis of the product formed



Preparation of ethanol by fermentation

Fermentation is the slow catalytic decomposition of complex organic compounds into simpler substances (compounds) through the agency of enzymes. Enzymes are organic catalysts secreted in a living cells. An essential reaction that occurs during fermentation is the conversion of glucose into alcohol by an enzyme called zymase which is present in the yeast.



In practice, glucose is not always used as the starting compound. Starch or cellulose may be used. In this case, they have to be converted to simple sugar (glucose or fructose) before the fermentation can occur. Some examples of fermentation process we come across in our daily life are souring of milk, palm wine and putrefaction of meat or food.

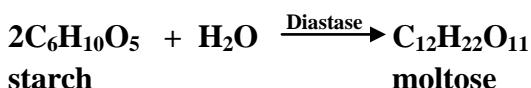
Local gin is produced from palm wine. Fresh palm wine contains sugar solution and yeast. After standing the palm wine for 3-4 days the yeast ferments the sugar to ethanol and carbon dioxide. The resultant solution is then distilled to obtain the local gin.

Sugar solution and yeast. After standing the palm wine for few days, the yeast, ferments the sugar to ethanol and carbon (iv) oxide. The resultant solution is then distilled to obtain the local gin.

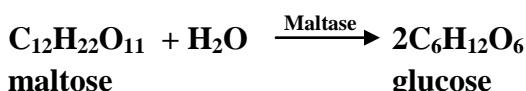
Yeast ferments palm wine until the concentration of ethanol is about 14%. This concentration of ethanol kills the yeast and stops the fermentation process. Exposure of the ethanol to atmospheric air causes its oxidation to ethanoic acid.

Preparation of ethanol from starch

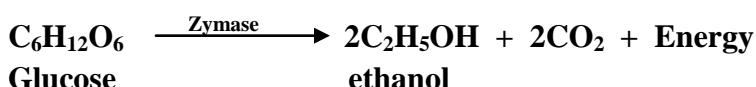
Starchy foods such as maize, cassave, potatoes, guinea corn, millet, etc. are common sources of ethanol. The starch is extracted by first of all grinding the food (e.g. Cassava) and treating it with super-heated steam. Malt- a partially germinated barley is added and kept warm at 60°C for an hour. Malt supplies the enzyme diastase which hydrolyses starch to maltose.



The resultant sugar solution is then cooled to about 25°C and yeast is added. Yeast contains two enzymes: (i) maltase and (ii) zymase. Maltase catalyses the hydrolysis of maltose to glucose.



Zymase then catalyses the conversion of glucose to ethanol.

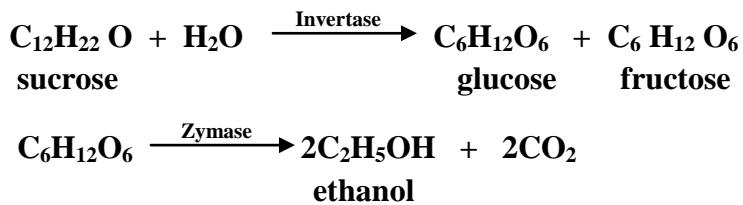


The alcohol is then separated from water by fractional distillation to get about 95% ethanol called rectified spirit. The rectified spirit can further be redistilled to get 99.5% ethanol called an absolute ethanol using quicklime (CaO) which is hydroscopic. Methylated spirit is made by

adding to rectified spirit 10% methanol and unpleasant tasting substances like pyridine and dye to make it unfit for drinking.

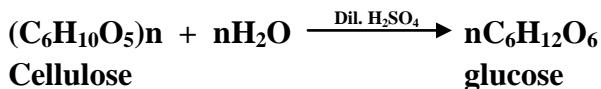
Preparation of ethanol from molasses

Molasses is a syrup liquid left after the crystallization of sugarcane (sucrose) from concentrated sugarcane juice. It contains about 50% uncrystallisable sugar. It is hydrolysed to two isomeric simple sugars (glucose and fructose) by an enzyme - invertase from the yeast. The enzyme zymase from the same yeast converts them (glucose and fructose) to ethanol.

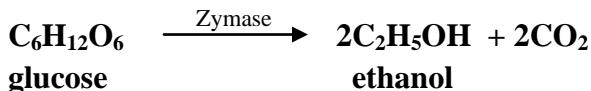


Preparation of ethanol from Cellulose

Saw dust and cotton are examples of cellulosic materials. Saw dust can be hydrolysed to glucose by the addition of dil. H_2SO_4 and steam, at about a pressure of 6 atmospheres.

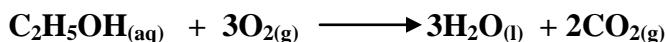


Excess H_2SO_4 is neutralized with calcium hydroxide. Yeast is then added to the resultant solution. The enzyme zymase present in the yeast converts glucose to ethanol.



Properties of alkanol (ethanol)

1. Ethanol is a colourless and volatile liquid with a typical smell and burning taste. It is soluble in water. Ethanol boils at 78°C .
2. It burns in air with a pale blue flame to form water and carbon (iv) oxide.



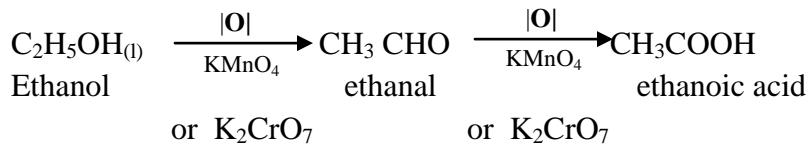
3. Ethanol reacts with sodium or potassium metal at room temperature to liberate hydrogen gas.



The product sodium ethoxide can be hydrolysed to form alkaline solution.

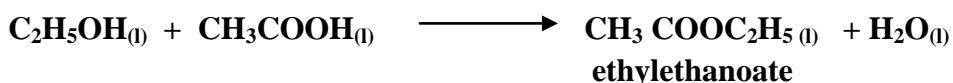


4. In the presence of acidified potassium tetraoxomanganate (vii), KMn_2O_4 or acidified potassium heptaoxo chromate (vi) ethanol is oxidized to ethanal and finally to ethanoic acid.



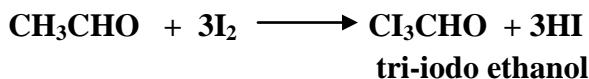
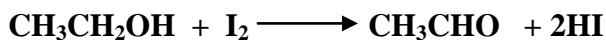
If palm wine is kept for so long, it turns sour because a bacteria in the air called mycodema aceti oxidized the ethanol to ethanoic acid.

5. Organic alcohols react with organic acids to form alkanoates or esters. Ethanol reacts with ethanoic acid to form a sweet-smelling compound called ethylethanoate, an ester.



This reaction is known as esterification reaction.

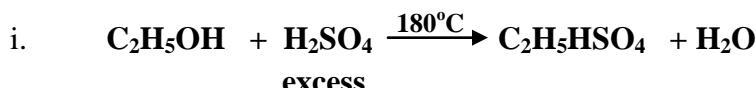
6. Ethanol reacts with iodine in alkaline solution to give yellow crystals of tri-iodoethanal.



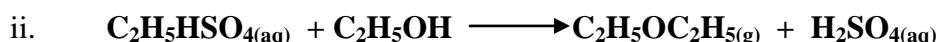
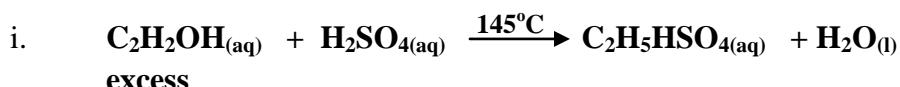
7. Ethanol is dehydrated with conc H_2SO_4

The reaction occurs in two ways:

- a. If the Conc. H_2SO_4 is in excess and at a high temperature of 180°C , a gaseous product - Alkene (ethene) is formed.



- b. If the ethanol is in excess and at a lower temperature of 145°C , a gaseous product, ethoxyethane is formed.



Methanol is produced by destructive distillation of wood. Unlike ethanol, it is poisonous if swallowed. It damages the optic nerves of the eye causing blindness.

Uses of alkanols (ethanol)

1. Ethanol is a good solvent used to dissolve soaps, perfumes, flavouring extracts, dyes, varnishes, drugs etc.
2. It is used in the manufacturing of alcoholic drinks such as wines, beers and spirits (e.g. Gin, whisky, rum, brandy etc).

16.3 Alkanoic Acid

Alkanoic acids are the organic acids. They belong to a homologous series with the general molecular formula $C_nH_{2n-1}COOH$. They have the carboxyl group - COOH as their functional group. Organic acids are also called the carboxylic acids. The higher members of this homologous series are sometimes called the fatty acids because they occur in fats and oil obtained from plants and animals. Examples of such fatty acids are:-

Molecular Formula	IUPAC Name	Old name
$C_{15}H_{31}COOH$	Hexadecanoic acid	Palmitic acid
$C_{17}H_{35}COOH$	Octadecanoic acid	Stearic acid
$C_{17}H_{33}COOH$	Octadec-9-enoic acid	Oleic acid

Octadecenoic acid, $C_{17}H_{33}COOH$ is an unsaturated fatty acid.

Alkanoic acids are named by replacing the 'e' in the parent alkane with 'OIC' followed by the -oic acid.

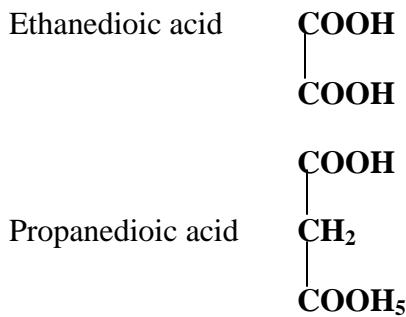
The first three members are as follows:

1. Methanoic acid $HCOOH$
2. Ethanoic acid CH_3COOH
3. Propanoic acid C_2H_5COOH

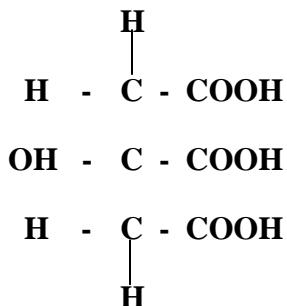
The carboxylic acid with one functional group per molecule is called the monocarboxylic acids

e.g. Ethanoic acid CH_3COOH

Those with two carboxyl groups are called dicarboxylic acids e.g.



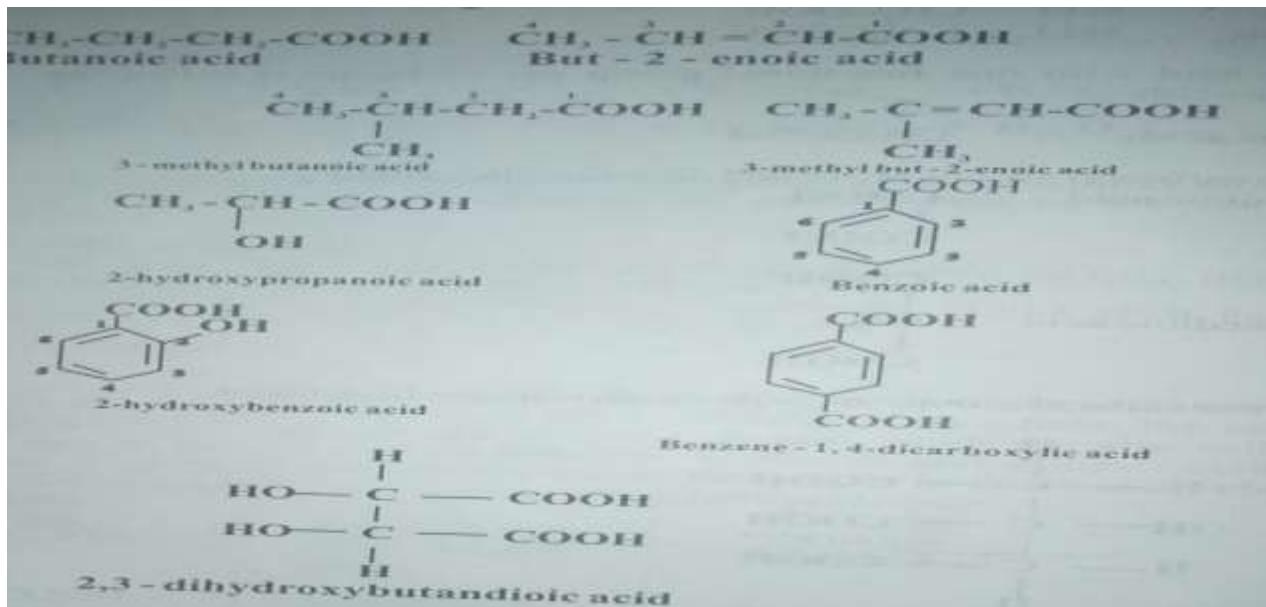
Those with three carboxyl groups are called the tricarboxylic acids eg.



2 - hydroxypropane 1, 2, 3 - tricarboxylic acid

IUPAC Nomenclature of organic acids

1. The longest carbon chain is taken as the parent carboxylic acid.
2. If a substituent is attached to the organic acid, the carbon chain is numbered starting with the carbon of the carboxyl group.



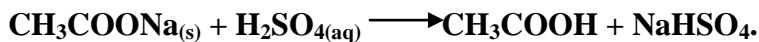
Preparation of ethanoic acid

Ethanoic acid is prepared in the laboratory by:

1. Complete oxidation of ethanol by sodium heptaoxodichromate(vi) solution, $\text{Na}_2\text{Cr}_2\text{O}_7$

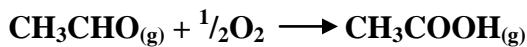


Reacting anhydrous sodium ethanoate with conc. H_2SO_4 .



Industrially it is prepared by:

1. Bacterial oxidation of ethanol: The bacteria called mycoderma aceti oxidizes the ethanol in vinegar to ethanoic acid in the presence of atmospheric oxygen.
2. Converting ethyne to ethanol at 95°C using mercury I tetraoxosulphate (vi) as catalyst. The ethanal is then oxidized with air to ethanoic acid at 60°C .



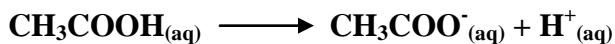
NOTE: Ethanoic acid vapour attacks cork and rubber. The apparatus used during its preparation must all be made of glass.

Properties of ethanoic acid

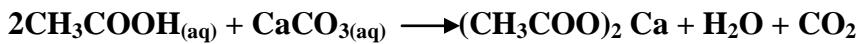
1. Like other acids, aqueous solution of ethanoic acid has a sour taste. It freezes at the temperature below 18°C .

It turns wet blue litmus paper red. Its boiling point is 118°C .

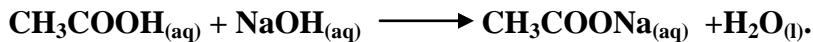
2. It is a weak monobasic acid



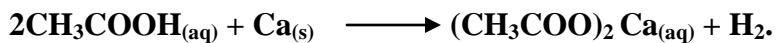
3. It reacts with trioxocarbonate (iv) and hydrogentrioxocarbonate (iv) to liberate carbon (iv) oxide.



4. It neutralizes base or alkali to form salt and water only.



5. It reacts with highly electropositive metals to liberate hydrogen gas.

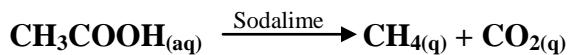


6. Ethanoic acid undergoes esterification reaction. Esterification reaction is the reaction between any alkanoic acid and an alkanol to form ester. Sweet smell of flowers and fruits is caused by the presence of esters in them.

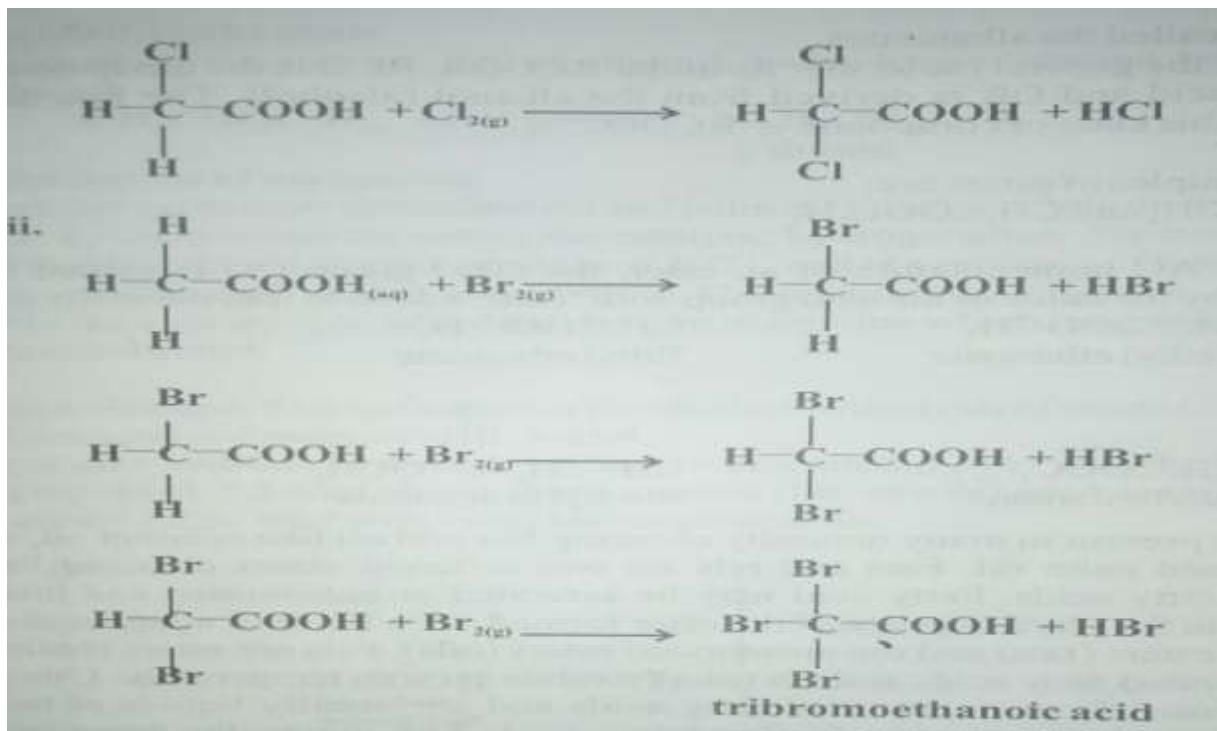
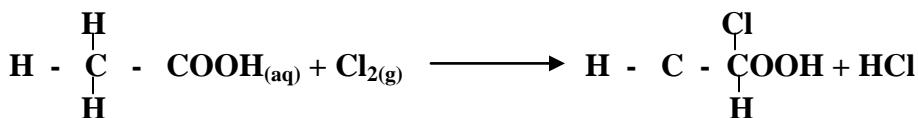


Tetraoxosulphate (vi) acid is used here as a catalyst.

7. An anhydrous ethanoic acid is decomposed to methane and carbon(iv) oxide if it is heated with sodalime. Sodalime is $\text{Ca}(\text{OH})_2$ slackened with NaOH solution. This reaction is known as decarboxylation of ethanoic acid.



8. If chlorine is made to react with the boiling ethanoic acid in the presence of sunlight, monochloroethanoic acid is formed. If chlorine is in excess, di and trichloroethanoic acid will be formed. Similar reaction occurs when the acid reacts with bromine. Sunlight catalyses the reaction.



Uses of ethanoic acid

It is used as an important solvent.

It is used in coagulating rubber latex.

It is used in the manufacture of synthetic fibre like rayon and artificial silk.

16.4 Esters, Saponification and Hydrogenation of Oil

Esters: Ester is a product of esterification reaction. Esterification reaction is the formation of an ester and water by the reaction between an alkanol (alcohol) and alkanoic acid (organic acid). It is a reversible reaction. Conc. H_2SO_4 is usually added as a catalyst and it also absorbs the water formed during the reaction. This action prevents the reverse reaction from taking place.



Esters are called the alkanoates.

They have the general molecular formula $RCOOR$. RCO is the group derived from the acid and OR is derived from the alkanol (alcohol). The functional group for this class of compound is $RCOO^-$.

Some examples of esters are:



In the IUPAC nomenclature of an ester, the alkyl group (R) is named first followed by the name of the acid group with 'Oate' added to it as the suffix e.g



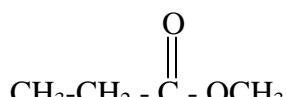
Methyl ethanoate



Ethyl ethanoate



Ethylbutanoate

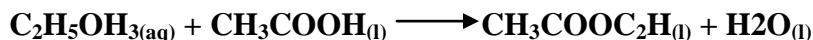


Methyl propanoate

Esters are present in many naturally occurring fats and oil like coconut oil, cod liver oil and palm oil. Fats and oils are two different esters obtained from different fatty acids. Fatty acid may be saturated or unsaturated and this is reflected in the physical state of the ester formed. Two kinds of alkanoates are saturated esters (fats) and the unsaturated esters (oils). Fats are esters obtained from saturated fatty acids and are usually solids at room temperature. Oils are esters formed from unsaturated fatty acids and are usually liquids at room temperature. An example of oil is oleic acid. The molecular formula is $C_{17}H_{33}COOH$ but its structural formula is $CH_3(CH_2)_7CH=CH(CH_2)_7COOH$.

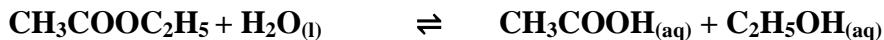
Preparation of ethylethanoate, $\text{CH}_3\text{COOC}_2\text{H}_5$.

$\text{CH}_3\text{COOC}_2\text{H}_5$ is prepared by the reaction between ethanol and ethanoic acid at 150°C in the presence of Conc. H_2SO_4 .



Properties

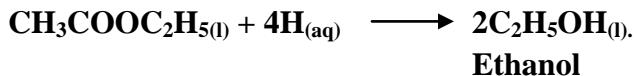
1. It is a colourless, volatile and sweet smelling liquid. Its boiling point is 75°C . It is soluble in organic solvents like benzene, ethanol etc.
2. It can be hydrolysed to ethanoic acid. The reaction is catalysed by hydrogen or hydroxide ions.



- 3 If NaOH is used in the hydrolysis, sodium ethanoate is formed.

Uses of ethylethanoate

It is used in the preparation of ethanol by its hydrogenation with nascent hydrogen.

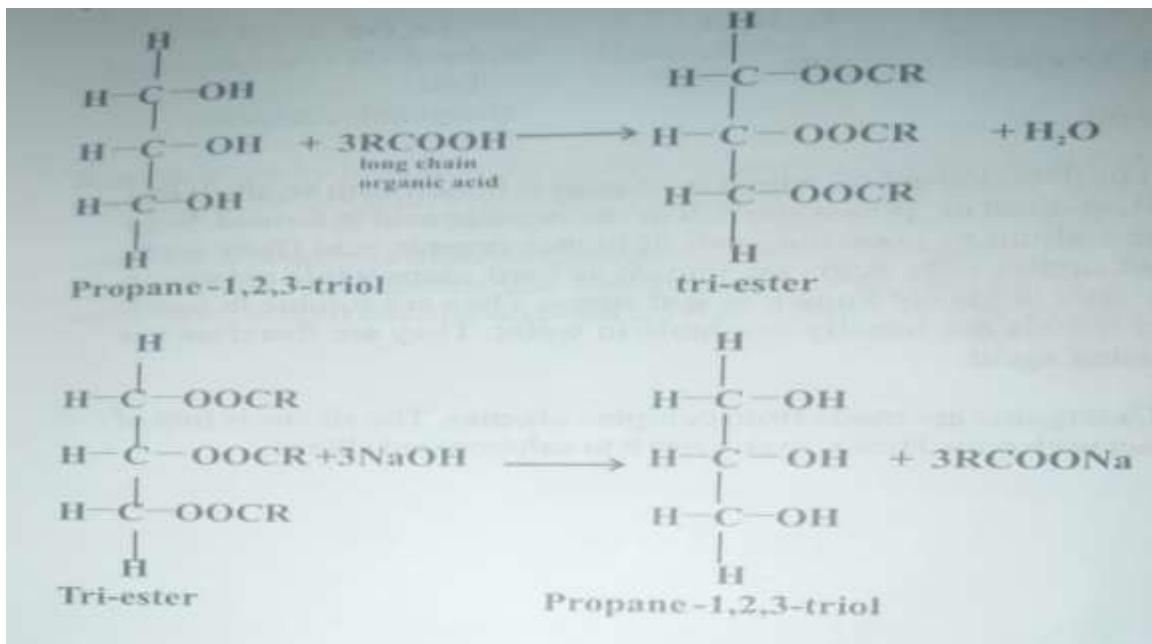


Manufacture of margarine

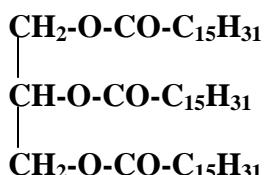
Vegetable oil such as groundnut oil and palm oil is refined. It is then converted to fat by the process known as the catalytic hydrogenation. Hydrogen gas is bubbled into the oil at a temperature of 200°C and at a pressure of 5 atmosphere in the presence of nickel as a catalyst. The oil hardens. Vitamins, milk, salt, food colourant are then added to it to form margarine which is a good substitute for natural butter.

Saponification: Saponification is the alkaline hydrolysis of esters (fats or oils) to form soap and propane -1, 2, 3-triol.

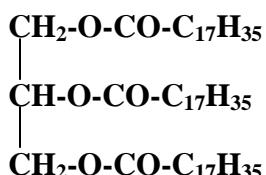
Fat or oil + sodium hydroxide \rightarrow Soap + propanediol, 1,2,3 - triol the formation of propane -1,2,3-triol shows that the esters (fats or oils) used are tri-esters of propane - 1,2,3 - triol with long chain organic acids.



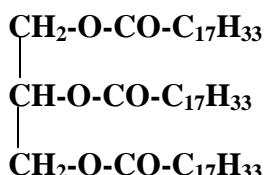
Saponification can also be defined as the hydrolysis of tri-esters of propane - 1,2,3-triol; to form soap. Fats and oils used during soap production are the tri-esters of propane- 1,2,3-triol. Some examples of tri-ester are;



Propane - 1,2,3 trihexade -
Canoate (Tripalmitin)
Occurs in palm oil



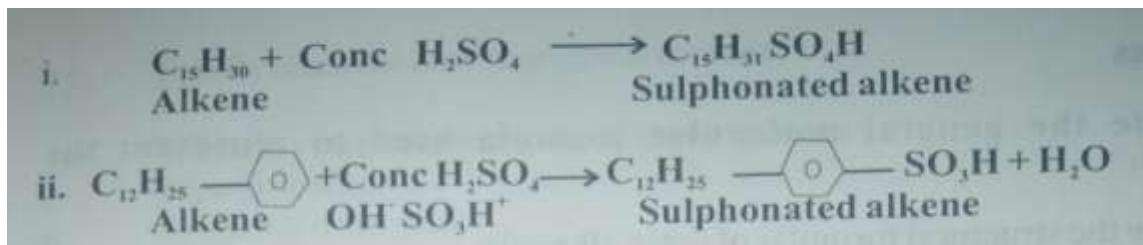
Propane-1,2,3-trioctadecanoate
(Tristearin) Occurs in beef or mutton fat



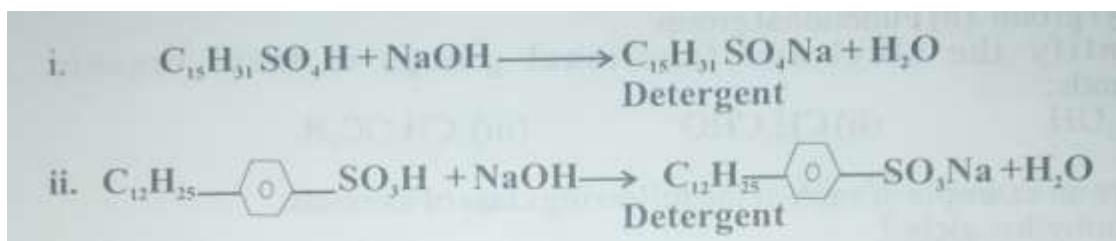
Propane - 1,2,3-trioctadec - 9-enoate (Triolein) Occurs in olive oil

When palm oil or palm kernel oil which is an ester is boiled with an alkali like NaOH or KOH, sodium or potassium salt of the organic acid is formed. Soap is therefore the sodium or potassium salt of higher organic acid (fatty acid). Sodium salts of higher fatty acids are known as hard soaps while potassium salts of higher fatty acids are known as soft soaps. They are soluble in water. Soaps of other metals are usually insoluble in water. They are therefore not used as a cleansing agent.

Detergents: Detergents are made from complex alkenes. The alkene is first of all made to react with conc H_2SO_4 to convert it to sulphonated alkene.

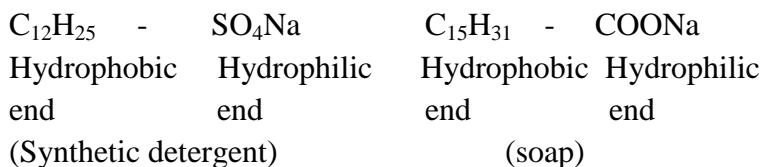


The sulphonated alkene is then hydrolysed with NaOH or KOH to form the sodium or potassium salt of sulphonated alkene called detergent.



Cleaning action of soap and detergent

Soap and detergents are polar molecules with water soluble (hydrophilic) and fat soluble (hydrophobic) ends.



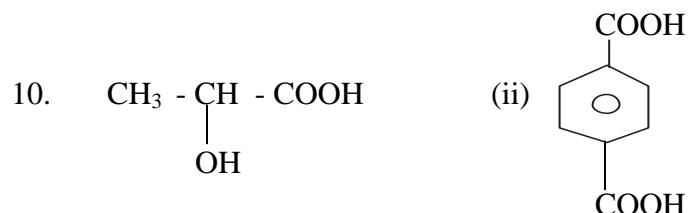
Soap cleanses by lowering the surface tension of water and by loosening the dirt particle on the cloth. The dirt then floats in the foam.

Synthetic detergent is more advantageous than soap because it is not affected by hard water. This means that calcium salt of detergent is very soluble in water. The calcium salt of soap is not soluble in water. It forms scum.

Exercises

1. Write the general molecular formula used to represent the alkanols.
2. Draw the structural formular of these alkanols:
 - a. Propan-2-Ol (b) 2-Chloroethanol

3. What do you understand by these terms:
- (a) Alkyl group (b) Functional group.
- (c) Identify the alkyl and functional groups in these organic compounds:
- i. $\text{C}_2\text{H}_5\text{OH}$ (ii) CH_3CHO (iii) $\text{CH}_3\text{OC}_2\text{H}_5$
4. Write an example of each of the following class of alcohols:
- (a) Monohydric alcohol
- (b) Dihydric alcohol
- (C) Trihydric alcohol.
5. Use an equation to show one chemical property of an alcohol.
6. Why is it that fresh palm wine becomes sour when it is kept for five days?
7. Name the first two members of the alkanoic acid.
8. Name two fatty acids we studied in this chapter.
9. Write the IUPAC nomenclature of these alkanoic acids.



11. Why is it that cork and rubber material is not used during laboratory preparation of ethanoic acid?
12. Write three properties of ethanoic acid.
13. What is an ester? Write the formular and IUPAC nomenclature of three esters.
14. What is the difference between fat and oil. Write a named structural formular of each.
15. Briefly describe how margarine is made.
- 16a. What do you understand these terms to mean:
- i. Saponification
- ii. Synthetic detergent

- iii. Soap
- b. Why is synthetic detergent said to be a better cleaning agent than soap?

PART 3

CHAPTER 1

QUANTITATIVE AND QUALITATIVE ANALYSIS

1.1 Acid-Base Titration: Read chapter 4, part 2 of the book.

1.2 What examiners look for in acid-base titrations

Introduction

Paper 1 in WAEC and NECO practical chemistry examinations at the SS level; is a two-hour practical test. School candidates usually take it.

The paper consists of three compulsory questions and carries 50 marks. The first question is on quantitative (volumetric) analysis, usually, acid-base titration; the second question is on qualitative analysis, and the third question is to test candidates familiarity with the practical activities suggested in the teaching syllabus (WAEC, 2004).

Areas of Emphasis

Recording of Titration Data

1. The students should obey all the instructions. All the questions are to be answered in ink; if not, marks will be deducted.
2. Solution of the acid must be in the burette and not the alkaline solution. If the alkaline solution is placed in the burette, mark will be deducted.
3. The burette readings (final and initial) must be tabulated. If a candidate puts down the titre values without indicating how they are obtained, marks are deducted.
4. Students should be trained on how to take and record the burette readings to two places of decimal. They are expected to be consistent when taking down such readings: If a student consistently or inconsistently record burette readings to the nearest whole numbers, he or she is penalized by deduction of marks (see Tables 1 and 2). In Table 1, the burette readings have been consistently recorded to the nearest whole numbers. The two recordings are wrong; some marks will be deducted.

Table 1:

Burette reading	I	II	III
Final (cm^3)	26	25	26
Initial (cm^3)	0	0	1
Volume of acid used (cm^3)	26	25	26

Table 2:

Burette reading	I	II	III
Final	26.0	26	27.6
Initial	0	1.40	2.0
Volume of acid used (cm ³)	26	25.60	25.6

5. The unit volume, cm³, must be attached to all the burette readings and the titre values, or, as shown in Tables 1 and 2.. usually, wrong unit or no unit of volume attracts deduction of marks.
6. The initial and final burette readings must not exceed 50.00 cm³- as in titration numbers III and IV in Table 3. If titre values obtained from such titrations are used in averaging, marks will be deducted.

Table 3:

Burette reading	I (cm ³)	II (cm ³)	III (cm ³)	IV (cm ³)
Final	24.50	49.00	73.50	98.00
Initial	0.00	24.50	49.00	73.50
Volume of acid used	24.50	24.50	24.50	24.50

7. Arithmetical error should be avoided (see Table 2, titration number II, and Table 4, titration number II). Such errors attract deduction of marks if the titre values are used in averaging.
8. The average volume of acid used should be from two or more titre values in which the difference between any two titres is not more than 0.20cm³. Such titre values are said to be consistent.

The teacher should note that when averaging, a student can use the titre value obtained from the first titration, called rough titre, as long as it is consistent with other titres used in averaging. (see Table 4 below).

Table 4:

Burette reading	Rough	I	II	III
Final (cm ³)	24.70	25.50	25.40	26.40
Initial (cm ³)	0.00	1.00	0.00	2.00
Volume of acid used	24.70	24.50	24.50	24.40

Question 1: Which combinations of titre values can be used in averaging from Table 4?

Answers: (Rough +I); (Rough +II); (I + III); (Rough + I +II); (I + II + III).

Question 2: Which combinations of titres **cannot** be used in averaging from Table 4?

Answers: (Rough + III); (Rough + I + III); (Rough +II + III); (Rough + I + II +III).

9. The average titre value should be expressed to two places of decimal. Consider the following combinations of titre values from Table 4:

a. Average titre from Rough + I + II:

$$\frac{24.70 + 24.50 + 24.50}{3} = \frac{73.70}{3} = 24.57\text{cm}^3 \text{ (correct)}; 24.6\text{ cm}^3 \text{ (wrong)}$$

b. Average titre from titre I + II + III:

$$\frac{24.50 + 24.50 + 24.40}{3} = \frac{73.4}{3} = 24.47\text{cm}^3 \text{ (correct)}; 24.5\text{cm}^3 \text{ (wrong)}$$

c. Average titre from titres II and III:

$$\frac{24.50 + 24.40}{2} = \frac{48.90}{2} = 24.45\text{ cm}^3 \text{ (correct)}; 24.5\text{ cm}^3 \text{ (wrong)}$$

Average titres approximated or rounded-up as in the above examples attract no mark.

10. Students must not alter any burette reading or titre value, or, cancel a table of titres outright. If done, some marks will be deducted.

Handling of Acid-Base Titration Variables

Students should be trained on how to handle titration variables. This is by practice.

1. Students should be advised to read the question on titration several times, in order to know the parameters to be determined. Standard solutions of acids and alkalis can be used to determine the following parameters:

- The concentration of an acid or alkaline solution;
- The molar mass of an acid or base;
- The water of crystallization in an acid or base;
- The percentage of purity of an acid or base;
- The solubility of an acid or base.

The procedure for solving problems on acid-base titration depends on the parameters to be determined.

2. The following variables should be identified from the titration results and the information provided:

C_a = Concentration of acid, in mol dm⁻³

V_a = Average volume of acid used, in cm³ (or dm³);

C_b = Concentration of base, in mol dm⁻³

V_b = Size of the pipette used, in cm³ (or dm³);

n_a = Amount of the acid: its coefficient in the balanced equation of reaction;

n_b = Amount of the base: its coefficient in the balanced equation of reaction.

3. C_a and C_b must be expressed in mol dm⁻³, while V_a and V_b must be expressed in the same unit of volume. The importance of units cannot be over-emphasized in chemical calculations.

- When the concentration of a pure acid or base is given in grams per dm³, it must be converted to moles per dm³. The expression to be used is: Conc, in mol dm⁻³ = (Conc, in g dm⁻³)/(Molar mass, g mol⁻¹)
 - When mass concentration is given as a fraction of a per dm³, it must be converted to grams per dm³. E.g. A solution which contains 2.0 g of an acid in 250 cm³ of solution has a mass concentration of 8.0 g dm⁻³.
4. The mole ratio expression for an acid-base titration given as $C_a V_a / C_b V_b = n_a / n_b$, can only be used if five of the six variables are known, or when only the mole ratio n_a/n_b is unknown.
5. The unit of molar mass is grams per mole (g mol⁻¹). Usually, no mark will be awarded when a wrong unit is attached to a correctly calculated molar mass.
6. The final answer to any calculation must be expressed to three significant figures; if not, marks may not be awarded.

Preparation of Solutions for Acid-Base Titrations

All the candidates in a particular center or school are supposed to be supplied with the same set of solutions for the acid-base titration. They are also expected to use the same size of pipette, as the one used by the teacher-in-charge. Hence, all the candidates are expected to state the size of their pipette (either 20.0 cm³ or 25.0 cm³). This is because they are usually assessed relative to the average titre value provided by the teacher-in-charge.

When the titre values used by a candidate in averaging are not within 0.40 cm³ of the average titre provided by the teacher-in-charge, the candidate does not score any mark for titration! One major cause of large variations in titre values is due to the use of other sources of water in the preparation of the solutions of the acid and base used for the titration-instead of distilled water. There is no alternative to distilled water. The chemistry teacher is supposed to use distilled water for the preparation of laboratory reagents. Any solution in which distilled water is not used

will be heterogeneous. In order to facilitate consistent titre values between those of the students and that of the teacher, distilled water should be used in the preparation of solutions for titrations.

1.3 Sample Questions and Answers

Question 1.

The following is a typical WAEC/NECO question on acid-base titration. It illustrates the use of a standard solution of a base for the determination of the percentage purity of an acid. The answers provided illustrates the procedure for solving problems on acid-base titration-based on the determination of percentage purity.

A contains 4.0 g of an impure tetraoxosulphate (VI) acid in 250 cm³ of solution.

B is 6.0 g dm⁻³ of sodium hydroxide solution.

In a titration experiment, 25.0 cm³ of **B** required an average of 16.78cm³ of **A** for complete neutralization. Calculate the:

- a. concentration of **B** in mol dm⁻³;
- b. amount of NaOH in 25.0cm³ of **B** used;
- c. concentration of **A** in mol dm⁻³;
- d. mass of the pure acid per 250 cm³ of **A**;
- e. mass of impurity per 250 cm³ of **A**;
- f. percentage purity of the acid used in preparing **A**.

Equation for the reaction is:



[H = 1, O = 16, Na = 23; S = 32]

Solution

- a. To calculate the molar concentration of NaOH.

Given: Mass concentration of NaOH = 6.0 g dm⁻³

The expression that gives the relationship between the molar and mass concentrations of a solution of a chemical substance is:

Molar conc = (mass conc)/(molar mass)

Molar mass of NaOH = $23 + 16 + 1 = 40.0 \text{ g mol}^{-1}$

b. To calculate the amount of NaOH in 25.0 cm^3 of B

Given: Molar concentration of B is $0.150 \text{ mol dm}^{-3}$

The amount of solute in a given volume of a standard solution is obtained from the expression:
Amount (mol) = Conc (mol dm⁻³) x Volume (dm³)

Substituting: Amount = $0.150 \times 25/1000 = 0.00375 \text{ mol}$ (or $3.75 \times 10^{-3} \text{ mol}$)

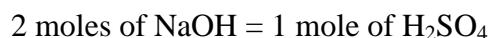
c. To calculate the molar concentration of A

Since the acid in A is impure, its molar concentration cannot be obtained from the mass concentration given. Instead, the molar concentration of the pure acid, which neutralized the given the volume of the standard solution of the base is obtained from the mole ratio expression. There are two methods:

Method 1: From the first principle (Direct proportion) see chapter 3.1, part 2 of this book

Amount of the base in B used is 0.00375 mol .

From the balanced equation of reaction :



$$\therefore 0.00375 \text{ mol NaOH} = \frac{1}{2} \times 0.00375 \text{ mol H}_2\text{SO}_4$$

$$= 0.00188 \text{ mol of H}_2\text{SO}_4$$

Given that the average volume of A used is 15.78 cm^3 ; i.e.

15.78 cm^3 of A contained $0.00188 \text{ mol of H}_2\text{SO}_4$

$\therefore 1000 \text{ cm}^3$ of A contained $0.00188 \times 1000/15.78 \text{ mol}$

$$= 0.119 \text{ mol of H}_2\text{SO}_4$$

Hence, the concentration of A is $0.119 \text{ mol dm}^{-3}$

Method 2: Formula method

The mole ratio expression is: $C_A V_A/n_A = C_B V_B/n_B$

$C_A = ? \text{ mol dm}^{-3}$; $V_A = 15.78 \text{ cm}^3$; $n_A = 1$;

$C_B = 0.150 \text{ mol dm}^{-3}$; $V_B = 25.0 \text{ cm}^3$; $n_B = 2$

Making the unknown, C_A , the subject of the formula, and substituting

$$C_A = \frac{0.150 \times 25.0 \times 1}{15.78 \times 2} = 0.119 \text{ mol dm}^{-3} \quad (\text{3 sig. fig.})$$

d. To find the mass of the pure acid in 250 cm^3 of A

First, calculate the mass of the pure acid in one dm^3 of A, i.e., the mass concentration, using the expression: Mass conc = molar conc x molar mass.

Molar conc of $\text{H}_2\text{SO}_4 = 0.119 \text{ mol dm}^{-3}$

Molar mass of $\text{H}_2\text{SO}_4 = 2(1) + 32 + 4(16) = 2 + 32 + 64 = 98.0 \text{ g/mol}$

Substituting: Mass conc = $0.119 \times 98.0 = 11.7 \text{ grams per dm}^3$

Hence, if 1000 cm^3 (1 dm^3) of A contained 11.7g;

250 cm^3 of A will contain $11.7 \times 250/1000$

$$= 2.93 \text{ g} \quad (\text{3 sig. fig.})$$

i.e. 2.93 g of pure acid are in 250 cm^3 of A.

e. To calculate mass of impurity per 250 cm^3 of A

Given that 250 cm^3 of A contained 4.0 g of impure acid; found that 250 cm^3 of A contained 2.93 g of pure acid. Hence, mass of impurity per 250 cm^3 of A will be:

$$(4.00 - 2.93) \text{ g} = 1.07 \text{ g}$$

f. To calculate the percentage purity of A

% Purity = $(\text{mass of pure acid per } 250 \text{ cm}^3 \text{ of A}) \times 100$

$\text{mass of impure acid per } 250 \text{ cm}^3 \text{ of A})$

$$= \frac{2.93 \times 100}{4.00} = 73.3\% \quad (\text{3 sig. fig.})$$

Questions 2.

This is also a typical WAEC/NECO question on acid-base titration. However, the pattern of the questions is such that the mole ratio expression cannot be used in the determination of the concentration of the unknown. The answers provided illustrate the procedure for solving problems on acid-base titration – based on the determination of the amount water of crystallization in an acid.

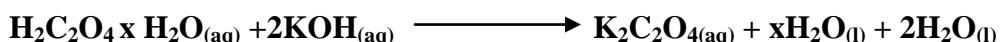
Be reminded that when a hydrated substance is used in an acid-base titration, only the anhydrous part takes part in the reaction; the water of crystallization remains in the solution of just like an impurity.

7.0 g of a hydrate of ethanedioic acid, $\text{H}_2\text{C}_2\text{O}_4 \times \text{H}_2\text{O}$ were dissolved in 1dm^3 of solution. An average of 22.2cm^3 of the resulting solution completely neutralized 25.0cm^3 of $0.101 \text{ mol dm}^{-3}$ potassium hydroxide.

- a. Write a balanced equation of reaction
- b. Suggest a suitable indicator for the titration;
- c. How many moles of ethanedioic acid react with one mole potassium hydroxide?
- d. Calculate the
 - i. amount of KOH in 25.0 cm^3 of $0.101 \text{ mol dm}^{-3}$ solution.
 - ii. amount of ethanedioic acid that reacted with 25.0cm^3 of $0.101 \text{ mol dm}^{-3}$ KOH.
 - iii. amount of the acid in the original 7.0g
 - iv. amount of water of crystallization in the original 7.0g
 - v. value of x [H = 1; C = 12; O = 16]

Solution

- a. The balanced equation of reaction is:



- b. The suitable indicator is phenolphthalein; since the reaction is between a weak acid and strong base.
- c. To state the mole-mole relationship

From the above equation: 2 moles of KOH = 1 mole of ethanedioic acid

$$\therefore 1 \text{ mole of KOH} = \frac{1}{2} \text{ or } 0.500 \text{ mol of the acid.}$$

- d. i. Amount of KOH in 25.0 cm^3 of $0.101 \text{ mol dm}^{-3}$ solution
 1000 cm^3 of solution contained 0.101 mol KOH .

$\therefore 25\text{cm}^3$ of solution will contain $0.101 \times 25/1000 \text{ mol KOH}$

$$= 0.00253 \text{ mol}$$

or Amount (mol) = conc (mol dm⁻³) x volume (dm³)

$$= 0.101 \times 25/1000 = 0.00253 \text{ mol}$$

ii. Amount of acid that reacted with KOH.

From the equation of reaction: 2 moles of KOH = 1 mole of ethanedioic acid

$$\therefore 0.00253 \text{ mol of KOH} = \frac{1}{2} 0.00253 \text{ mol}$$

$$= 0.00127 \text{ mol of acid.}$$

iii. Amount of acid in the original 7.0 g

22.20 cm³ of solution contain 0.00127 mol of acid.

$\therefore 1000 \text{ cm}^3$ of solution will contain $0.00127 \times 1000/22.2 \text{ mol of the acid}$

$$= 0.0572 \text{ mol of acid.}$$

iv. Amount of H₂O in the original 7.0 g

First, find the mass of anhydrous acid, given that the mass of the hydrate = 7.0 g

Amount of the anhydrous acid = 0.0572 mol

Molar mass of H₂C₂O₄ = 2(1) + 2(16) = 2 + 24 + 64 = 90 g mol⁻¹

Mass of the anhydrous acid = amount (mol) x molar mass (g mol⁻¹)

$$= 0.0572 \times 90 = 5.15 \text{ g}$$

\therefore Mass of water crystallization = (7.00 – 5.15) g = 1.85 g

Amount of H₂O = Mass/Molar mass

$$= 1.85/18 = 0.103 \text{ mol}$$

v. To find the value of x.

Amounts of anhydrous acid and water of crystallization can be in a simple ratio;

i.e. acid : water = 0.0572 mol : 0.103 mol = 1 : 1.80 = 1 : 2

Hence, the value of x is 2.

Formula of the hydrate is H₂C₂O₄ . 2H₂O

Conclusion

If our students are familiar with what examiners look for in acid-base titrations, and possess the required knowledge on the procedure for solving problems on acid-base titrations, then, there would be marked improvement in the performance of students in WAEC/NECO practical chemistry examinations.

1.4 Redox Titrations Involving Fe^{2+} , KMnO_4 , C_2O_4 , I_2 , KI and $\text{S}_2\text{O}_3^{2-}$

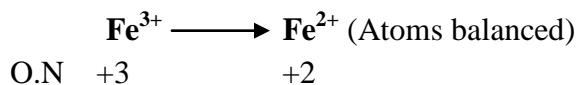
After redox titration, the chemistry teacher should teach the students; writing and balancing Redox Equation. For the teacher to simplify the teaching and learning of the concept, he will start by teaching oxidation or reduction equation independently. An equation which represents oxidation or reduction process is called a half equation.

Balancing Half Equations

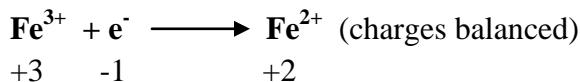
Balancing Simple Half Equations: Simple half equations contain simple ions. Usually, the product of oxidation or reduction of the reactant in the half equation is provided. There are three steps to be followed in order to balance a simple half equation.

- i. Assign oxidation number (O.N) to each atom of the element in the half equation;
- ii. Balance the number of atoms on both sides, so as to conform to the law of conservation of mass.
- iii. Add the appropriate numbers of electrons to account for the number of electrons lost or gained, i.e to balance the charges.

Example 1 A half reaction involving ions is:



The change in O.N of Fe atom is from +3 to +2. It is a decrease by 1 unit to the left. Hence, add Ie^- to the left of the equation:



(Note: change on O.N of Fe = (+2) – (+3) = -1.

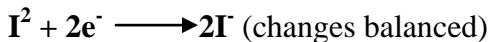
It corresponds to a gain of 1 electron. Add Ie^- to the lefty.

Example 2: A half reaction involving iodine is :



O.N. O

The change in O.N of each iodine atom is O to $-\text{I}$. It is a decrease by I unite to the left. Hence, for the 2 iodine atoms, add 2e^- to the left.



O.N O -2 = -2

(Note: Change in O.N of iodine, 1 = $(-1)-(0) = -1$. It corresponds to a gain of 1 electron for each atom. Add 2e^- to the left).

b. Balancing Complex Half Equations

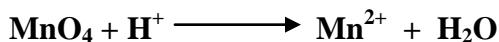
Complex redox reactions are those that involve oxyanions as the oxidizing or reducing agents, or both. Such reactions take place in a neutral, an acidic, or alkaline medium.

To balance a complex half equation acid medium:

There are six steps to be followed in order to balance a complex half equation in acid medium.

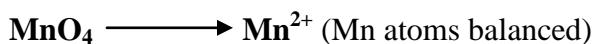
- i. Balance all the atoms except those of H and O
- ii. Assign oxidation number to the atoms that have changed
- iii. Add appropriate number of electrons to balance electron gain or loss.
- iv. Add appropriate number of H_2O molecules to balance oxygen atoms.
- v. Add appropriate number of H^+ to balance hydrogen atoms.
- vi. Check on charge.

Example 1; Redox reactions involving tetraoxomanganese (VII) ions in acid medium is written as:

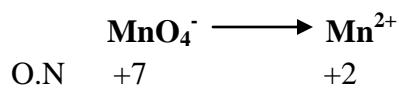


This is a half equation. The presence of H^+ shows that the reaction is in acid solution. the equation can be balanced as follows:

- i. Balance atoms of Mn



- ii. Assign O.N on Mn atoms



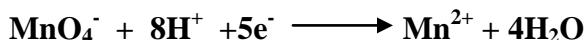
- iii. Add appropriate number of electron-gain or electron loss. The change in O.N of Mn atom is from -7 to +2. It is a decrease by 5 units to the left. Hence, add $5e^-$ to the left of the equations: $\text{MnO}_4^- + 5e^-$

Mn^{2+} . (Change in O.N of Mn = (+2) – (+7) = $5e^-$. It corresponds to gain of electrons. Add $5e^-$ to the left)

- iv. Add appropriate number of H_2O molecules to balance oxygen atoms. Add H_2O to the right to balance O atoms.



- v. Add appropriate number of H^+ to balance hydrogen atoms. Add 8H^+ to the left, to balance H atoms:



- vi. Check on charge

$$-1 + 8(+1)(-5) = +2 + 0$$

$$-1 + 8 - 5 = +2$$

$$+2 = +2$$

Balancing Redox Equations

Once the students have learnt the balancing of half equations, teach balancing of redox equations. In order to obtain the equation of a redox reaction, an oxidation half equation, must be paired with a reduction half equation. In order to balance such a reaction, the number of electrons lost by oxidation must be equal to the number electrons gained by reduction. Two methods of balancing redox reactions are (1) Ion-electron method and (2) Oxidation number method.

1. Ion Electron Method:

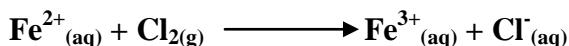
Ion- electron method involves splitting the redox equation into oxidation and reduction halves. The reactants and products of such reactions are usually provided. Six steps are involved in order to balance a redox equation.

- i. Assign oxidation number (ON) to each atom of element that has changed in the equations.

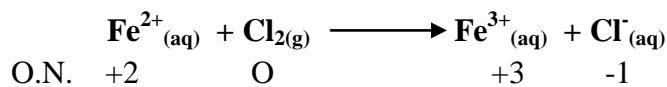
- ii. Separate the equation into oxidation half and reduction half.
- iii. Balance each half equation as treated above for neutral, acidic or alkaline medium.
- iv. Multiply each half equation by the appropriate coefficients, to balance electron-loss and electron gain.
- v. Add the two half equations canceling the common terms.
- vi. Check on the charges

Example 1

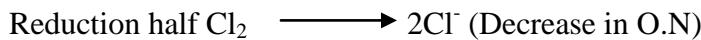
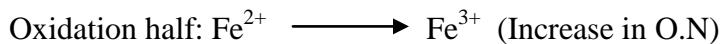
Balance the following redox equation by ion electron method:



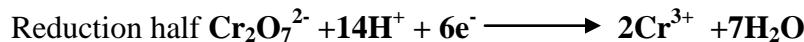
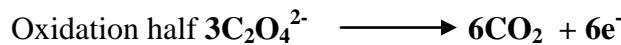
- i. Assign oxidation numbers (O.N):



- ii. Separate the reaction into oxidation and reduction half:

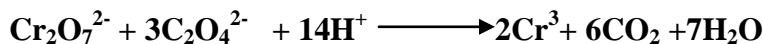
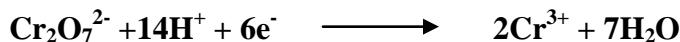
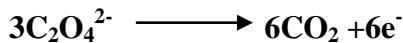


There are loss of two electrons in the oxidation half and a gain of six electrons in the reduction half. Hence multiply oxidation half by 3.



- v. Add the two half equations, canceling the common terms:

cancel the 6e^- common to them:



- vi. Check on the charges:

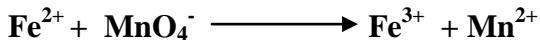
$$(-2) + 3(-2) + 14(+1) = 2(+3)$$

$$-2 - 6 + 14 = +6$$

$$+6 = +6$$

2. Oxidation Number Method:

E.g 1: Consider the redox reaction in acid solution:

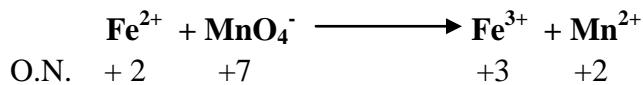


The steps to follow are:

Balance all the atoms except those of H and O.

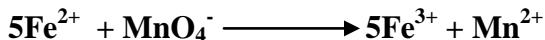


- Assign oxidation numbers, to atoms that have changed:

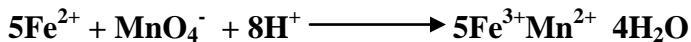


Change in O.N of Fe is $+3 - (+2)$: Ie loss of $1e^-$ by the reducing agent, Fe^{2+} . Change in O.N of Mn is $+2 - (+7) = -5$; Ie. Gain of $5e^-$ by the oxidizing agent, MnO_4^- .

- Equalize changes in oxidation number for each MnO_4^- reduced, five Fe^{2+} must be oxidized. Hence, multiply Fe atoms by 5 and Mn atoms by 1 to obtain:



- Balance hydrogen atoms with H^+ : Add 8H^+ to the left:



- Check on the charges:

$$5(+2) + (-1) + 8(+1) = 5(+3) + 2$$

$$+10 \quad -1 \quad +8 \quad = 15 +2$$

$$+17 \quad +17$$

1.5 Laboratory Tests for Redox Reagents

The teacher is expected to perform simple test tube reactions here.

- Tests for a Reducing Agent

Generally, elements especially metals in groups I, II and III in the periodic table are reducing agents. Other common reducing agents included hydrochloric acid; aqueous solutions of potassium iodine, tin (II) chloride, SnCl_2 , iron (II) salts, and iron (II) hydroxide.

Sulphur (IV) oxide, SO_2 gas can act as an oxidizing and as a reducing agent. Some non metals act as oxidizing and reducing agents, e.g. carbon, sulphur and phosphorus.

An oxidizing agent is used to test for a reducing agent. Any substance that reacts in one or more of the following reactions is a reducing agent · To the solid substance in a test tube, add a few drops of bench dilute hydrochloric acid. There is effervescence; the gas given off is colourless, odourless, has no action on moist red or blue litmus paper, but gives a pop sound with lighted splint. The gas is hydrogen, while the substance is a metal above hydrogen in the activity series.

- To the solution of the substance, add a few drops of purple acidified KMnO_4 solution. The KMnO_4 solution is decolourised.
- To the solution of the substance, add a few drops of yellow acidified $\text{K}_2\text{Cr}_2\text{O}_7$ solution. the yellow colour of $\text{K}_2\text{Cr}_2\text{O}_7$ solution turns green.
- To the solution of the substance, add a few drops of yellow or brown solution FeCl_3 . The solution turns green.

Tests for An Oxidizing Agent

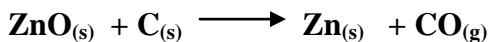
Generally, elements that are diatomic and in groups VI and VII in the periodic table are oxidizing agents, e.g. oxygen, fluorine, chlorine, bromine and iodine. Others are hydrogen peroxide, H_2O_2 ; manganese (IV) oxide, $\text{MnO}_{2(s)}$; concentrated HNO_3 and H_2SO_4 ; acidified solutions of KMnO_4 and $\text{K}_2\text{Cr}_2\text{O}_7$; and solutions of iron (II) salts and iron (II) hydroxide.

A reducing agent is used to test for an oxidizing agent. Any substance that reacts in one or more of the following reactions is a oxidizing agent:

- To the solution of the substance, add a few drops of colourless acidified potassium iodine solution.
The solution turns brown, due to the evolution of iodine, which turns starch solution blue-black.
- Heat a little of the substance with HC1 .
A greenish-yellow gas with pungent Smell is evolved. The gas, which bleaches litmus paper is chlorine.
- Heat the solid substance strongly in attest tube.
The gas evolved is colourless, odourless, has no action on litmus paper, but relights a glowing splint. The gas is oxygen.

Worked Examples

1. In the reaction:



- Is the law of conservation of matter obeyed? Explain.
- Mention the reactant that has been i. oxidized, and the product of oxidation ii. Reduced, and the product of reduction.
- What name is given to the substance that has been I. oxidized?
ii. Reduced?

Tests For Ions

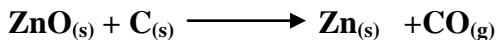
- For Fe^{2+} and Fe^{3+} read chapter 4, part 3 of this book.
- For chlorine and hydrogen, chloride, read chapter 9 part 2 of this book.
- For NO_3^- and ammonia gas read chapter 10, part 2 of this book.
- For sulphide (S^-) and SO_4^{2-} , read chapter 11, part 2 of this book.

For oxygen, read chapter 8, part 2 of this book. Then read chapter 8 of part 3 for test of the macro-molecules and simple sugar. Fats and oils, protein, starch and simple sugar.

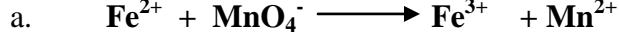
Exercises

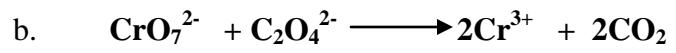
1. Define: (a) Neutralization reaction (b) half equation (c) Redox equation

2. In the reaction:



- Is the law of conservation of matter obeyed? Explain
 - Mention the reactant that has been (i) oxidized, and the product of the oxidation (ii) reduced, and the product of reduction (iii) reduced, and the product of reduction.
 - What name is given to the substance that has been (i) oxidized? (ii) reduced?
3. Balance these redox equations using either (i) ion-electron method or (ii) oxidation number method.





CHAPTER 2

PETROLEUM

2.1 Origin and Composition of Crude Oil

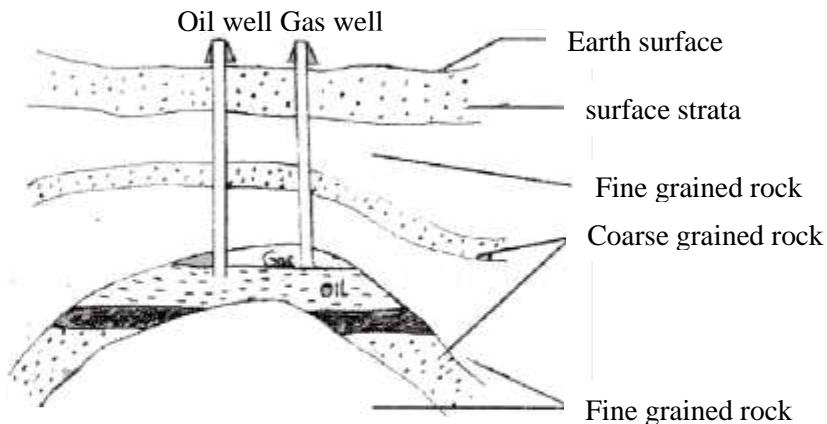
The term petroleum is made up of two words; petra meaning rock and 'oleum' meaning oil. It is usually applied to dark coloured oily liquid found below the earth's crust. It is obtained by the process of mining. Petroleum is therefore known as mineral oil or crude oil. The petroleum industry came into existence in the year 1959 when Col. Edwin L. Drake drilled the first well after finding the oil at Titusville, Pennsylvania in U.S.A. Crude oil or petroleum is a mixture of many hydrocarbons. It is believed to have been formed from the decay of dead marine plants and animals at the bottom of sea. The remains of these equatic plants and animals were covered with many layers of the earth. They were subjected to high temperature and pressure. Then after many thousands and millions of years, these plants and animals remains were converted into petroleum and natural gas. The sea water contains the crude oil (mixture of hydrocarbons). The marine plants and animals absorb and drink the sea water respectively. The hydrocarbons in the sea water are therefore incorporated into the tissues and cells of these aquatic organisms. When the aquatic plants and animals die and thier remains decay, the hydrocarbons in their body are converted to petroleum or crude oil after many years. Hydrocarbons are organic compounds composed of only two elements-carbon and hydrogen. The main source of hydrocarbons is crude oil. Petroleum or crude oil is a mixture of many hydrocarbons. The hydrocarbons include the gaseous, liquid and solid alkanes, alkenes, cycloalkanes, aromatic hydrocarbons and others. Natural gas consists mainly of methane.

2.2 Nigerian and World Crude Oil Reserves

Nigeria is among the top ten oil producers of the world. Crude oil was recently found in Otu-Aguleri in Anambra State of Nigeria. On 30th of August 2012, the President, Goodluck Jonathan Ebele Azikiwe commissioned Oleum Petroleum Industry at Otu-Aguleri. He declared Anambra State as the tenth (10th) oil producing state in Nigeria. Since 1956, when crude oil was first discovered at Olobiri, Nigeria has acquired much money from it. Large deposits of oil have been found in the Niger delta-Afam, Abata, Egbema, ughelli, Owaza and Bornu. Crude oil was also found in calabar and Burutu (Delta State). Crude oil is pumped from these areas through pipelines to the oil terminals at Bonny or Forcados for export. The oil is refined outside Nigeria because Nigerian oil refineries are presently not functional. Nigerian oil refineries are located at Port-Harcourt, Warri and Kaduna.

China and Asia are two Asia countries that are rich in crude oil. Asia has about three-fifths of the world's oil reserves but most of it are found in the Middle East. Russia has the next largest oil reserves. Other important oil reserves are found in United states, Canada, Australia, Venezuela, Northern and Western Africa.

2.3 Exploration and Drilling of Crude Oil



Petroleum occurs naturally as a dark, viscous liquid in huge subterranean deposits in many countries of the world such as Nigeria, Egypt, Russia, Kuwait, U.S.A., Saudi Arabia, Iran, Iraq, Venezuela, Mexico, Romania etc. Crude oil is generally found under the rocky strata of the earth's crust, trapped deep below. A certain amount of gas called "natural gas" is associated with the oil.

Coarse grained rock layers or oil-bearing rock layers are explored or located by:

- Aerial photography
- Examining the surface rocks,
- Core drilling and
- Mapping the earth's rock layers.

Geologists do the mapping by exploding dynamite in deep holes in the earth's crust to detect possible oil-bearing areas. Crude oil is extracted by drilling deep holes or wells into the ground when a well reaches any oil-bearing layer, crude oil and natural gas may rise suddenly into the air under pressure. The crude oil then is pumped out and stored in large steel tanks. It is then sent to oil refineries by pipelines or tankers.

2.4 Fractional Distillation of Petroleum and Major Products

Petroleum or crude oil is said to be a mixture of many hydrocarbons. It also contains 1% of sulphur compounds and traces of oxygen and nitrogen compounds. Fractional distillation technique is used in separating mixture of liquids with boiling points that are close together. This method is applied in the purification of crude oil since it is a liquid mixture containing many hydrocarbons with relatively close boiling points. Petroleum contains some very important hydrocarbons like kerosene, petrol, engine oil, gas oil, bitumen, lubricating oil etc. these

products are separated by fractional distillation. The crude oil is heated at the furnace. It vaporizes and enters into the fractionating column where the real separation occurs. The column is cooler at the top than at the bottom. Thus the higher the vapour mixtures rises in the column the faster it condenses. The different components or fractions are taken out from different levels (outlets) of the tower.

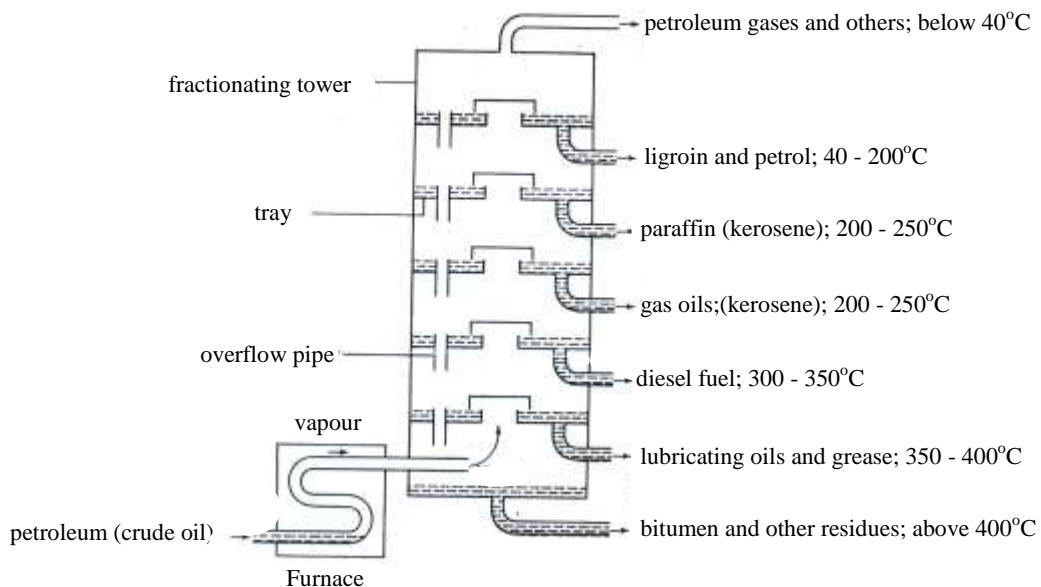


Fig 2.2 Occurrence and drilling of petroleum

Important Fractions of Petroleum and their Uses

1. Gases are mostly used as fuel and for manufacturing of chemicals.
2. Petrol or Gasoline are used as motor fuel and solvents.
3. Kerosene is used as an aeroplane (jet) fuel and as a domestic fuel.
4. Heavy oil is used as fuel in diesel engines.
5. Residue is distilled to obtain greases, wax, oil, bitumen etc. The oil and greases are used as engine lubricants and the wax is used in the manufacture of polishes, candles, waxed paper, body pomades. Bitumen is used for road making and for insulation.

2.5 Petrochemicals

Petroleum or crude oil was originally used as a source of fuels but in 1930 and 1950 in the U.S.A and European countries, it started to be a very important source of many individual chemicals.

The term 'Petrochemicals' is applied to those numerous organic chemicals that are derived from petroleum sources. Petroleum are produced either from natural gas or any of the fractions obtained during the refining of crude oil. Today petroleum has assumed a supreme position as a potent source of a very large number of organic compounds required for industry.

Petrochemicals are chemicals that are chemicals derived from petroleum and natural gas. These chemicals are of great importance to chemical industries. Such essential chemicals include benzene, phenol, methane, ethane, propane, butane, ethane, butane etc. These chemicals are also used in the production of a wide variety of other useful chemicals. Carbon black is easily obtained at the top of a burning kerosene lamp. Carbon black is important as an essential ingredient in the making of tires and in the manufacturing of printer's ink.

Benzene is used in the making of synthetic detergents. The petrochemicals known as butadiene and styrene are used in the making of synthetic rubber. The petrochemicals called polythene is used in making of plastic materials like polythene bags, plastic buckets, balls, spoons, plates, tables, chairs etc. Petrochemicals are used in the manufacture of a lot of other products like body pomade, polishes, disinfectants, insecticides, herbicides, shampoo, hand lotions, body cream, lipsticks, rouge, stain removers and drugs. The formulae and structure of some of the petrochemicals used in the manufacturing of synthetic detergents are as follows.

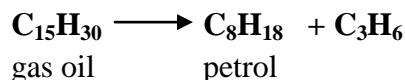
	Common Name	Chemical Formulae	Structural Formulae
1.	Benzene	C_6H_6	
2.	Dodecane	$C_{12}H_{26}$	
3.	Benzene duodecane	$C_6H_5C_{12}H_{26}$	

2.6 Cracking; Polymerization; Reformation; Octane Rating

The first stage in refining of petroleum is the separation of petroleum into various components (fractions) by fractional distillation. All the conversion is the second stage in the refining process. It involves three main processes namely: (i) Cracking (ii) Polymerization and (iii) Reformation.

These processes are designed to convert unwanted fractions into the products in high demand.

Cracking is the breaking of the complex fraction into smaller or simpler fractions that are in high demand. This is brought about by subjecting the complex fraction to high temperature and pressure often in the presence of a catalyst. Eg Gas oil fraction is cracked to yield more petrol.



Two types of cracking are:

- (i) Thermal cracking and (ii) Catalytic cracking

Thermal Cracking is the breaking of the complex fraction into smaller units by heating it to high temperature up to 700°C and high pressure of about atmosphere.

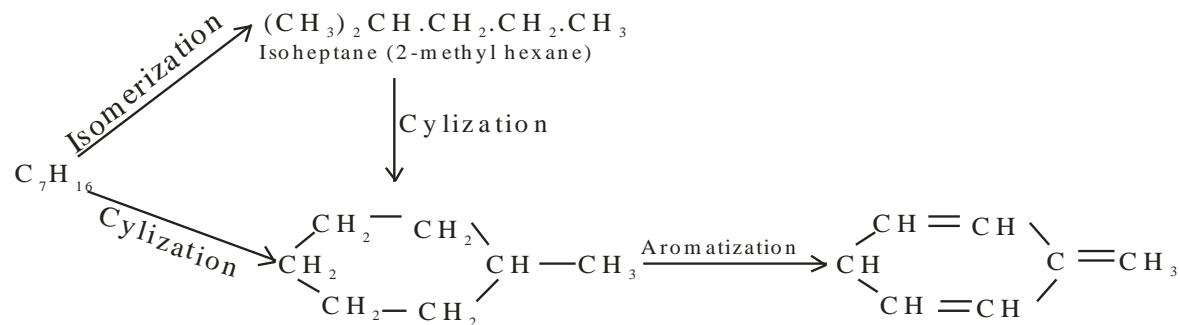
Catalytic Cracking is the breaking of the complex fraction by heating it in the presence of a catalyst at a reduced temperature and pressure.

Examples of such catalyst are: Silica alumina and a mixture of silicon (iv) oxide and aluminum oxide.

Polymerization involves joining small molecules of hydrocarbon gases in the $\text{C}_1 - \text{C}_4$ range to produce molecules in the $\text{C}_5 - \text{C}_{10}$ range. Both cracking and polymerization are used to increase the yield of petrol from crude oil.

Reformation is the rearrangement of carbon atoms in a molecule. Reformation is used to improve the quality of petrol by changing linear or straight chain alkanes to branched chain and cyclic hydrocarbon which yield better or high performance fuel.

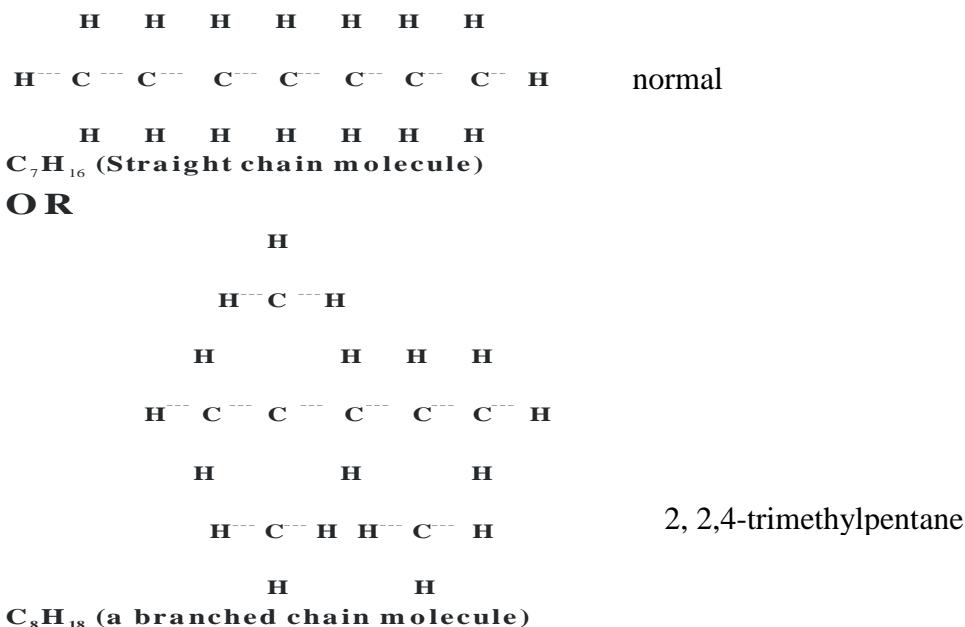
In the process of reforming, the gasoline is heated to a temperature of about 600°C for a very short time, in presence of a platinum catalyst. The pressure (28-50 atmosphere) is maintained in such a way that molecules do not crack but are reformed. The straight chain hydrocarbons may be converted to branched chain hydrocarbons by **Isomerisation** or **Alkylation**; alkanes to cycloalkanes and cycloalkanes to aromatic hydrocarbons. The overall effect is increase of octane number of the fuel and the process is known as reforming.



Reforming carried out in the presence of hydrogen is known as hydroforming

Octane rating: Petrol is a hydrocarbon containing carbon in the C₅-C₁₀ range.

The major compounds in petrol are C₇ (Heptane) and C₈ (Octane) hydrocarbon. These hydrocarbons may occur as straight chain or branched chain molecules as in



Octane rating of petrol is a measure of its performance in an internal combustion engine. Straight chain alkanes do not burn smoothly in internal combustion engines. They burn rapidly and unevenly, and generate an explosion which disturbs the up and down movement of pistons in the engine. This causes a strange sound, usually referred to as 'engine knock'. Octane number is the percentage of 2,2,4-trimethyl pentane in a petrol or fuel. For example, the petrol with an octane number of 90, means that the fuel is made up of 90% 2,2,4-trimethyl pentane and 10% normal heptanes. A fuel which contains only the straight chain hydrocarbon (n-heptane) is a very poor quality fuel and its octane number is zero. Thus higher grade petros are said to have a high octane rating or number. Motor vehicles run well on petrol with a high octane rating of about 80-90.

Ethyl fluid and methyl fluid are added to petrol as an antiknock compounds. They raise the octane rating of petrol. Ethyl fluid is a mixture of tetraethyl lead (iv) (TEL), Pb (C₂H₅)₄ (60%), Dibromo ethane, (26%), Dichloro ethane (9%) and a red dye (2%).

Methyl fluid is a mixture of tetramethyl lead (iv) (TML), Pb (CH₃)₄ (60%), Dibromomethane (26%), Dichloro-methane (9%) and the red dye (2%). 1-3ml of these mixtures (Ethyl fluid or methyl fluid) is added to a gallon of petrol to increase its antiknock properties.

Exercises

1. Define these terms
 - a. Petroleum
 - b. Petrol
 - c. Petrochemical
2. Name three petrochemicals we studied in this chapter and say their usefulness in chemical industries.
3. List three important fractions of petroleum. Write the uses of each of them.
4. What two Asia countries are rich in crude oil?
5. What are the four ways of locating oil-bearing rock layer?
6. Define the terms;
 - a. Thermal cracking
 - b. Catalytic cracking
 - c. Reformation
 - d. Octane number
7. a. What causes engine knock in moving vehicle?
b. Name two antiknock compounds we studied in this chapter

CHAPTER 3

METALS AND THEIR COMPOUNDS

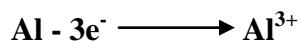
Metals are a class of elements that ionize by loss of their valence electrons. They are mainly found in groups 1-3 in the periodic table. They combine with non-metals electrovalently to form ionic or electrovalent compounds.

3.1 Physical Properties of Metals

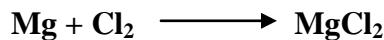
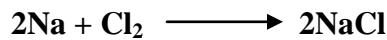
1. Metals have high densities.
2. They have characteristic lustre and can be polished.
3. With the exception of mercury which is liquid at room temperature, all metals are solid at room temperature with high melting and boiling points.
4. They have high tensile strength and as a result can withstand stress and strain.
5. Metals are malleable and ductile and as a result can be hammered into sheets or drawn into wire.
6. Metals are good conductors of heat and electricity.
7. Metals give off sound when hit, which differs from metal to metal. This means that they are sonorous.

Chemical properties of metals

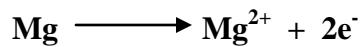
1. Metals are electropositive.
2. They ionize by the loss of their valence electron or electrons to form cations.



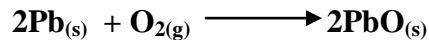
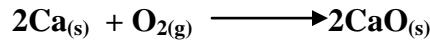
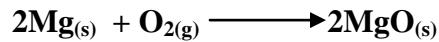
3. They form ionic or electrovalent compounds.



4. Metals are reducing agents or electron donor.

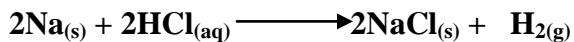
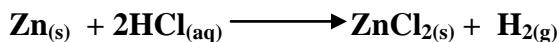


5. Almost all metals will combine with oxygen to form metallic oxide.



Noble metals like gold, silver and platinum do not easily combine with oxygen. Metallic oxides are bases. Na_2O and K_2O will dissolve readily in water to form alkalis. CaO is sparingly soluble in water while the oxides of the other metals are insoluble.

6. Metals higher than hydrogen in the electro-chemical series can displace hydrogen from acids.

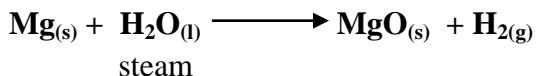


7. In metals, atoms are held together by metallic bonds.

8. Metals react with water in a number of ways depending on the position of the metal in the electrochemical series. K, Na and Ca react with cold water to liberate hydrogen

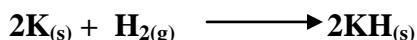


Mg, Zn and Fe react with steam to liberate hydrogen



Cu, Ag, Au and Hg can neither react with cold nor hot water.

9. More electropositive metals combine with hydrogen to form ionic hydrides. Hydride is a binary compound of hydrogen.



3.2 Principles of Extraction of Metals

I. Extraction of metals is a reduction process. K, Na, Ca, Mg and Al which are very reactive metals, naturally occur as oxides or salts. They occur mainly as sulphides and trioxocarbonate (iv). Zn, Fe, Sn and Pb which are moderately reactive occur as oxides or sulphides. The least reactive metals (noble metals) are mainly found in an uncombined form. These metallic compounds are ionic in nature. During extraction, the reducing agent donates the electrons and the metallic ion takes up the appropriate number of electron and it's converted to free atom. The process of electron gain is known as reduction. The more common methods used in the extraction of metal are:

1. Electrolytic reduction
2. Chemical reduction and
3. Thermal reduction.

K
Na
Ca
Mg
Al

Very reactive or more electropositive metals.
Occur as stable oxides and salts. Extracted by electrolysis.

Zn
Fe
Sn
Pb

Moderately reactive metals. Extracted by chemical reduction with C or CO.

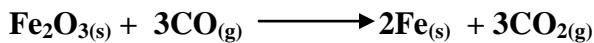
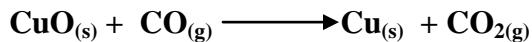
Cu
Hg
Ag
Au

Least reactive or noble metals. May be found as free elements.

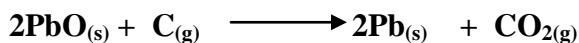
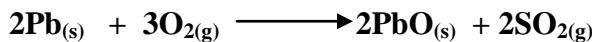
1. Electrolytic Reduction: is used in the extraction of very reactive or more electropositive metals. During Electrolysis, reduction and oxidation reactions take place in the cathode and anode respectively. The cathode supplies the electrons which are gained by the metallic ions in the electrolyte. The metallic ions are converted to their free atoms and are deposited at the cathode.

2. **Chemical Reduction** is used in the extraction of less electropositive (moderately reactive) metals. The two methods used are:

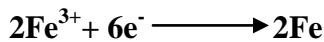
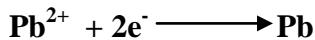
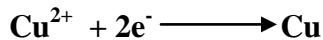
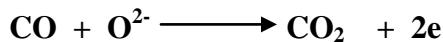
- i. Reducing metallic oxides to their metals with Coke or Carbon (ii)oxide.



- ii. Oxidation of the metallic sulphide to the oxide followed by reduction of the oxide by Coke.



In these reactions, CO reacts with O^{2-} ion in the oxide and releases the electrons which reduce the metallic ions.



3. Thermal Reduction: In this method, metals are obtained by heating the metallic sulphide in the presence of air



Summary of Extraction Methods

Metal	Most Common Ore	Method Of Extraction
K Na	Chlorides	Electrolysis of fused hydroxide or chlorides
Ca Mg	Chlorides and Carbonates	Electrolysis of fused Chlorides
Al	Oxides	Electrolysis of oxide
Zn Fe Sn	Oxides, Carbonates and sulphides	Roasting of carbonates and sulphide to form oxides, Reduction of oxides by carbon or carbon(ii)oxide.
Pb Cu Hg	Sulphide	Roasting in air Heating in air
Ag Au	Free elements	Mined as free elements

Classification of metals:

Elements found under groups I, II and III (in the periodic table) are called metals. Metals are therefore classified under groups I, II and III. Read groups I, II and III elements in chapter 6 of part I in this book. Group I elements are known as the alkali metals. They are Lithium, Sodium, Potassium, Caesium and Francium. Group II elements are called the alkaline earth metals. They are Beryllium, magnesium, calcium, strontium and barium. Group III elements are boron, aluminum, gallium etc.

3.3 Alkali Metals

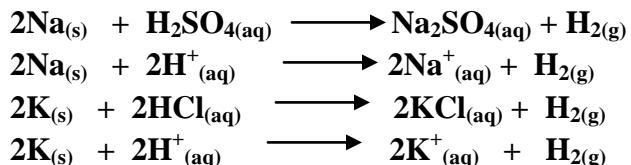
The alkali metals are Lithium, Sodium, Potassium, Rubidium and Caesium. (See chapter 6 of part 1).

Physical Properties of Alkali Metals

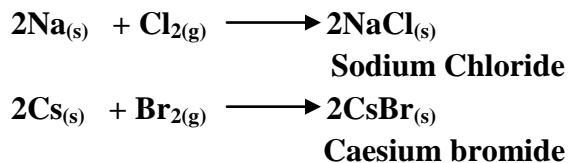
Property	Li	Na	K	Rb	Cs
Atomic number	3	11	19	39	55
Outer electronic configuration	1s ² 2s ¹	2p ⁶ 3s ¹	3p ⁶ 4s ¹	4p ⁶ 5s ¹	5p ⁶ 6s ¹
First Ionization Energy (KJ/mol)	520	494	418	403	374
Ionic radius A°	0.60	0.95	1.33	1.48	1.69
Atomic radius A°	1.33	1.57	2.03	2.16	2.35
Electronegativity	1.15	1.0	0.9	0.9	0.85
Melting point (K)	454	371	336	312	302
Boiling point (K)	1615	1156	1033	959	942

Some chemical properties of alkali metals:

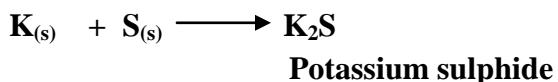
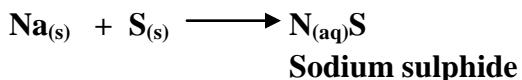
- a. They react with dilute acids liberating hydrogen gas.



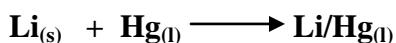
- b. They react with halogen to form halides

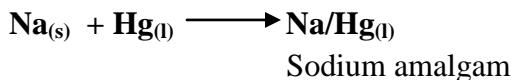


- c. They react with sulphur forming their corresponding sulphides.



- d. They dissolve in mercury to form their corresponding amalgams.





Read other chemical properties of alkali metals under group I elements in chapter 6 of Part I.

3.4 Extraction of Sodium

Sodium and potassium are extracted by electrolysis of their fused (molten) chlorides. Sodium is extracted industrially by Down's process in which fused or molten sodium chloride is electrolysed.

As the melting point of sodium chloride is high (800°C), Calcium Chloride is added to lower the melting point to 600°C . Downs' cell is made up of Carbon anode and iron cathode.

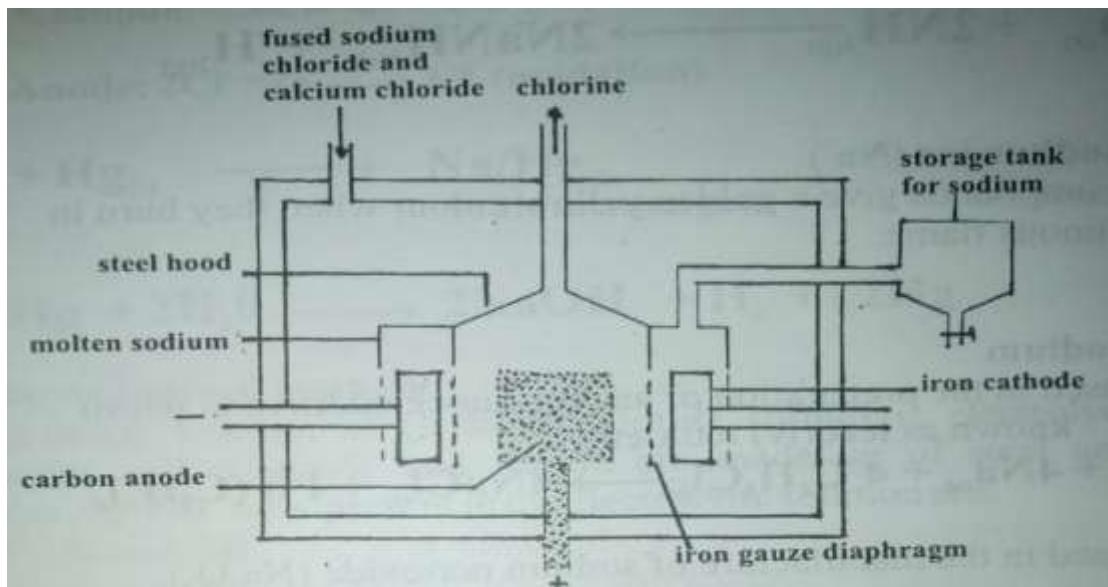


Fig. 3.1 Downs' Cell

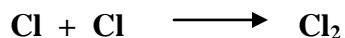
Chemistry of the reaction



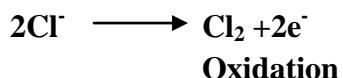
At the Cathode



At the Anode



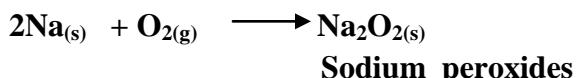
OR



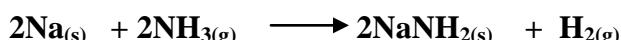
The sodium is tapped and chlorine is a valuable by-product of the process.

Some properties of Na

1. The density of Na is 0.97 g cm^{-3}
2. It burns with a golden yellow colour
3. It burns in an excess supply of air to form sodium peroxides



4. Sodium reacts with ammonia to liberate hydrogen.

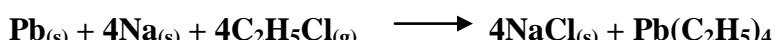


Test for sodium ion (Na^+)

Sodium compounds give a golden yellow colour when they burn in a non-luminous flame.

Uses of sodium

- a. It is used in the preparation of an anti-knock additive in petrol known as lead(iv) tetraethyl.



- b. It is used in the manufacture of sodium peroxide (Na_2O_2), Sodium cyanides (NaCN), and Sodaamide (NaNH_2).

3.5 Sodium Hydroxide (NaOH)

NaOH can be prepared in the laboratory by the Gossage's method. Calcium hydroxide commonly known as slacked lime is heated with 10% sodium trioxocarbonate (iv) solution. Calcium trioxocarbonate (iv) formed is insoluble. It is then filtered off. The filtrate which is a solution of sodium hydroxide is then concentrated by evaporation.

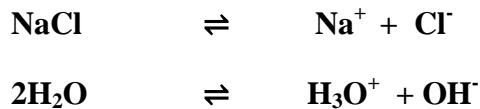


Two electrolytic methods of preparing Caustic Soda (NaOH) are by:

- (i) Mercury cathode cell method
- (ii) Diaphragm cell method

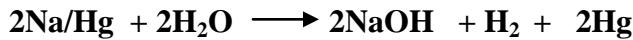
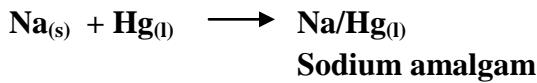
(i) Mercury cathode cell method

Caustic soda, NaOH is prepared industrially by the electrolysis of concentrated brine. The ions present in the concentrated brine are: Na^+ , Cl^- , H^+ and OH^- . At the mercury cathode, Na^+ is discharged instead of hydrogen ion (H^+). Chlorine is liberated at the carbon Anode.



At the Cathode: $\text{Na}^+ + \text{e}^- \longrightarrow \text{Na}$ (reduction)

At the Anode: $2\text{Cl}^- - 2\text{e}^- \longrightarrow \text{Cl}$ (oxidation)



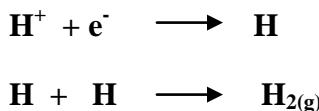
ii. Diaphragm cell method

Concentrated solution of (sodium chloride (brine) is electrolyzed in the diaphragm cell. Cathode and Anode are made up of steel and carbon respectively. The ions present in the electrolytic solution are:

Na^+ , Cl^- , H^+ and OH^- .

At the steel Cathode:

H^+ ions are preferentially discharged



At the Carbon Anode

Cl^- ions are preferentially discharged because of its high concentration.



The removal of H^+ and Cl^- ions leave Na^+ and OH^- (NaOH) in the solution. The solution is concentrated by evaporation to obtain the pellets of NaOH .

Physical properties of sodium hydroxide

1. NaOH solution is a strong alkali and turns red litmus paper to blue
2. It is a white deliquescent crystalline solid.
3. It is very soluble in water. Its dissolution in water is accompanied by the evolution of heat.

Chemical properties of sodium hydroxide

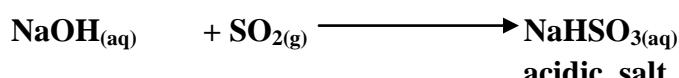
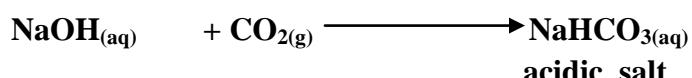
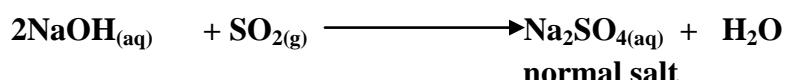
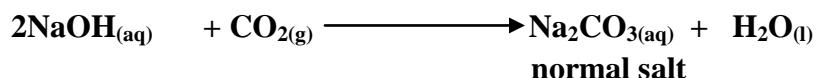
1. It dissolves or reacts with ethanol.



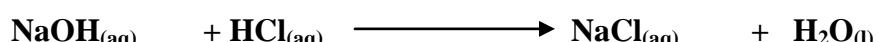
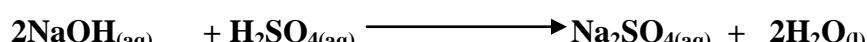
2. It ionizes completely in water.



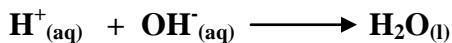
3. It reacts with the acidic oxides like carbon(iv) oxide and sulphur(iv) oxide to form their corresponding acidic or normal salts.



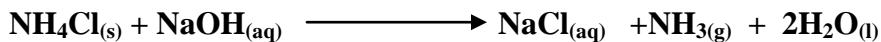
4. As a base, it neutralizes all acids to form salt and water only



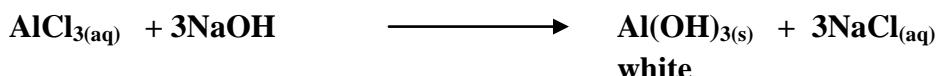
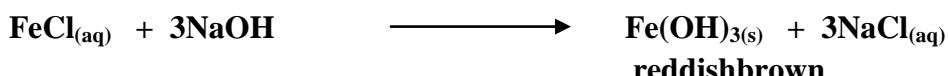
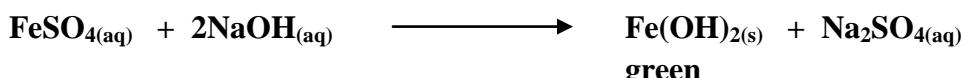
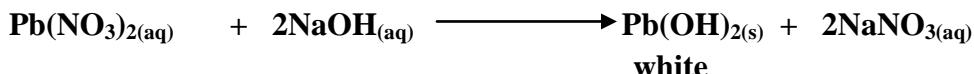
Ionically



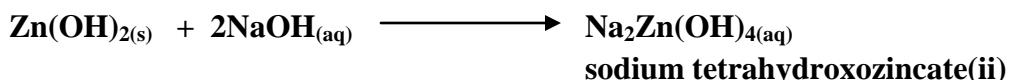
5. NaOH solution reacts with any ammonium salt when warmed to liberate ammonia gas (NH_3). The gas turns moist red litmus paper blue. This is used as a selective test for ammonium salts.

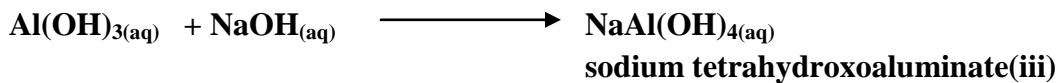


6. Dilute, concentrated and molten NaOH attack or react with glass to form sodium trioxosilicate(iv). This is the reason why glass stoppers become glue to reagent bottles containing the NaOH solution. It also explains why glass volumetric flask stoppers and burette taps tend to stick after the solution has been used in these equipments. To preserve these instruments, they should be thoroughly washed after NaOH solution has been used in them.
7. NaOH solution precipitates insoluble metallic hydroxides from solutions of their metallic salts. The cations in the metallic salts should not be Na^+ , K^+ or NH_4^+ because their hydroxides are soluble in water.



The hydroxides of Aluminum, Zinc and lead are soluble in excess NaOH solution (because they are amphoteric) forming soluble complex salts.



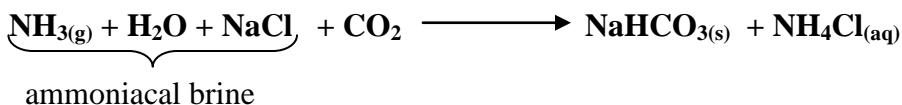


Uses of sodium hydroxide

1. It is used in the science laboratory as a strong alkali for analytical and precipitating purpose.
2. It is used in the manufacture of: soap, rayon (artificial silk) and paper.
3. It is used for absorbing carbon (iv)oxide.

3.6 Sodium Trioxocarbonate(iv), Na_2CO_3 : Sodium Hydrogen Trioxocarbonate(iv), NaHCO_3 and Sodium Chloride.

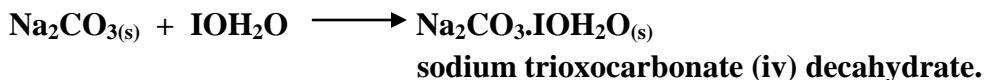
The most common industrial process for the production of Na_2CO_3 is the solvay process in which CO_2 is forced up the tower through the falling ammoniacal brine (mixture of NH_3 and brine).The tower contains baffle plates which slow down the flow of the liquid and present large surface area for the reaction to occur. The ammoniacal brine combine with the CO_2 to form NaHCO_3 and NH_4Cl .



The NaHCO_3 is only slightly soluble in water. It is then filtered, washed and decomposed by heat if Na_2CO_3 is needed.



To obtain the washing soda, the soda ash can then be dissolved in hot water and recrystallized.



Uses of Na_2CO_3

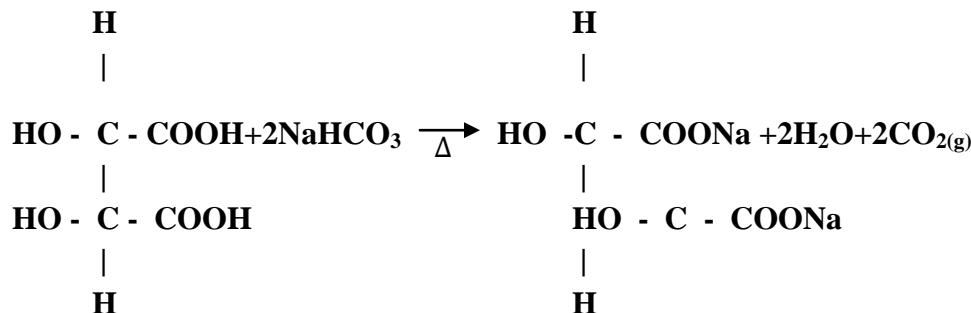
1. It is used in the manufacture of water glass.
2. It is used as water softener
3. It is used in the making of detergent.

Uses of Na_2HCO_3

1. It is used in fire extinguishers.

- It is used in the making of baking powder.

When a mixture of Na_2HCO_3 and $(\text{HOCHCOOH})_2$ or tartaric acid is heated, CO_2 gas liberated causes bread and cake to rise



2, 3-dihydroxy butandioic acid
or tartaric acid

Sodium chloride NaCl

It occurs naturally as a solid in the form of rock salt deposits in Poland and other parts of the world. It is also found in a very small quantity in sea water and sea weeds.

NaCl is produced or obtained by

- Mining it directly in the solid form
- Evaporation of sea water.
- The neutralization reaction between $\text{NaOH}_{(aq)}$ and $\text{HCl}_{(aq)}$



- Direct combination reaction between Sodium and Chlorine gas.



Uses of NaCl.

- It is used in the making of: sodium, chlorine, hydrochloric acid, bleach, sodium hydroxide, sodium trioxocarbonate(iv) decahydrate i.e. washing soda and sodium hydrogen trioxocarbonate(iv) i.e. baking powder.
- It is used as a ‘salting out’ agent during the manufacture of soap.

3.7 Alkaline Earth Metals

Group 11A elements of the periodic table are known as the alkaline earth metals. They consists of beryllium, magnesium, calcium, strontium, barium and radium.

(Read Chapter 6 of Part 1).

Physical Properties Of Alkaline Earth Metals

Property	Be	Mg	Ca	Sr	Ba
Atomic number	4	12	20	38	56
Outer electronic configuration	$1s^2 2s^2$	$2p^6 3s^2$	$3p^6 4s^2$	$4p^6 5s^2$	$5p^6 6s^2$
First Ionization Energy(KJ/mol)	900	738	590	549	502
Ionic radius A°	0.31	0.65	0.99	1.13	1.35
Metallic radius A°	1.12	1.60	1.97	2.15	2.22
Atomic Volume (Cm ³ /mol)	4.9	14.0	26	33.7	39.3
Electronegativity	1.5	1.25	1.05	1.0	0.95
Melting point (K)	1553	923	1124	1062	993
Boiling point (K)	3243	1380	1760	1607	1413

Chemical Properties of Alkaline Earth Metals:

Read group (II) elements under chapter 6 of Part I.

3.8 Extraction of Calcium

Calcium is extracted by the electrolysis of fused calcium chloride. Some calcium fluoride is added to the fused calcium chloride to lower the melting point from 850°C to about 650°C. The mixture is placed in a crucible lined with graphite anode. The cathode consists of an iron rod.

NOTE: In this process, Calcium hydroxide cannot be used as an electrolyte because on heating, it decomposes to calcium oxide and water. On further heating, the CaO formed becomes white hot. It does not melt since it is one of the refractive substance known.

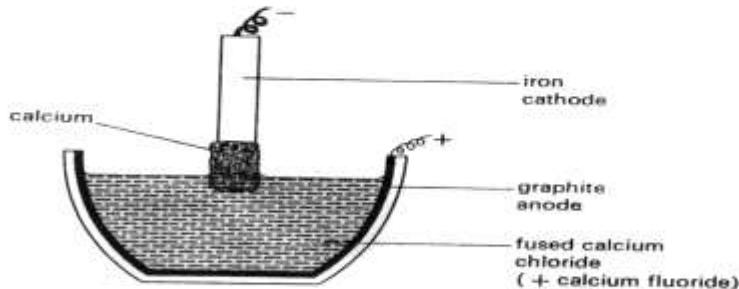


Fig. 3.2 Extraction of Calcium

Chemistry of the reaction

The ions present are Ca^{2+} and Cl^-

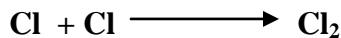
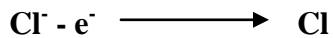


At the Cathode: The calcium ion gains two electrons and it's reduced to calcium atom.



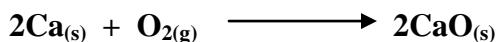
At the Anode: The chloride ion loses an electron to form chlorine atom. Two chlorine atoms pair up to form chlorine molecule which is liberated at the anode.

liberated at the anode.



Chemical properties of calcium

- It is a good conductor of heat and electricity. Calcium metal is very reactive and electropositive. It rapidly tarnishes and losses its metallic lustre when it is exposed to air. This is caused by the formation of a white film of calcium oxide (quick lime) on the surface of the metal.



The $\text{CaO}_{(s)}$ absorbs the atmospheric moisture to form Calcium hydroxide.



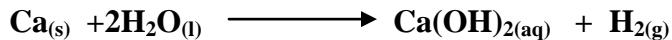
The Ca(OH)_2 combines with the atmospheric carbon(iv)oxide to form calcium trioxocarbonate(iv)



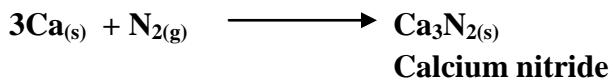
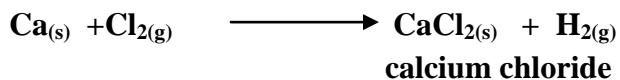
When heated in air, calcium burns with a brick-red flame forming calcium oxide.



2. Calcium combines very slowly with cold water to form calcium hydroxide and hydrogen. The reaction is rapid with hot water.



3. On heating, calcium reacts and forms binary compounds with these non-metals-nitrogen, chlorine, hydrogen and sulphur.



4. Calcium reacts vigorously with dilute acids liberating hydrogen

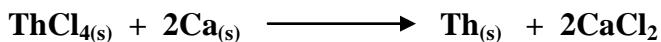


5. If ammonia is passed over heated calcium, the two react to liberate hydrogen gas.



Uses of calcium

- It is used in the manufacture of CaO, Ca(OH)₂, and CaSO₄.H₂O.
- It is also used in the making of calcium fluoride and calcium hydride.
- It is used in the extraction of uranium and thorium.



- It is used as a deoxidant.

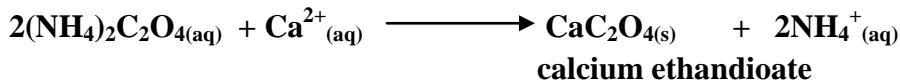
Test for Ca²⁺

- Calcium compounds give a brick-red colour to a non-luminous flame.

2. When few drops of NaOH is added to a solution of a calcium compound, a white precipitate which is insoluble in excess NaOH solution, indicates the presence of calcium ions.



3. When few drops of ammonium ethanedioate solution is added to a solution of an unknown salt, the formation of white precipitate which is soluble in dilute hydrochloric acid but insoluble in ethanoic acid, indicates the presence of calcium ions.



4. When ammonium trioxocarbonate (iv) solution is added to a solution of an unknown salt, the formation of a white precipitate indicates the presence of calcium ions.



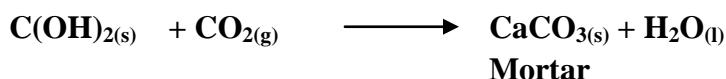
Cement and Mortar

Cement is obtained when a mixture of limestone and clay is heated to a high temperature. Clay comprises of aluminum trioxosilicate(iv).



Concrete is obtained when gravel, sand and water are mixed with cement.

Mortar is obtained as follows:



3.9 Group (III) Metals

Group (III) elements in the periodic table are boron, aluminum, gallium, indium and thallium. Read group (III) elements under Chapter 6 of part I. Aluminium is the third most abundant element found in the earth's crust.

The principle ores of aluminium are:

1. Bauxite (hydrated oxide, $\text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$)
2. Cryolite (sodium hexaflouroaluminate (iii), Na_2AlF_6).
3. Mica or Feldspar ($\text{K}_2\text{Al}_2\text{Si}_6\text{O}_{16}$).

4. Kaoline or China Clay ($\text{Al}_2\text{Si}_2\text{O}_7 \cdot 2\text{H}_2\text{O}$)
5. Corundum (Al_2O_3)
6. Spinel (aluminate, $\text{MgO} \cdot \text{Al}_2\text{O}_3$)

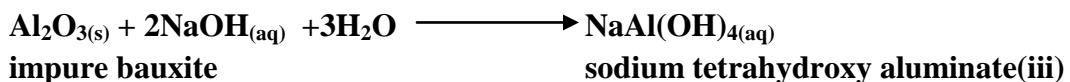
3.10 Extraction of Aluminum

Aluminum is extracted from bauxide by electrolysis. The extraction proceeds in two stages;

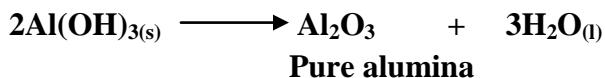
1. Purification of the Bauxite.
2. The electrolysis of the pure alumina.

A. Purification of the bauxite

The impure bauxite is heated with sodium hydroxide solution to form soluble sodium tetrahydroxy aluminate (iii).



The impurities in the ore which are iron (iii) oxide and trioxosilicate (iv) compounds are not soluble in the alkali. They are therefore filtered off as a sludge. Aluminum hydroxide crystals are then added to filtrate, NaAl(OH)_4 solution to induce the precipitation of Aluminum hydroxide.



B. Electrolysis of the pure alumina

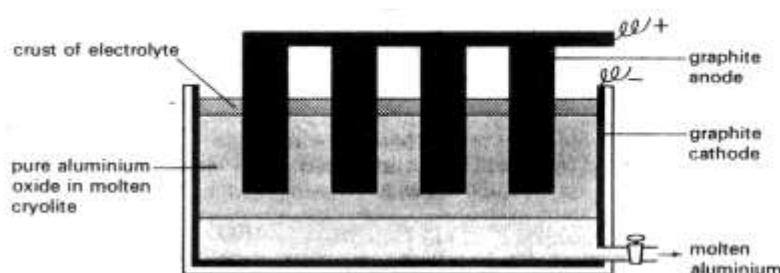


Fig. 3.3 Extraction of aluminum

An electrolytic cell used in this electrolysis is as shown above. Both cathode and anode are made of graphite. A solution of pure aluminum oxide (alumina) in molten cryolite, Na_3AlFe , is poured into the cell and the temperature of the system is maintained at about 100°C . Aluminum is therefore extracted by the electrolysis of pure Al_2O_3 in molten cryolite. During the electrolysis, oxygen ions migrate to the anode where they are oxidized to oxygen atoms. The atoms pair up to

form the oxygen molecule which is liberated at the anode. Aluminum ions gains three electrons each from the cathode and are reduced to aluminum atoms. The molten aluminum collects at the bottom (cathode) of the cell from where it is tapped off. The solid crust formed at the surface of the electrolyte prevents the molten aluminum from being oxidized by the atmospheric oxygen. The O₂ formed during the electrolysis cannot oxidize the molten Al because it reacts with the anode and is given off as carbon(iv) oxide.

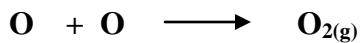
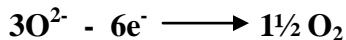
Chemistry of the reaction

The ions present are: Al₂O₃ → 2Al³⁺ + 3O²⁻

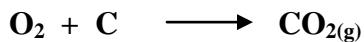
At the Cathode



At the anode:



The O₂ reacts with the anode



Physical properties of AL

Property	Al	Property	Al
Atomic number	13	Metallic radius	1.43A°
Outer electronic		Common oxidation state	+3
Configuration	3s ² 3p ¹	Melting point	660°C
Atomic mass	27	Boiling point	2450°C
Ionic radius	M3+: 0.50A°		
Density	2.70gCm ⁻³		

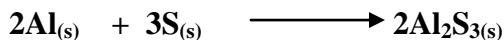
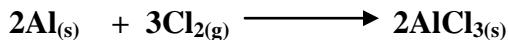
Al is a good conductor of heat and electricity.

It is ductile and malleable

It is silvery-white in colour.

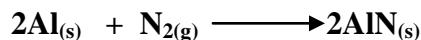
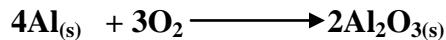
Chemical properties of aluminum

- Al combines directly with carbon, sulphur, nitrogen, phosphorus and halogens to form their corresponding binary compound. The reaction is exothermic.



- On exposure of Al to moist air, a thin layer of Al_2O_3 is formed on its surface. This prevents further attack of the metal by the atmospheric oxygen, water or steam under normal condition. Aluminum sheets are therefore said to be corrosion free, but if the temperature of the aluminum sheet is raised to about 800°C , the remaining aluminum metal will combine with the air forming $\text{Al}_2\text{O}_{3(\text{s})}$ and $\text{AlN}_{3(\text{s})}$

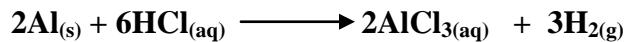
remaining aluminum metal will combine with the air forming $\text{Al}_2\text{O}_{3(\text{s})}$ and $\text{AlN}_{3(\text{s})}$



- Aluminum metal when mixed with the liquid mercury forms an alloy known as aluminum amalgam. An alloy is a mixture of metals. The amalgam reacts with moist air or water to form aluminum hydroxide.



- Like other metals, aluminum displaces hydrogen from hydrochloric acid when it reacts with it.



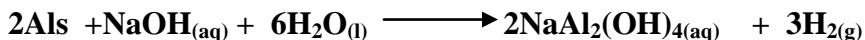
It reacts slowly with dil. HCl but rapidly with conc. HCl.

- Al does not react with either the dil. Or concentrated trioxonitrate (v) acid, HNO_3 . The acid oxidizes the surfaces of the metal forming a thin protective layer of aluminum oxide which renders the aluminum metal passive.

- Al does not react with dil. H_2SO_4 . Concentrated H_2SO_4 oxidizes the metal liberating sulphur(iv) oxide.



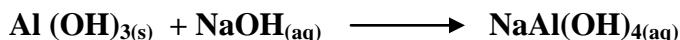
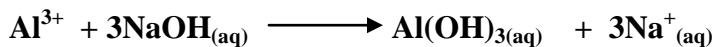
- Al reacts with or dissolves in both sodium and potassium hydroxide to sodium or potassium tetrahydroxy aluminate (iii).



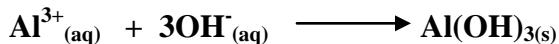
Washing soda is not used in washing aluminum containers because it is alkaline.

Test for Al³⁺

	Test	Observation	Inference
1.	Solution of an unknown compound + NaOH soln.	White gelatinous ppt Soluble in excess NaOH is observed.	Al ³⁺ suspected



	Test	Observation	Inference
1.	Solution of an unknown compound + NH ₄ OH	White gelatinous ppt insoluble in excess NH ₄ OH is observed.	Al ³⁺ present



Alloys of aluminum

Some important alloys of aluminum are:

- i. Magnalium which is a mixture of Al and Mg.
- ii. Aluminum bronze which is a mixture of Cu and Al
- iii. Duralumin which is a mixture of Al, Cu, Mg and Mn.

These alloys are strong, hard, light and they are resistant to corrosion. They are used in the making of cars, ships and aircrafts.

Uses of Al

- 1. It is used in making foil for wrapping foods.
- 2. It is used in making domestic materials like spoons, pans, buckets, kettles, plates etc.
- 3. Aluminum powder suspended in oil is used as paint for mirror and car.

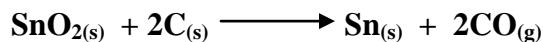
3.11 Group (iv) Metals

Group (IV) elements in the periodic table are: Carbon, Silicon, Germanium, Tin and Lead. Our attention here should be concentrated on Tin (Sn) and Lead. The principal ore of tin is

Tin(iv)oxide, SnO_2 commonly known as cassiterite . It is mined in Jos, located in Plateau State of Nigeria.

Extraction of Tin

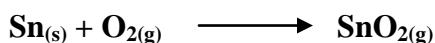
After mining the tin ore, it is ground into powder and washed with water. The impurities like sulphur and arsenic oxide are then removed from the ore by roasting it in the air. Pure tin (iv) oxide obtained is then reduced to tin by heating it with coke or anthracite.



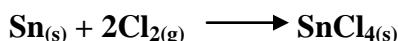
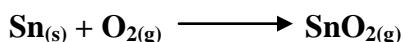
The three allotropic forms of tin are (1) Grey tin (2) White tin and (2) Brittle tin. Allotropy of solids is caused by differences in structure. These three allotropes are reversible at different temperatures. Grey tin is formed at a temperature below 13°C . White tin is stable at a temperature between 13°C and 160°C . Brittle tin is formed at temperatures above 160°C .

Properties of tin

1. It is a white lustrous metal
2. The transition temperature for the three allotropes of tin is 1302°C .
3. It does not react with either cold or hot water.
4. When tin is heated to a high temperature, it reacts with air to form SnO_2 .



5. When heated, tin react with non-metals.



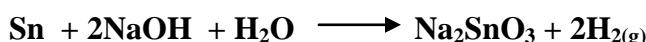
6. It reacts slowly with dilute H_2SO_4



It reacts with hot concentrated H_2SO_4 to liberate $\text{SO}_{2(\text{g})}$



7. Tin reacts with hot sodium hydroxide liberating hydrogen.



Uses of tin

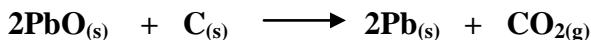
1. It is used in the coating or electroplating of steel containers e.g. milk cans.
2. Tin foil is used for protecting food.
3. It is used in the making of alloys, e.g. Bronze is an alloy of copper and tin.
Soft solder is an alloy of Pb and Sn.
Type metal is an alloy of Pb, Sb and Sn.

Lead: The principal ores of lead are PbS, PbSO₄ and PbCO₃.

Extraction: The ore, PbS is roasted in air to obtain lead(ii)oxide.

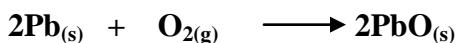


The PbO is then heated with some Coke, and limestone. The limestone reacts with the impurities and forms slag which is removed from the tap. PbO is reduced to lead, Pb by Coke.



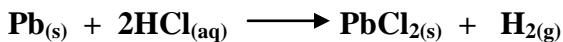
Properties of lead

1. It is a greyish white lustrous solid
2. It is a good conductor of heat and electricity.
3. Lead is attacked by moist air forming a thin layer coating of PbO, Pb(OH)₂ and PbCO₃ on the surface of the metal.
4. When Pb is heated in air, it reacts with oxygen to form yellowish solid, lead(ii) oxide.



5. Lead does not dissolve in either dilute HCl or dilute H₂SO₄ due to the formation of an insoluble thin layer coat of lead(ii)chloride or lead tetraoxosulphate(vi). The coat prevents the metal from further attack.

Lead reacts with hot concentrated HCl and H₂SO₄ to liberate hydrogen and sulphur(iv) oxide respectively.



Pb reacts with both dilute and concentrated $\text{HNO}_{3(\text{aq})}$ acid to liberate nitrogen (ii) oxide and nitrogen (iv) oxide respectively.

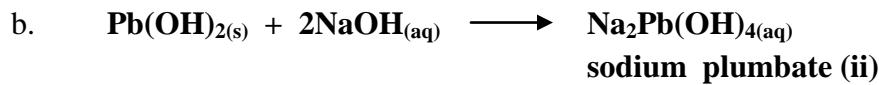


Uses of lead

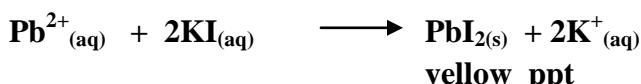
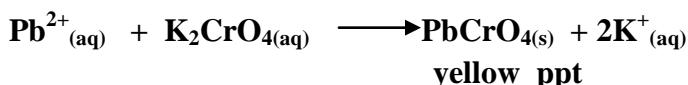
1. It is used in making water and gas pipes.
2. It is used in the making of lead accumulators
3. It is used in the manufacture of Lead (iv) tetraethyl, and PbO , PbCO_3 and PbSO_4 .
4. It is used in the making of alloys like: Type metal, soft solder and pewter.

Test for lead (ii) ions

1. Test	Observation	Inference
Soln of salt + a few drops of NaOH soln	White precipitate which dissolves in excess NaOH Solution is formed.	Pb^{2+} Suspected



2. Test	Observation	Inference
Soln of salt + K_2CrO_4 soln	A yellow precipitate is formed	Pb^{2+} is present
3. Soln of salt +KI Soln	A yellow precipitate is formed	Pb^{2+} is present



Lead tetraoxochromate (vi), K_2CrO_4 and potassium iodide, KI are yellow solids

4.	Test	Observation	Inference
	Soln of salt + a few drops of Conc HCl	White precipitate is formed. ppt dissolves when warmed but reappears on cooling	Pb ²⁺ confirmed



Exercises

- 1a. What is a metal? (b) Write three examples of metal.
2. Write two physical and two chemical properties of metal.
3. Briefly explain 'Metallic bonding'.
4. Name three common methods employed in the extraction of metal.
Explain only one of them.
5. Write the chemistry of the reactions that occur during the extraction of sodium using the Down's cell. What are the cathode and anode made of?
6. Write one physical property of Sodium hydroxide.
7. Write three chemical properties of sodium hydroxide.
8. What are the two methods of obtaining Sodium Chloride?
9. What are the two important uses of sodium Chloride?
10. Write the chemistry of the reactions that occur during the extraction of sodium using the Down's cell. What are the cathode and anode made of?
11. Write one physical property of Sodium hydroxide.
12. Write three chemical properties of sodium hydroxide.
13. What are the two methods of obtaining Sodium Chloride?
14. What are the two important uses of sodium Chloride?
15. What are the two principal ores of: (a) tin? (b) lead?
16. Use a balanced equations only to show the reduction of: (a) tin ore to tin (b) lead ore to lead

17. Write two physical properties of (a) tin (b) lead.
18. Write two chemical properties of (a) tin (b) lead.

CHAPTER 4

TRANSITION METALS

Transition metals (elements) are the elements which have partially filled d orbitals. Electronic configuration of the first transition series are as follows:

Transition metals	Electronic Configuration	
Scandium, Sc	(Ar structure)	3d14s2
Titanium, Ti	„	3d24s2
Vanadium, V	„	3d34s2
Chromium, Cr	„	3d54s1
Manganese, Mn	„	3d54s2
Iron, Fe	„	3d64s2
Cobalt	„	3d74s2
Nickel, Ni	„	3d84s2
Copper, Cu	„	3d104s1
Zinc, Zn	„	3d104s2

Fig 4.1

4.1 Properties of Transition Metals

1. Transition elements or metals have variable oxidation states as shown below:

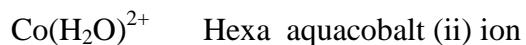
Element	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
Maximum oxidation numbers	3+	+4	+5	+6	+7	+6	+4	+4	+2	+2
	+3	+4	+3	+6	+3	+3	+3	+2	+1	
		+2	+2	+4	+2	+2	+2			
				+3						
				+2						

Transition metals form coloured compound. Coloured ions are formed by metals with incompletely filled d-orbitals. Metals with completely empty d-orbitals and those with completely filled d-orbitals form colourless ions. Sandium and zinc do not form coloured ion because they cannot absorb colours from white light. The colours of some compounds of transition metals are shown below.

Ion	Colours	Compounds in which ions exist
Mn ⁴⁺	black	MnO ₂
Mn ⁷⁺	intense purple	KMnO ₄
Fe ²⁺	black	FeO
Fe ³⁺	reddish-brown	FeO ₂ O ₃
Fe ²⁺	green	FeSO ₄
Fe ³⁺	from colourless to yellow	Fe ₂ (SO ₄) ₃ .YH ₂ O
Cr ⁶⁺	yellow	K ₂ CrO ₄
Cr ⁶⁺	deep orange	K ₂ Cr ₂ O ₇
Cr ²⁺	blue	CoCl ₂ .6H ₂ O
Cu ²⁺	black	CuO
Cu ²⁺	blue	CuSO ₄ .7H ₂ O
Cu ²⁺	green	CuCl ₂ .2H ₂ O
Cu ¹⁺	orange yellow	Cu ₂ O

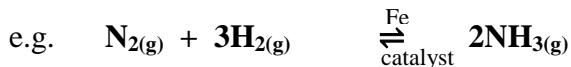
Fig. 4.3

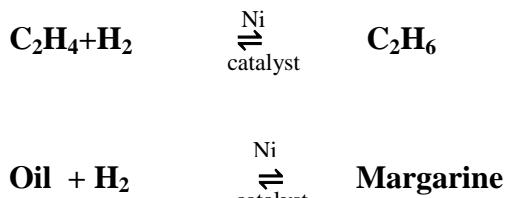
3. Transition metals form complex ions or compounds or complexes. Complex compounds are often referred to as the co-ordination compounds. A complex compound consists of a central atom to which atoms, ions or molecules are surrounded. The surrounding ions or molecules are known as ligands. Some examples of complexes are:



See the structure od some of these complexes in chapter 13.3 of part 1

4. They act as catalyst because they can change their oxidation number with ease.





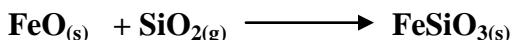
4.2 Copper, Cu: Copper is a transition metal. It has a completely filled 3d orbitals ($3d^{10}$) . Copper is one of the less reactive metals. It is below hydrogen in the electrochemical series. It can be found as a free metal because it is stable in air. Like silver and gold, copper is also known as a noble metal. The principal ore of copper is: Copper pyrites (CuFeS_2). Other ores of it are cuprite (Cu_2O), Malachite ($\text{CuCO}_3\cdot\text{Cu}(\text{OH})_2$), Bornite ($\text{Cu}_2\text{S}\cdot\text{CuS}\cdot\text{FeS}$) and Chalcocite (Cu_2S).

Extraction of copper: Copper is extracted from copper pyrites, CuFeS_2 . The two processes involved in this extraction are (i) the concentration of the ore by froth floatation to remove the impurities. (ii) Roasting the concentrated ore in air. CuFeS_2 is crushed and mixed with water. Air is blown into the mixture. The denser impurities then sink to the bottom of the container. The froth floating on the surface the concentrated ore which is separated out. The ore is then roasted in the air to oxidize the iron and some sulphur in the ore to iron (ii) oxide and sulphur(iv)oxide respectively.

contains the concentrated ore which is separated out. The ore is then roasted in the air to oxidize the iron and some sulphur in the ore to iron (ii) oxide and sulphur (iv) oxide respectively.



Silicon(iv) oxide, is added to the remaining solid mixtures and is heated. FeO is converted to iron(ii) trioxosilicate (iv) which float as slag on top of the molten Copper(i) sulphide, Cu_2S .



FeSiO_3 is removed as slag and Cu_2S is then reduced to Cu by heating it with a limited supply of air.



The copper formed is an impure substance.

Pure copper is obtained from the impure form by electrolysis. CuSO_4 is the electrolyte. The impure Cu is used as the anode while pure copper is used as the cathode. During electrolysis, the impure Cu enters into the solution and ionizes. It migrates to the cathode where it gains electrons to form Cu atom, which is deposited at the cathode as a pure Cu.

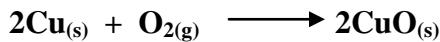
Properties of Cu

- It has a melting point of 1083°C .

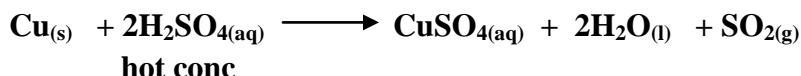
2. Its relative density is 8.95
3. It is a reddish brown metal
4. It is a good conductor of heat and electricity.
5. It is very malleable and ductile
6. It has a high tensile strength.
7. If Cu is exposed to moist air, it is covered with a thin layer green coating of basic copper (ii) tetraoxosulphate (vi),



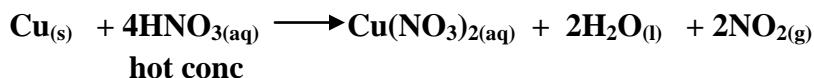
8. Cu is oxidized to black copper(ii)oxide when it is heated in air



9. Cu cannot displace hydrogen from dilute acids since it is lower than hydrogen in the electrochemical series.
10. It cannot react with dilute H_2SO_4 and HCl acids. However Cu reacts with the hot concentrated oxidizing acids like H_2SO_4 and HNO_3 .



Cu reacts with hot concentrated HNO_3 acid to liberate nitrogen(iv)oxide



Cu reacts with dilute HNO_3 to liberate nitrogen(ii)oxide



Uses of copper

1. It is used in making electric wires.
2. It is used in the making of alloys like bronze, brass etc.
3. It is used in roofing and plumbing.

Alloys of copper: The alloys of copper are:

- i. Duralumin which is a mixture of Al, Cu, Mg and Mn.
- ii. Bronze which is a mixture of Cu and Sn.
- iii. Brass which is a mixture of Cu and Zn.
- iv. Cupronickels which is a mixture of Cu and Ni.

Test for copper ion, Cu^{2+}

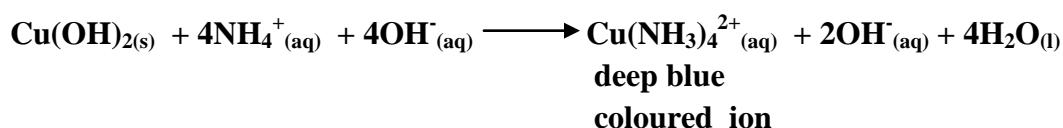
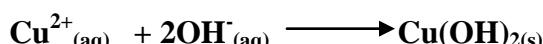
Common hydrated Copper(ii) salts are greenish or bluish in colour. They are soluble in water to form blue solution.

Copper(ii) salts impart bluish-green colour to a non-huminous flame

1. Test	Observation	Inference
Solution of salt + a few drops of NaOH soln	A blue gelationus precipitate which is insoluble in excess NaOH solution is formed.	Cu^{2+} present



2. Test	Observation	Inference
Solution of salt + a few drops of Aqueous NH_3	A blue precipitate which dissolved in excess aqueous NH_3 to give a blue solution is observed.	Cu^{2+} present and confirmed



4.3 Iron

Iron is the second most abundant metal in nature after aluminium. It commonly occurs in these combined forms known as the iron ores.

Haematite or iron (iii) oxide Fe_2O_3

Limonite or hydrated iron (iii) oxide, $\text{Fe}_2\text{O}_3 \cdot 3\text{H}_2\text{O}$

Magnetite - Iodestone or iron (ii) diironoxide, Fe_2O_4

Iron pyrite - firestone or iron (ii) persulphide, Fe_3S_2 .

Siderite or iron (ii) trioxocarbonate (iv), FeCO_3 .

Fe is a constituent of the pigments chlorophyll and hemoglobin.

Extraction of Iron

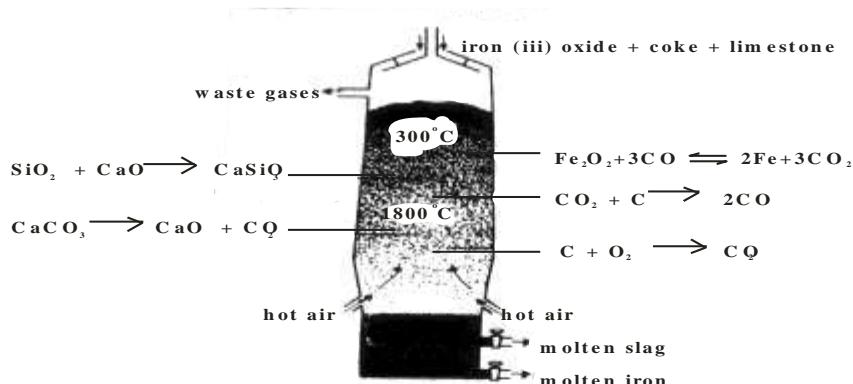


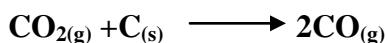
Fig 4.4 Blast Furnace

Blast furnace shown above is the apparatus used in the extraction of iron. The principal iron ore used in the extraction of iron is haematite. Others ores can also be used. If the ore is not an iron (iii) oxide, it is first of all roasted in air to convert it to iron (iii) oxide. Iron (iii) oxides, coke and limestone are introduced into the blast furnace from its top. A blast of hot air is also introduced into the furnace through small pipes called tuyeres. These reactions occur within the furnace.

At the lower part of the furnace; coke combines with the oxygen present in the hot air to form carbon (ii) oxide, CO. The reaction is exothermic.

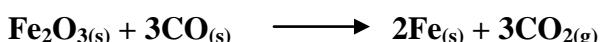


As the quantity of O_2 in the hot air decreases, CO_2 is then reduced to CO by the coke.



In the upper part of the furnace:

The gas CO moves up and reacts with iron (iii) oxide reducing it to iron.



At high temperature, limestone decomposes to form calcium oxide



The CaO reacts with silicon (iv) oxide (impurity) to form calcium trioxosilicate (iv) known as slag.



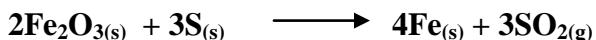
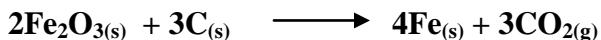
The Molten iron formed sinks to the Bottom of the blast furnace while the molten slag floats on top of it. Slag is then removed from the top as shown in Fig 7.1. The iron formed is hard and brittle and is known as pig iron.

Kinds of iron

Pig iron: Is the iron obtained directly from the blast furnace. It is hard and brittle. It contains about 2-5% carbon and other elements as impurities. The elements include sulphur, phosphorus, silicon and manganese.

Cast iron: Is got from pig iron by re-melting it with some scrap iron and then cooled in mould. The percentage of impurities it contained is lower than that in the pig iron. It is used in the making of cookers, stove, radiators and base of bunsen burners.

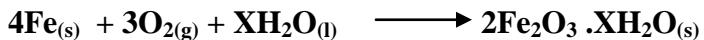
Wrought iron: Is the purest form of iron. It contains about 0.2% carbon. It is obtained from cast iron by the removal of some of its impurities (C and S). It is done by reacting the cast iron with Fe_2O_3 which oxidises the impurities to CO_2 and SO_2 respectively.



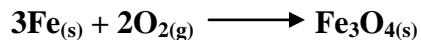
Properties of iron

1. It is a white lustrous solid with a relative density of 7.9. Its melting point is 1530°C . It is a good conductor of heat and electricity and can be magnetized with ease. It is a very malleable and ductile metal with high tensil strength.

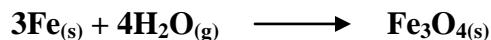
2. Iron reacts with moist air to form reddish brown hydrated iron (iii) oxide, $\text{Fe}_2\text{O}_3 \cdot \text{XH}_2\text{O}_{(s)}$ known as rust.



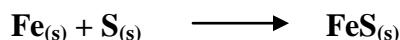
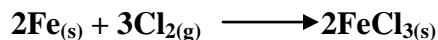
3. Iron burns in air when heated to a high temperature to form iron (ii) diiron (iii) oxide, Fe_3O_4 .



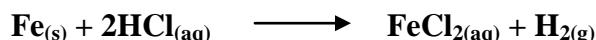
4. Red hot iron filings react with steam to form iron (ii) diiron (iii) oxide.



5. When heated, iron reacts with non-metals such as chlorine, sulphur, phosphorus and carbon. It does not react with nitrogen.



6. Iron dissolves in dilute HCl and H₂SO₄ acids to liberate hydrogen gas.



7. Iron reacts with dilute HNO₃ to form many products such as NH₄NO₃, Fe(NO₃)₂, NH₃ and oxides of nitrogen. Concentrated HNO₃ acid does not react with iron instead it renders it passive. This is due to the formation of the layer of protective oxide. This explains why concentrated HNO₃ acid is stored in iron containers.

Uses of iron

Cast iron is used in the making of pipes, radiators, retort stands etc.

Wrought iron is malleable. It is used in the making of nails, iron rods, chains, iron sheets, ornamental gates and agricultural implements.

Alloys: The alloys of iron includes steels and alloy steels.

1. Steel or Ordinary steel is a mixture of iron and about 0.1-1.5% carbon.
2. Alloy steels are formed when other elements are added to the molten steel to impart different desired properties.

Steel is classified into mild, medium and hard depending on its carbon content and the heat treatment it receives. Mild steel contains 0.1 -0.25% carbon, medium steel contains 0.25-0.6% carbon while hard steel contains 0.6-1.5% carbon.

Alloy steels include the following:

1. Stainless steel which is a mixture of steel, chromium and nickel. It is used in making scissors, cutlery etc.
2. Magnetic steel which is a mixture of steel and cobalt. It is used in the making of permanent magnets.

3. Hard steel-a mixture of steel and tungsten is used for making cutting and drilling tools.
4. Manganese steel is a mixture of steel and manganese.
5. Spring steel is a mixture of steel and silicon.
6. Molybdenum steel is a mixture of steel and molybdenum.

Spring steels is a mixture of iron and silicon. Manganese steels and molybdenum steels are also alloy steels.

4.4 Rusting of Iron

When iron is exposed to moist air, it rusts and forms reddish brown hydrated iron (iii) oxide, $\text{Fe}_2\text{O}_3 \cdot \text{XH}_2\text{O}$ known as rust. Rust does not protect the metal from further reaction because it is non-coherent and permeable.

Methods of preventing rusting

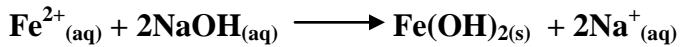
1. Painting and greasing are methods of preventing rusting. Oil paint or grease is applied on the surface of iron to prevent it from coming in contact with air and water.
2. Electroplating is a method of coating the surface of one metal with another. The purpose of electroplating is to prevent rusting and to improve the appearance of some metals like iron and steel. Metals used in electroplating include silver, chromium, tin and zinc. In silver plating a spoon, the anode is a silver rod, the cathod is the spoon while the electrolyte is a soluble silver salt eg AgNO_3 solution. As the current passes through the cell, the anode dissolves in the solution and forms ions. The ions migrate to the cathode and are discharged and deposited as a thin layer coat on the object to be plated.
3. Galvanizing is the electrolytic method of spraying molten zinc on the surface of iron and steel or by dipping the iron and steel into the molten zinc. Zinc is mainly used for galvanizing iron and steel, but galvanized containers are not used for storing food because they can cause zinc poisoning.

Test for iron (ii), Fe^{2+}

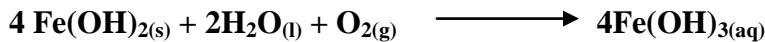
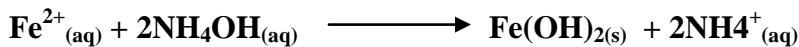
Common soluble salts of iron (ii) dissolve to form green solutions.

1.	Test	Observation	Inference
	Solution of salt + a few drops of NaOH soln	A dirty green gelatinous precipitate which is insoluble in excess NaOH soln is observed. The green	Fe^{2+} is present

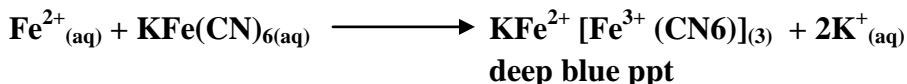
ppt gradually turns brown
on exposure to air.



2.	Test	Observation	Inference
	Solution of salt + a few drops of NH ₄ OH soln	A dirty green gelatinous precipitate which is insoluble in excess NH ₄ OH is observed. The green Ppt turns brown on exposure to air.	Fe ²⁺ is present



3.	Test	Observation	Inference
	Solution of salt + a few drops of potassium Hexacyanoferrate (iii) soln	A deep blue precipitate is observed.	Fe ²⁺ is Confirmed

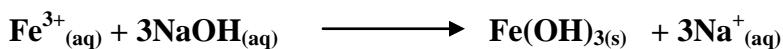


4.	Test	Observation	Inference
	Solution of salt + potassium Thiocyanate soln.	No visible reaction is observed.	Fe ²⁺ is Present

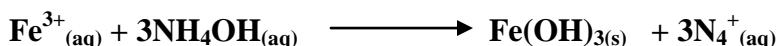
Test for iron (iii), Fe³⁺

Common soluble salts of iron (iii) dissolve to form yellow or brown solutions.

1.	Test	Observation	Inference
	Solution of salt + a few drops of NaOH soln.	A reddish brown gelatinous precipitate, which is insoluble in excess NaOH	Fe ³⁺ is confirmed



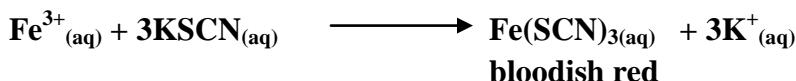
2.	Test	Observation	Inference
	Solution of salt + a few drops ppt, of NH ₄ OH	A reddish brown gelatinous which is insoluble in excess NaOH solution is observed.	Fe ²⁺ is present



3.	Test	Observation	Inference
	Solution of salt + a few drops of potassium hexacyano- ferrate (ii)	A blue precipitate is observed.	Fe ³⁺ is present



4.	Test	Observation	Inference
	Solution of salt + a few drops of potassium thiocyanate solution	A bloodish solution is observed.	Fe ³⁺ ion confirmed.



Exercises

- What is transition metal?
- Mention three properties of transition metals.
- Briefly describe how an impure copper is purified.
- Write two chemical properties of copper.
- Name four alloys of copper.
- What are the components of each?
- Name three ores of iron.
- Name three kinds of iron.

- a. Dilute HCl acid on iron.
- 6. Use chemical equations only to show the effect of:
- b. Dilute H₂SO₄ acid, on iron.
- 7a. Write the chemical formular of rust.
- b. What is the colour of rust?
- c. How can rusting be prevented?

CHAPTER 5

ETHICAL, LEGAL AND SOCIAL ISSUES ON CHEMICAL WASTES:

A lot of benefits is derived from industrialization. It helps a nation to make maximum profit. It helps the younger generation to get employment. On the other hand, industrialization has adversely affected the life of living things (man, animal and plant) in various environment. The chemical wastes produced from the chemical industries are introduced into the environments (air, land and water). The quantity or concentration of these chemical wastes sometimes become so high that the living things in the particular environment are no longer comfortable. High quantity of the industrial wastes in the environment pollutes it. The pollution of the environment then, causes emigration and death of living organisms in the environment. Some of these chemical wastes or industrial pollutants includes; gases eg sulphur (iv) oxide, carbon (ii) oxide, carbon (iv) oxide, hydrogen sulphide, nitrogen (iv) oxide, nitrogen (ii) oxide and others.



Fig. 5.1 Evolution of smoke in an industry

Rapid advances in science and technology have produced enormous benefits but have also created undesirable hazardous side-effects that impact human health and the environment. The toxicological sciences strive to understand and evaluate the health and environmental effects of chemical and physical agents. The impact of this expanding body of science on society has grown enormously in the last 100 years, and with that have arisen corresponding financial, legal, and individual implications. Despite the increased scientific data and understanding, decision making has become more difficult and complex. It is thus increasingly important to consider the ethical, legal, and social issues that confront toxicologists, public health professionals, and decision makers. ***Toxic means poisonous. Toxicology is the scientific study of poisons or toxic chemicals or substances. Toxicologists are the scientists that specialized in the scientific study of poisons or toxic chemicals.***

The fundamental principles that an ethical toxicologist should consider can be summarized as: 1) dignity, which includes respect for the autonomy of human and animal subjects; 2) veracity, an

adherence to transparency and presentation of all the facts so all parties can discover the truth; 3) justice, which includes an equitable distribution of the costs, hazards, and gains; 4) integrity, an honest and forthright approach; 5) responsibility, an acknowledgment of accountability to all parties involved; and 6) sustainability, consideration that actions can be maintained over a long period of time .

Beyond these basic principles it is important to have a vision of environmental health that is grounded in ethical considerations.

5.1 Historical Perspective

Looking back, it is easy to see the beginnings of an ethical framework for decision making in the Greek physician Hippocrates (460-377 BC), who studied the effects of food, occupation, and climate on causation of disease and is credited with the basic medical tenet of "do no harm." Bernardo Ramazzini (1633 - 1714), an Italian physician, examined the health hazards of chemicals, dust, metals, and other agents encountered by workers in 52 occupations, which he documented in his book "Diseases of Workers".



Fig. 5.2 Rachel Carson

Aldo Leopold , considered by many to be America's first bioethicist, summarized ethical responsibilities in a simple statement in 1949. "A thing is right when it tends to preserve the integrity, stability, and beauty of the biotic community. It is wrong when it tends otherwise." It can be extrapolated from this ethical statement that exposing people, particularly children, to harmful agents robs them of their "integrity, stability, and beauty," indeed their potential, and is therefore wrong. Health, ecological, and ethical concerns about chemical exposures were highlighted by Rachel Carson, in Silent Spring, first published in 1962. Carson sounded one of the first alarms about the effects of environmental contaminants and catalyzed numerous regulatory changes related to chemical use.

One relevant principle adopted by toxicologists states that: "To prevent harm is the best method of environmental protection and, when knowledge is limited, apply a precautionary approach."

5.2 Legal Issues

There is a wide range of laws and regulations that shape the role of toxicology in society. One of the first laws dealing with toxicology, passed in 82 BCE by the Roman Emperor Sulla, was intended to deter intentional poisonings because women were poisoning men to acquire their wealth. In 1880, food poisonings spurred Peter Collier, chief chemist, U.S. Department of Agriculture, to recommend passage of a national food and drug law. In 1938, the Federal Food, Drug, and Cosmetic Act was adopted following an incident in which Elixir Sulfanilamide, containing the poisonous solvent diethylene glycol, killed 107 people, many of whom were children. The need to control chemical contamination was recognized in the 1976 when the U.S. Congress passed the Toxic Substances Control Act (TSCA) to "prevent unreasonable risks of injury to health or the environment associated with the manufacture, processing, distribution in commerce, use, or disposal of chemical substances." TSCA became largely ineffective following court decisions and there is now an effort to pass chemical policy reform legislation. Meanwhile Europe has moved forward with REACH, Registration Evaluation and Authorization of Chemicals, a system that requires testing and evaluation of chemicals before their introduction into commerce.

In Nigeria, many bodies or agencies are responsible for the control of drugs and Chemicals. They examine the testing and evaluation of chemicals before their introduction into commerce. These bodies include: Chemical Society of Nigeria (CSN), Institute of Chartered Chemists of Nigeria (ICCON), National Agency for Food and Drug Administration and Control (NAFDAC), and National Drug Law Enforcement Agency (NDLEA). The functions of the two drug control agencies (NAFDAC and NDLEA) are: They regulate control the manufacture, importation, distribution, advertisement, sale and use of drugs, medical devices, food, cosmetics, chemicals and all bottled drinks including our popular "Pure water".

5.3 Social Considerations

Toxicologists and public health professionals play an important role in society in protecting and promoting public health. There has been an extra focus on ethical and social issues related to children's health. The U.S. Society of Toxicology code of ethics indicates that toxicologists should be thoughtful public health advocates. While seldom explicitly stated, professional codes of ethics such as those for SOT are often based on the following social responsibilities: (1) to share and use knowledge, (2) to promote the health and well being of children, and (3) to maintain the right of all species to reach and maintain their full potential.



Fig. 5.3 A child playing with an animal

5.4 Ethical Considerations

A toxicologist is also concerned with issues of integrity and honesty in the conduct and interpretation of toxicological studies. It is important to examine and acknowledge conflicts of interest. Toxicology associations as well as state and federal agencies, nonprofits, and universities have statements and guidelines on conflict of interest and disclosure. In addition, toxicologists must adhere to rules and regulations regarding the use of animals and humans in scientific studies. The conduct of studies involving humans has a rich history that has become increasingly well defined and regulated to ensure adequate knowledge and consent of subjects involved.

5.5 Ethical and Environmental Background to the Management of Waste

The development and welfare of modern societies depend to a large extent upon the contribution of technology and industrial processes, such as the generation and widespread use of electricity. These processes are, in general, associated with the production of wastes, some of which are unavoidable, unrecyclable and hazardous. Such wastes require careful management to ensure adequate protection of humans and the environment. The timescales over which such protection is required can extend, in the case of wastes containing toxic chemical elements or long-lived radioactive isotopes, well beyond the lifespans of current or forthcoming generations, i.e., many thousands of years into the future. Hence there is an ethical imperative to care about future generations and to act in such a way as to preserve, as much as possible, their options to enjoy and benefit from the Earth's resources. Such a concern for the protection of human health and the environment in a developing world has been illustrated by the concept of "sustainable development" put forward by the World Commission on Environment and Development, "the Brundtland Commission", in 1987 . This concept, which is principally an ethical one, was defined as "satisfying the needs of the present, without compromising the ability of future generations to meet their own needs".

The concept of "sustainable development" was chosen as the main theme of the United Nations Conference on Environment and Development in Rio de Janeiro in 1992, and was therefore

extensively discussed. It is appropriate that the principles of this concept be applied to complex environmental issues such as the ones resulting from the production of potentially harmful wastes. Current environmental protection policies are increasingly concerned with issues of a global nature related to long-term consequences of, for example, ozone depletion and climate changes. In this context, an evolving environmental consciousness, coupled with the emergence of strong ethical concerns, indicate the importance attached to morally correct human conduct . This trend should contribute to the adoption of public policies integrating both technical and ethical considerations to maximise the benefits and limit the potential adverse effects of industrial development now and in the future. It is therefore welcome that ethical issues are currently being integrated into the environmental debate.

In the management of wastes having a long-term potential for harm, interest focusses on two classes of ethical concerns: Intergenerational equity and Intragenerational equity.

The first is the achievement of "intergenerational equity" by choosing technologies and strategies which minimize the resource and risk burdens passed to future generations by the current generations which produce the wastes. It is a fact of life that each generation leaves a heritage to posterity, involving a mix of burdens and benefits, and that today's decisions may foreclose options or open new horizons for the future. This is unavoidable, but our actions and decisions will be more acceptable if appropriate degrees of equity or justice are respected, and we do not unduly restrict the freedom of choice of future generations. The objective is to manage the wastes in such a way that potential future impacts are kept at a level that is acceptable both ethically and in terms of safety.

The second concern is the achievement of "intragenerational equity" and in particular an ethical approach to the handling, within current generations, of questions of resource allocation and of public involvement in the decision-making process. The need for public involvement and its importance in making key decisions, such as the timing of waste disposal actions, will be emphasised. When considering resource allocation, risks from radioactive wastes must be kept in perspective with competing projects in the area of human health and environmental protection.

Consideration of these concerns leads to a set of principles to be used as a guide in making ethical choices about waste management strategy:

- the liabilities of waste management should be considered when undertaking new projects;
- those who generate the wastes should take responsibility, and provide the resources, for the management of these materials in a way which will not impose undue burdens on future generations;
- wastes should be managed in a way that secures an acceptable level of protection for human health and the environment, and affords to future generations at least the level of

safety which is acceptable today; there seems to be no ethical basis for discounting future health and environmental damage risks;

- a waste management strategy should not be based on a presumption of a stable societal structure for the indefinite future, nor of technological advance; rather it should aim at bequeathing a passively safe situation which places no reliance on active institutional controls.

Clearly, the development of waste management policy and plans should be openly discussed with representatives from all concerned sections of society.

Summarily, it can be said that: the purest of ethical behavior and decision making requires the thoughtful development and articulation of fundamental principles upon which to base any action. The ethical toxicologist must consider and integrate basic ethical principles into the decision making process. This approach moves beyond what is legally required: an ethical approach requires ongoing discussion and considerations as the toxicological sciences and society evolve. Toxicologists must not only be familiar with the rules and regulations regarding the ethical conduct of research, but also with the underlying ethical principles. The challenge is to move beyond a purely legal adherence to the rules but toward an ethical approach grounded in carefully considered and articulated ethical principles that drive the responsible conduct and application of research in modern societies.

5.6 Pollution

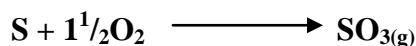
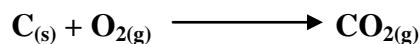
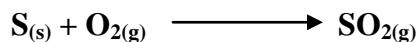
Pollution is a phenomenon that occurs when a pollutant is introduced into the atmosphere. If a pollutant is introduced into air, it causes air pollution. If it is introduced into the water, it causes water pollution and if it is introduced into the soil, it causes soil pollution.

A Pollutant is substance or a chemical which when introduced into the atmosphere causes atmospheric pollution. It can also be defined as a substance which when introduced into an environment, makes the environment unfit or unhealthy for the life of living things in that environment. Pollutants can be classified into three. These are: air pollutants, water pollutant and land or soil pollutants. Air pollutants are airborne materials such as dust, smoke and soot, gases such as sulphur (iv) oxide (SO_2), Carbon (ii) oxide (CO), oxides of nitrogen and hydrogen sulphide (H_2S). Others include insecticides, herbicides, etc.

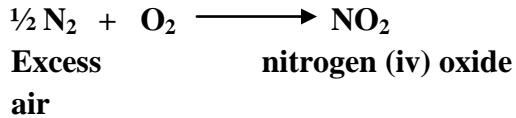
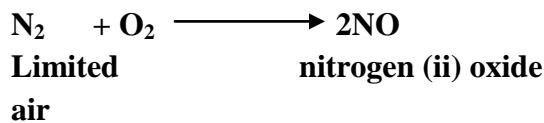
Water pollutants are too many. This is because; some air and soil pollutants can dissolve either directly in or indirectly in a body of water to cause water pollution. This means that some air and soil pollutants can also be listed as water pollutants. Some examples of water pollutants are SO_2 , CO , nitrogen (iv) oxide (NO_2), nitrogen (ii) oxide (NO), fertilizers, heat, mercury, lead, synthetic organic compounds such as insecticides, fungicides, herbicides and oil.

Land or soil pollutants includes plastics, tyres, old cars, solid waste from kitchens and industries such as papers, cans vegetable materials, soap and detergents, fertilizers, faeces and dungs etc.

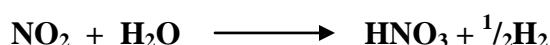
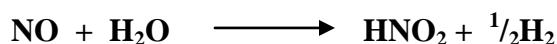
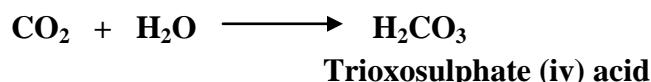
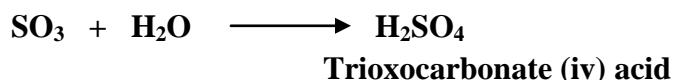
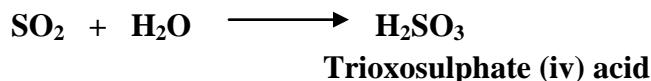
5.7 Air Pollution: any substance present in the air other than oxygen, nitrogen, noble gases (ie Ar, He, Kr etc) are impurities and causes pollution of air. Most pollution is caused by burning of fuels such as oil, coal or gas which contain small amounts of sulphur. When the fuel is burnt, apart from the release of smoke and soot into the air, the sulphur impurity and coal are converted into the gases sulphur (iv) oxide and carbon (iv) oxide respectively



Air pollution is also caused by oxides of nitrogen released from chemical industries and car exhausts. Nitrogen (ii) and nitrogen (iv) oxides are produced when nitrogen gas is heated to a high temperature in a limited and in excess supply of air.

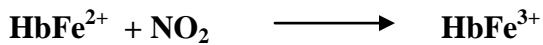


These gases: SO₂, sulphur (vi) oxide (SO₃), CO₂, NO and NO₂ enter into the atmosphere and dissolves in rain to form acid rain.

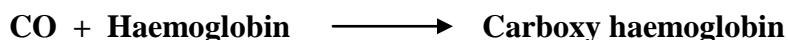


Acid rain then falls into rivers and the land and building causing damage to life and property. Acid rain is harmful to living things such as trees, life in rivers and lakes etc. it also causes damage to buildings and works of sculpture outside by washing away beautiful paint used in

decorating them. It makes metals to corrode faster. Dioxonitrate (iii) ion (NO_2) from $\text{HNO}_{2(\text{aq})}$ can oxidize Fe^{2+} in haemoglobin to F^{3+} specie. Such haemoglobin is ineffective as oxygen carrier.



Carbon (ii) oxide (CO) is formed when the hydrocarbon fuel in a car is not burnt completely to form carbon (iv) oxide. CO is extremely poisonous. It readily combines with haemoglobin to form a very stable complex compound known as carboxy haemoglobin which does not carry oxygen to the body cells. The victim – very quickly becomes starved of oxygen. He suffers from suffocation and eventually dies.



Trtraethyl – lead (iv) is added to petrol to improve its performance as a fuel. As the petrol burns, lead compounds are released into the atmosphere. Lead compounds are harmful because they damage nerve cells. They can cause brain damage in children. Noise is an air pollutant.

5.8 Water Pollution: some of the water pollutants are (i) dissolved impurities from the atmosphere e.g. CO_2 , SO_2 etc (ii) sewage and garbage (iii) industrial waste, (iv) Agricultural waste, treatment of water before use is vital to our health.

Industrial Waste: many factories pour chemical waste directly into rivers. These waste include acids, alkalis, cyanides, mercury compounds, phenols. Many of these are harmful both to fish and to humans who use untreated water from these rivers.

Agricultural Waste: Many substance used as fertilizers or pesticides in Agriculture are washed away by rain. The amount of such substances in some water sources can pose a health problem both to fish and human beings

Water collected from asbestos roof is not good for drinking. Asbestos fibres scratch the walls of the air passages including the walls of lungs. It leads to lung cancer.

Oil spillage and heating up of rivers and lakes by sun can cause the death of aquatic life.

Control of Pollution: It can be controlled by:

- i. Making people aware of the danger of pollution
- ii. Treating sewage in sewage plants before emptying the effluent into rivers and seas.
- iii. Treating industrial waste before emptying them into rivers and seas.
- iv. The government should make it a law that chemical industries should be sited very far away from the residential places.

Exercises

1. What are the two classes of ethical concerns where interest is focused in the management of wastes having a long-term potential for harm?
2. What are the six fundamental principles that an ethical toxicologist should consider?
3. What are the full meaning of these abbreviations?
 - a. TSCA
 - b. REACH
 - c. ICCON
 - d. NAFDAC
 - e. NDLEA
 - f. CSN
4. i. Name two drug control agencies in Nigeria
 - i. What are their functions?
5. Define these terms: (i) Pollution (ii) Pollutant (iii) Air pollution
6. What are the three classes of pollutants? (ii) Write three examples of each.
7. What are the social effects of environmental pollution?

CHAPTER 6

FATS AND OILS

6.1 Fats and Oils

Fatty acids are the organic acids. They belong to a homologous series with the general molecular formula $C_nH_{2n+1}COOH$. They have the carboxyl group - COOH as their functional group. Organic acids are also called the carboxylic acids. The higher members of this homologous series are sometimes called the fatty acids because they occur in fats and oil obtained from plants and animals. Examples of such fatty acids are:-

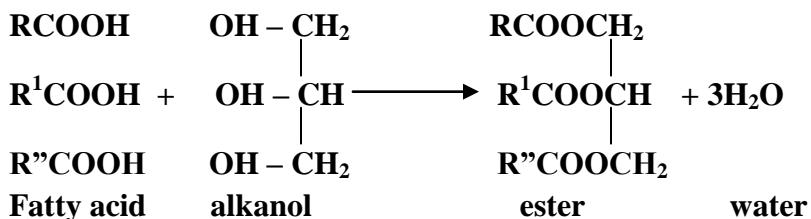
Molecular Formula	IUPAC Name	Old name
$C_{15}H_{31}COOH$	Hexadecanoic acid	Palmitic acid
$C_{17}H_{35}COOH$	Octadecanoic acid	Stearic acid
$C_{17}H_{33}COOH$	Octadec-9-enoic acid	Oleic acid

Octadecenoic acid, $C_{17}H_{33}COOH$ is an unsaturated fatty acid.

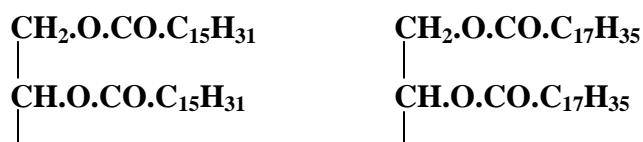
Alkanoic acids are named by replacing the 'e' in the parent alkane with 'OIC' followed by the - word acid.

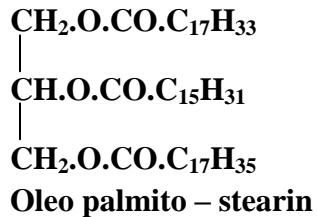
6.2 Sources of Fats and oil

Fats are obtained from plants. Both fats and oils belong to a group of compounds known as lipids. They are esters of the trihydric alkanol, propan – 1, 2, 3 – triol. Each molecules of propane -1, 2, 3-triols reacts with three molecules of fatty acids as follows:



R, R' and R'' may be the same or different. eg





The fatty acids found in lipids are long chain alkanoic acids where R contains about 10 to 18 carbon atoms. They have the general molecule formula; RCOOH .

6.3 Physical Properties of Fats and Oils

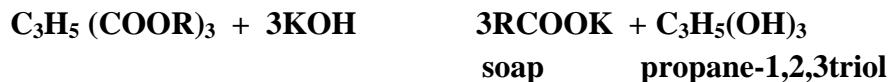
- i. Fats and oil are colourless (yellow, reddish or brown due to impurities) solid or liquids
 - ii. They are lighter than water and immiscible with it
 - iii. They are soluble in organic solvents such as benzene, petroleum ether etc.
 - iv. They are nonvolatile, but decompose on strong heating producing irritating odour.
 - v. They are very weak acids

6.4 Chemical Properties of Fats and Oils

I. On long storage, fats and oils came in contact with air and moisture and undergo slow decomposition. They develop unpleasant smell. The process is known as rancidification. It is caused by enzymatic hydrolysis of fats and oil by certain micro-organisms. Hydrolysis of fats and oil by alkalis is known as **saponification**.

ii. **Saponification:** saponification is the reaction between esters (fats and oils) and an alkalis to form soap and propane -1,2,3-triol.

fat or oil + sodium hydroxide → soap + propane -1,2,3-triol.



iii. **Hydrogenation:** See manufacture of margarine on chapter 7 of this book.

6.5 Test for Fats and Oils

i. **Paper test:** A drop of oil or melted fat on a piece of filter paper forms a translucent grease spot.

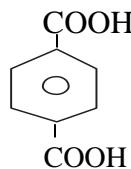
ii. **Sudan (iii) test:** Shake a few cm³ of coconut oil with water. Add about four drops of sudan (iii) stain. The oil is stained red. This is a specific test for fats and oils.

6.6 Uses of Fats and Oils

- i. Many fats and oils are used as food material to provide energy.
- ii. They are used in the manufacture of soaps, hair oils etc.
- iii. Castor oil is used as purgative and cod liver oil as a source of vitamin A and D.
- iv. Linseed oil are extensively used in the manufacture of paints, varnishes, etc.

Exercises

1. Name two fatty acids we studied in this chapter.
2. List two sources of fats and oils.
3. Write two physical properties of fats and oils.
4. Write the IUPAC nomenclature of these alkanoic acids.



5. Why is it that cork and rubber materials are not used during laboratory preparation of ethanoic acid?
6. What is an ester? Write the formulae and IUPAC nomenclature of three esters.
7. What is the difference between fats and oils? Write a named structural formulae of each.

CHAPTER 7

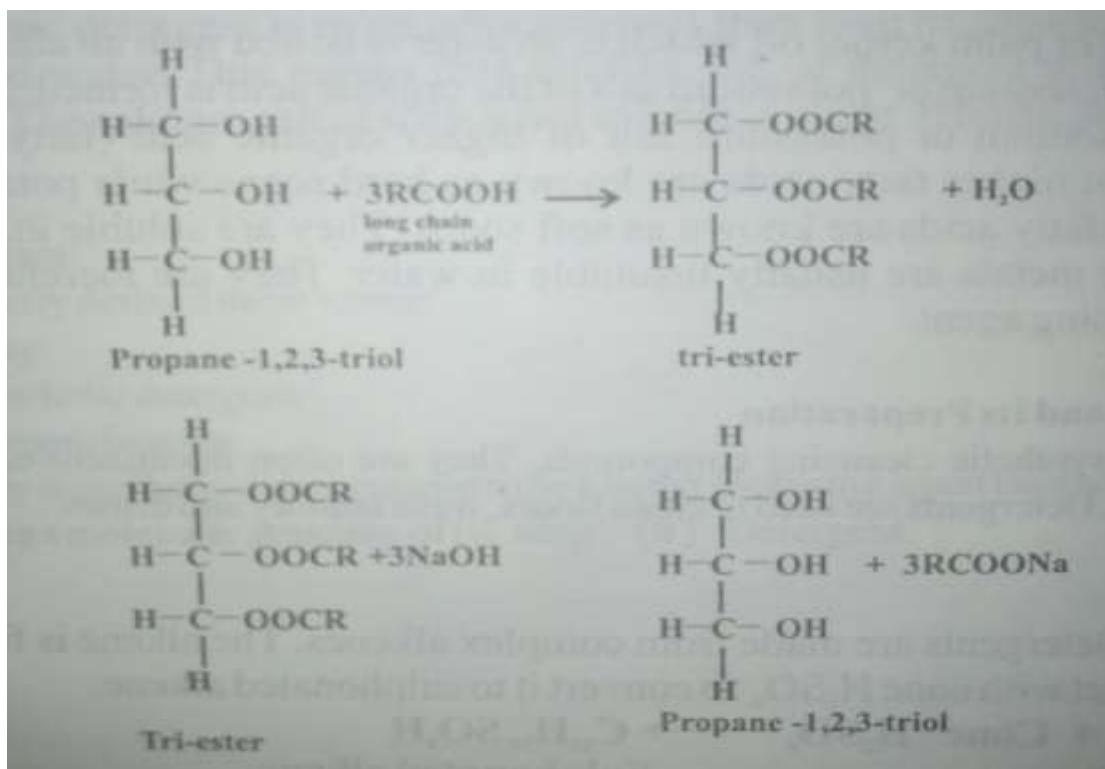
SOAP AND DETERGENTS

7.1 Soap and Its Preparation

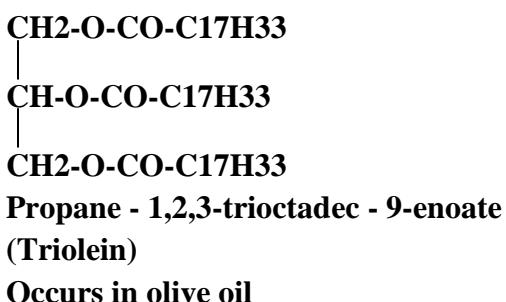
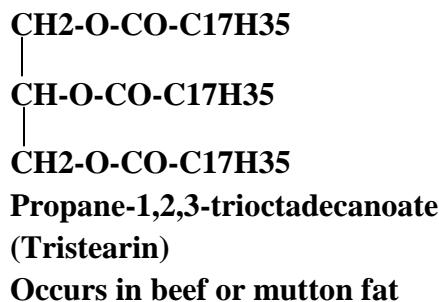
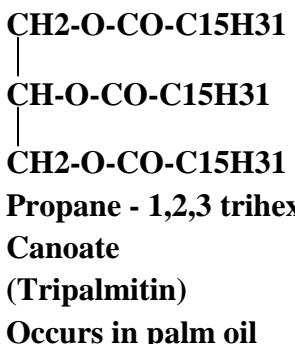
Soap is a chemical compound formed from the reaction between fats or oil and alkalis. It is manufactured by introducing a measured quantity of sodium hydroxide solution into a measured quantity of melted fat or oil. The mixture is boiled and vigorously stirred. The semi-solid boiling mixture is then treated with brine (sodium chloride solution). This causes the contents to separate into two layers the upper layer contains an impure soap. This is then washed to remove the excess alkali and salt.

Saponification: Saponification is the alkaline hydrolysis of esters (fats or oils) to form soap and propane -1, 2, 3-triol.

Fat or oil + sodium hydroxide Soap + propane, 1,2,3 - triol. The formation of propane -1,2,3-triol shows that the esters (fats or oils) used are tri-esters of propane - 1,2,3 - triol with long chain organic acids.



Saponification can also be defined as the hydrolysis of tri-esters of propane - 1,2,3-triol; to form soap. Fats and oils used during soap production are the tri-esters of propane- 1,2,3-triol. Some examples of tri-ester are;

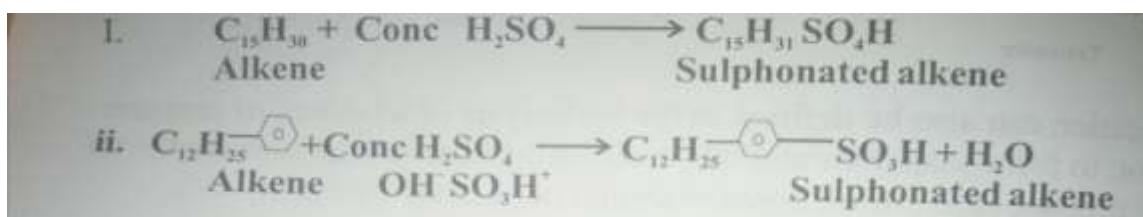


When palm oil or palm kernel oil which is an ester is boiled with an alkali like NaOH or KOH, sodium or potassium salt of the organic acid is formed. Soap is therefore the sodium or potassium salt of higher organic acid (fatty acid). Sodium salts of higher fatty acids are known as hard soaps while potassium salts of higher fatty acids are known as soft soaps. They are soluble in water. Soaps of other metals are usually insoluble in water. They are therefore not used as a cleansing agent.

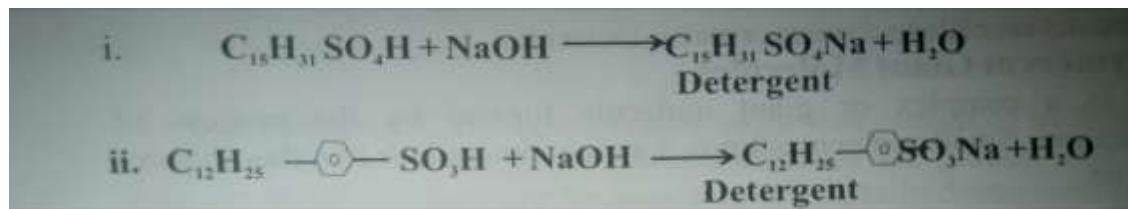
7.2 Detergent and Its Preparation

Detergents are synthetic cleansing compounds. They are often manufactured from petrochemicals. Detergents are used to clean floors, wash laundry and dishes.

Detergents: Detergents are made from complex alkenes. The alkene is first of all made to react with conc H₂SO₄ to convert it to sulphonated alkene.



The sulphonated alkene is then hydrolysed with NaOH or KOH to form the sodium or potassium salt of sulphonated alkene called detergent.



Cleaning action of soap and detergent

Soap and detergents are polar molecules with water soluble (hydrophilic) and fat soluble (hydrophobic) ends.

C12H25 -	SO4Na	C15H31 -	COONa
Hydrophobic	Hydrophilic	Hydrophobic	Hydrophilic
end	end	end	end
(Synthetic detergent)		(soap)	

Soap cleanses by lowering the surface tension of water and by loosening the dirt particle on the cloth. The dirt then floats in the foam.

Synthetic detergent is more advantageous than soap because it is not affected by hard water. This means that calcium salt of detergent is very soluble in water. The calcium salt of soap is not soluble in water. It forms scum.

Exercises

1 Briefly defined these terms:

- (i) Soap
- (ii) Synthetic detergent
- (iii) Saponification.

2. Why is synthetic detergent said to be a better cleansing agent than soap?

3 Write a molecular structure of (i) soap (ii) Detergent.

CHAPTER 8

GIANT MOLECULES

Giant molecules are called polymers or macro-molecules.

8.1 Polymers or Giant Molecules

Polymer is a complex or giant molecule formed by the process of polymerization. Small molecules known as monomers link together to form the polymer.

Some examples of polymer are starch, cellulose, protein, polyethene. There are three principal methods of polymerization. These three methods of preparing synthetic polymers are:

- i. Addition polymerization
- ii. Condensation polymerization
- iii. Cross-linkage polymerization

Read polymerization in more details in part II, chapter 15.5 of this book.

Some examples of addition polymers are: polyethene, Examples of condensation polymers are polyester, Nylon-6,6, maltose.

An example of cross-linkage polymer is vulcanised rubber.

Two types of polymer are natural polymers and artificial polymers. Examples of natural polymer are starch, protein, cellulose. They are natural polymers in that they are found in the body of living things.

Some polymers and their monomers are as follows:

Polymer	Monomer
Starch	Glucose
Protein	Amino acid
Cellulose	Glucose
Polychloroethene	Chloroethene
Polythene	ethene

8.2 Sources and Classification of Sugars

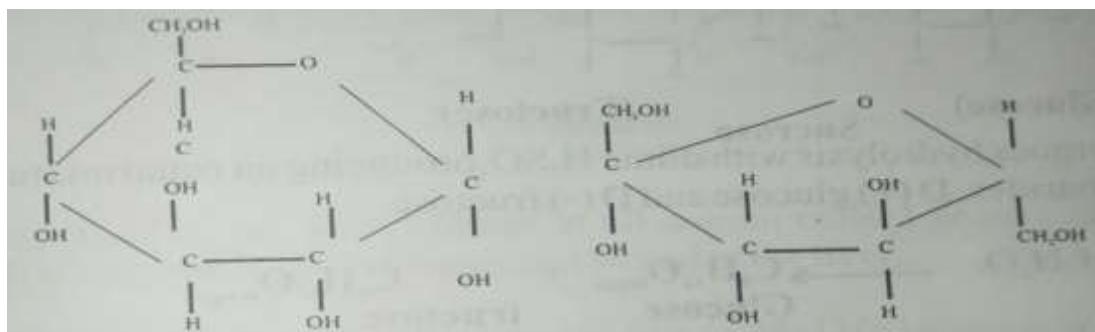
Carbohydrates: are a family of naturally occurring compounds containing carbon, hydrogen and oxygen. They are called carbohydrates meaning hydrates of carbon because they contain

hydrogen and oxygen in the same proportion as in water. They have the general formula $C_x(H_2O)_y$. Carbohydrates are classified into three. These are:

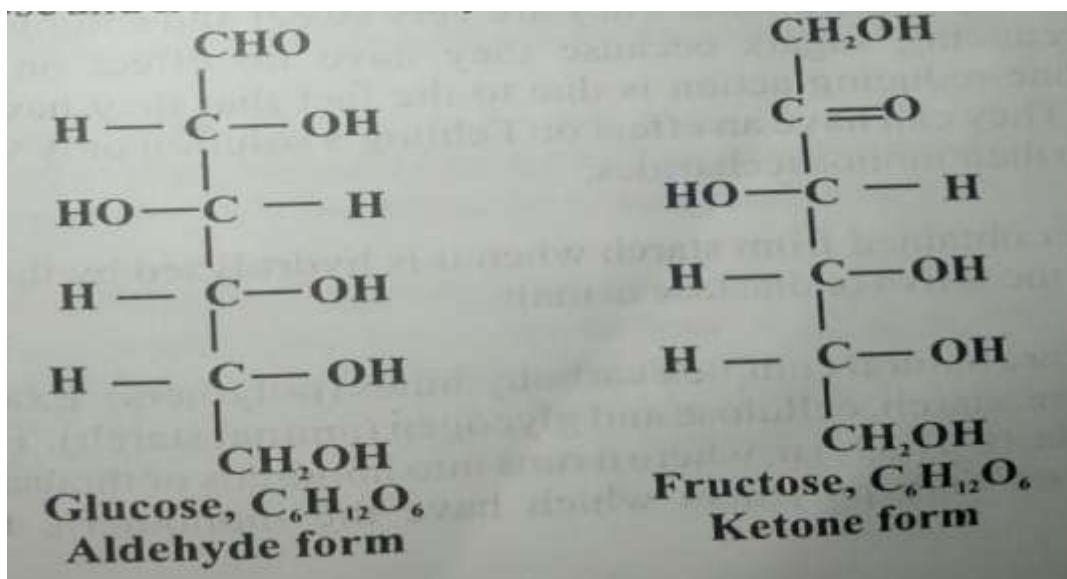
- i. The monosaccharides
- ii. The disaccharides
- iii. The polysaccharides

Monosaccharides, $C_6H_{12}O_6$, are simple sugar that cannot be hydrolysed into simple molecule. Examples of monosaccharides are glucose and fructose. They are called hexose sugar because they are carbohydrates with six carbon atoms. Glucose ($C_6H_{12}O_6$) occurs naturally in ripe grapes and in many sweet fruits. It is sometimes called grape sugar. It also occurs in the blood of animals. Fructose is another simple sugar present in honey. It is a colourless crystalline compound like glucose. It is the sweetest of all sugars. It is more soluble in ethanol and water than glucose.

Both glucose and fructose are isomers. The solution of glucose is dextrorotatory but that of fructose is laevorotatory. Glucose and fructose are reducing sugars. They reduce Fehling's solution to red copper (I) oxide and ammoniacal silver oxide to silver mirror.



Both glucose and fructose have cyclic and open chain structure.

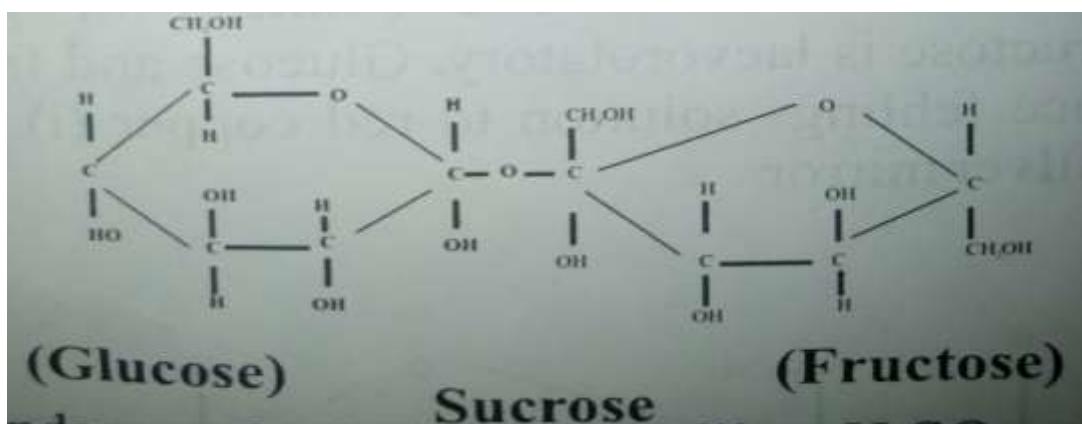
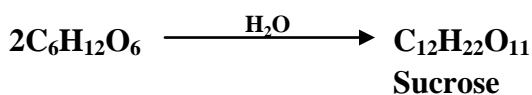


It has been pointed out that both glucose and fructose are hexose sugars, due to the fact that they contain six carbon atoms each. Glucose contains an aldehyde group (CHO) while Fructose contains ketone group ($\text{C} = \text{O}$).

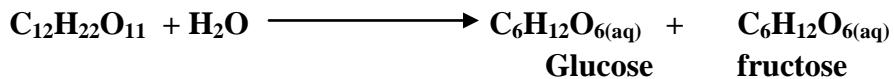
Glucose and Fructose are therefore known as the aldehyhexose and ketohexose sugars respectively.

The disaccharides, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$

Disaccharides are dimeric sugars with twelve (12) carbon atoms in each molecule. Examples of disaccharide are: Sucrose, maltose and lactose. Sucrose is a dimer or a disaccharide with the formula $\text{C}_{12}\text{H}_{22}\text{O}_{11}$. It is the sugar we eat everyday. It is sold in shops and markets. It occurs naturally in sugar cane. It is formed by condensation reaction of two glucose molecules.



Sucrose undergoes hydrolysis with dilute H_2SO_4 producing an equal mixture of its monosaccharides, D (+) glucose and D (-) fructose.



Disaccharides are crystalline solid. They are very sweet and soluble in water. They are non-reducing sugars because they have no effect on Fehlings solution. The none-reducing action is due to the fact that they have no free carbonyl group. They can have an effect on Fehling's solution only when they are hydrolysed to their monosaccharides.

Lactose is a dimer obtained from starch when it is hydrolysed by the enzyme such as ptyalin in the saliva or diastase in malt.

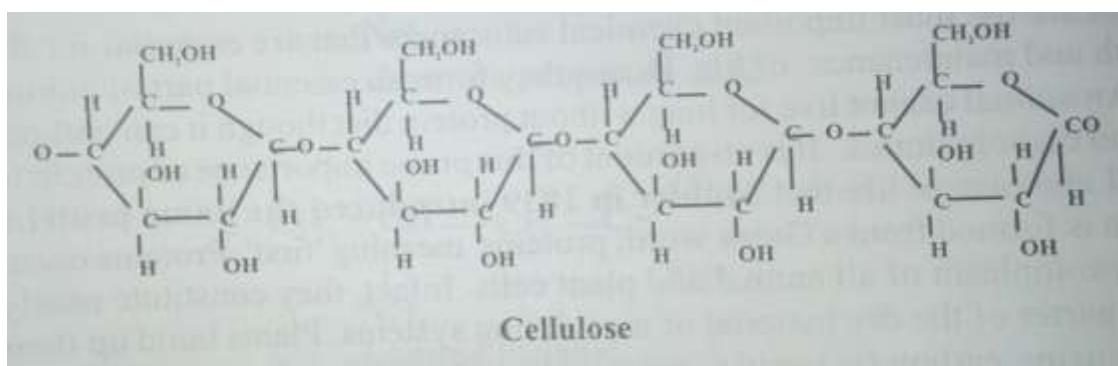
Polysaccharides are natural complex carbohydrates (polymers). Examples of polysaccharides are: starch, cellulose and glycogen (animal starch). They have the general formula, $(C_6H_{10}O_5)^n$, where n runs into hundreds or thousands.

Polysaccharides are high polymers which have the monomeric forms as simple sugar.

Starch is a white powder. It is insoluble in cold water but soluble in hot water to form colloidal solution. Starch gives deep blue colour with iodine solution.

Cellulose is the substance of which the cell walls of plants are made of. Wood has about 50% cellulose. It has the molecular formula of $(C_6H_{10}O_5)_n$, where n is a whole number from 20,000 to 5,000,000.

‘n’ is greater in cellulose than in starch. Cellulose is a whitish solid which is insoluble in water. Complete hydrolysis of cellulose yields only glucose units as the monomeric form. Incomplete or partial hydrolysis of cellulose yields the disaccharides.



Glycogen, $(C_6H_{10}O_5)_n$ is present in all animal cells. Excess carbohydrate or starch is converted to glycogen and stored in the liver.

Hydrolysis of glycogen with dilute H₂SO₄ yields D (+) glucose.

When the glucose in the body is used up, glycogen is broken down or is reconverted to glucose to be oxidized for the release of energy.

8.3 Test for Starch

Add a few drop of iodine solution to a boiled starch solution. A dark blue colour will be produced which disappears on heating and reappears on cooling.

8.4 Test for Simple Sugar (glucose):

Add a few drops of fehling's solution to 5cm³ of glucose solution in a test-tube. A brick-red precipitate is obtained on boiling.

8.5 Uses of Starch

- i. It is used mainly as a food.
- ii. It is used in the production of ethanol and glucose
- iii. It is used as an adhesive to fasten paper.
- iv. It is used as a stiffening agent in the laundry.

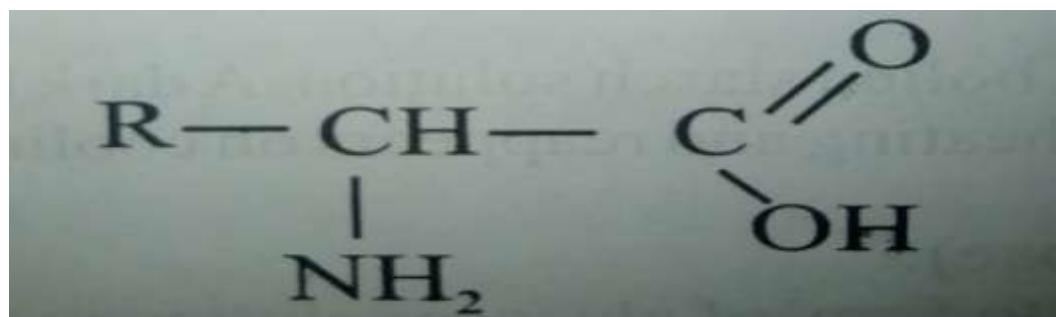
8.6 Uses of Sugar

- i. It is used in the manufacture of jam and sweets.
- ii. Glucose is given to patients and sportsmen as an immediate source of energy.

8.7 Protein and Its Sources

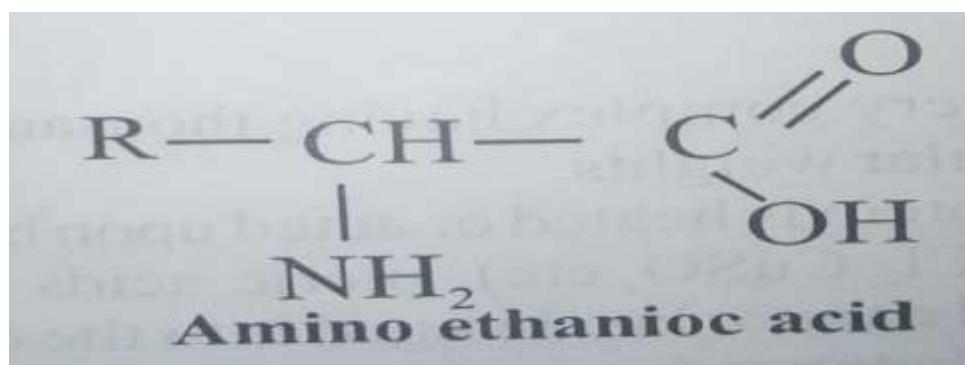
Protein are the most important chemical substances that are essential for the growth and maintenance of life. Hence they form an essential part of animal diet. An animal cannot live for long without protein diet though it can live without fats and carbohydrates. It is on account of this prime importance of protein to animal and human life that Mulder in 1839 introduced the name protein. Protein is formed from a Greek word; proteios' meaning 'first'. Proteins occur in the protoplasm of all animal and plant cells. Infact, they constitute nearly three quarter of the dry material of most living systems. Plants build up their protein using carbon (iv) oxide, water, trioxonitrate (v) and ammonium salts with the help of energy from the sun. Animals and human beings take plant and animal protein in their food which are hydrolysed to amino-acids in the system by enzymes. Protein is resynthesised in the animal and human system from these amino-acids.

8.8 Structure and Properties of Protein



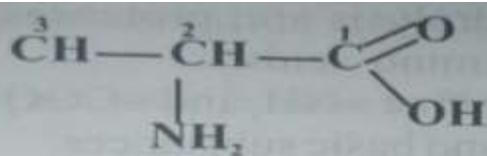
Amino Ethanioc acid is monomeric form of protein. Protein is a polymer.

This formular is derived from carboxylic acid in which one hydrogen atom has been replaced by amino group (-NH₂)

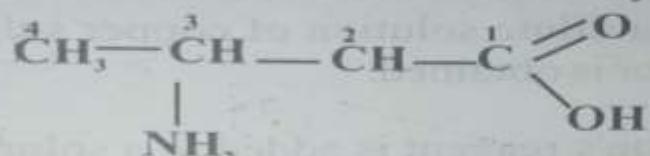


IUPAC Nomenclature of Amino Acids

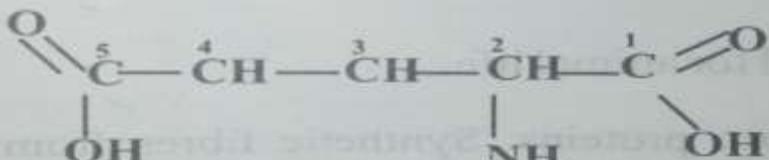
Amino acids are named as derivatives of carboxylic acids.



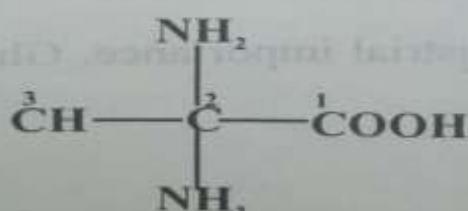
2-amino propanoic acid



3-amino butanoic acid



2-amino penta-1,5-dioic acid



2, 2-diamine propanoic acid

8.9 Properties

- Protein molecules are very complex having thousands of amino-acids and hence very high molecular weights.
- Denaturation; When protein is heated or acted upon by reagents like alcohol, heavy metal salts (HgCl_2 , CuSO_4 etc), picric acids or alkaline-earth salts, proteins coagulates and cannot be changed into the original substance. The solubility of the new substance (and other physical properties) are different from the original ones. This phenomenon is referred to as denaturation.
- Hydrolysis; Proteins are hydrolysed by dilute acids and alkalis or by enzymes. The products of hydrolysis are; proteoses, peptones, poly-peptides, simple peptides and amino-acids.

- iv. Amphoteric Nature: Proteins have free $-NH_2$ and $-COOH$ groups. Hence they react with both the acidic and basic substances.

8.10 Test For protein

- i. Biuret Test: On addition of a dilute solution of copper sulphate to alkaline solution of protein, a violet color is obtained.
- ii. Million's Test: When Million's reagent is added to a solution of protein , a white precipitate which turns red on heating may be formed.

Uses of protein

- i. Protein are needed as food for animal life.
- ii. Wool and natural silk are proteins. Synthetic fibres from caseine and soyabean protein have now been made or manufacture.
- iii. Many of the anti-virus vaccines are protein in denatured form.
- iv. Protein like glue are substance of industrial importance. Glue is used as adhesive in timber industry.

Exercises

- 1a. What is a polymer? (b). Write two examples of polymer.
- 2a. Define polymerization. (b). What are the three methods of polymerization?
- 3a. Why are protein and starch said to be a natural polymer? (b). What are the monomers of protein and starch?
4. Differentiate between these classes of carbohydrate and give two examples of each. (a). Monosaccharide (b). Disaccharide (c) . Polysaccharide
5. Is protein important to human health? Why?
6. List the products of hydrolysis of protein.
7. How can you identify the presence of protein in a substance?

CHAPTER 9

LIKELY EXAMINATION QUESTIONS:

9.1 Qualitative Analysis

Credit would be given for strict adherence to instruction, for observations precisely recorded and for accurate inferences. All tests, observations and inferences must be clearly entered in your answer book, in ink, at the time they are made.

1. C is a simple salt. Carry out the following exercise on C. Record your observations and state the conclusion you draw from the result of each test.
 - a. Put all of C into a beaker or boiling tube and add about 10cm³ of distilled water. Stir thoroughly to make a solution.
 - b. (i) Divide the solution into three portions. To the first portion, add NaOH_(aq) in excess.

- (ii) To the second portion, add aqueous ammonia in drops until it is in excess.
- (iii) To the third portion, add dilute Trioxonitrate(v) acid, $\text{HNO}_{3(\text{aq})}$ followed by silver trioxonitrate (v) $\text{AgNO}_{3(\text{aq})}$ solution.

Answer

TEST	OBSERVATION	INFERENCE
Specimen C	Non-crystalline white powder that is fused into round solids	Calcium, Zinc, lead and aluminium salt suspected.
Specimen $\text{C} + \text{H}_2\text{O}$	Soluble clear solution is formed	Specimen C is soluble in H_2O .
Specimen C (1st portion)	The white precipitate (ppt) formed is insoluble	Ca^{2+} present
Solution + NaOH(aq) in Drop and in excess	in excess NaOH(aq)	

TEST	OBSERVATION	INFERENCE
2nd portion of Sp.C solution + NH_4OH in drops and in excess	No ppt is formed	Ca^{2+} confirmed
3rd portion of Sp.C solution + $\text{Dil. HNO}_{3(\text{aq})}$ + $\text{AgNO}_{3(\text{aq})}$	White precipitate insoluble in dil HNO_3 but soluble in excess. NH_4OH .	Cl^- present and confirmed
	$\text{Cl}^{-(\text{aq})} + \text{Ag}^+ \rightarrow \text{AgCl(s)}$ White ppt	

Therefore specimen C contains calcium, Ca^{2+} and chloride, Cl^- ions.

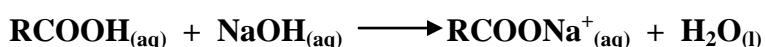
2. Divide specimen D into three portions
- To the first portion add distilled water and test with litmus paper.
 - To the second portion add sodium hydroxide solution drop by drop until in excess.
 - To the third portion add a saturated solution of sodium hydrogen trioxocarbonate (iv), NaHCO_3 .

Answer

TEST	OBSERVATION	INFERENCE
Specimen D	White salt	
Specimen D + H_2O + wet blue or red litmus paper	D is insoluble in H_2O . D floats on water. The liquid is neutral to litmus	Organic compound suspected.

TEST	OBSERVATION	INFERENCE
Specimen D + NaOH (aq)	D dissolves to form a clear colourless solution	Phenol, alkanotic acid,
Specimen D + saturated NaHCO_3 (aq)	D dissolves . There is effervescence. A colourless, odourless gas that turns lime water milky is given out.	Gas is CO_2 . -COOH is present. The -COOH is from an alkanoic acid or Benzoic acid.

Equation for the reactions:





Where R is alkyl or Benzyl group.

Therefore D contains a carboxyl group -COOH from either a higher alkanoic acid or Benzoic acid.

3. Specimen E is a liquid compound. Divide it into three portions.

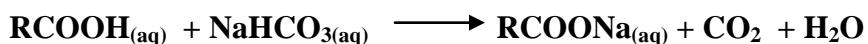
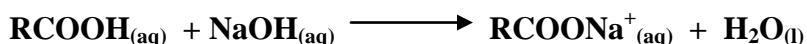
- (i) To the first portion add distilled water and test with litmus paper.
- (ii) To the second portion add NaOH solution.
- (iii) To the third portion add a saturated solution of $\text{NaHCO}_3_{(\text{aq})}$

Answer

TEST	OBSERVATION	INFERENCE
Specimen E is a liquid	The liquid is colourless with a sharp pungent smell.	An organic compd with a-COOH
E + distilled H_2O + litmus paper	E is soluble in H_2O and turns blue litmus paper red .	An alkanoic acid suspected.

TEST	OBSERVATION	INFERENCE
E + saturated solution of NaHCO_3	There is effervescence. E dissolves. A colourless and odourless gas is evolved. The gas turns lime water milky.	CO_2 gas from CO_3^- or HCO_3^- .

Equation for the reaction:



Where R is an alkyl group.

Question

Credit will be given for strict adherence to instructions, for observations precisely recorded and for accurate inference. All tests, observations and inferences must be clearly entered in your answer book, in ink, at the time they are made.

X is a mixture of salts. Carry out the following exercise on X. Record your observation and identify any gas evolved. State the conclusion you draw from the result of each test.

- a (i) Put half of X in a test tube and heat strongly.
 (ii) Test the gas evolved with litmus papers.
- b. Make a solution of about 10cm³ of X and divide it into 3 portions.
 - (i) To the first portion, add sodium hydroxide solution in drops, and then in excess. Heat the product obtained.
 - (ii) To the second portion, add aqueous ammonia in drops and then in excess.
 - (iii) To the third portion add Barium Chloride solution followed by a few drops of dilute HCl or HNO₃

Answer

TEST	OBSERVATION	INFERENCE
Specimen X	The mixture is greenish white and crystalline	
X + heat	Cracking sound is heard. H ₂ O condenses on the upper sides of Test tube. A colourless gas with a choking smell is evolved. The Gas turns red litmus paper blue. Dirty-yellow residue is left.	NH gas is evolved. NH ₄ ⁺ present H ₂ O of crystallization is present
		460

X + strong heat	The residue above decomposes. The gas evolved turns white Blue litmus paper red.	Fe ²⁺ suspected. The gas evolved is acidic
X + Water	X is soluble. Yellowish-green solution is formed	X contains soluble salts.
Solution of X + NaOH _(aq) in drops and in excess	Dirty-green gelatinous ppt. The ppt is insoluble in excess NaOH. On standing, the ppt turns brown.	Fe ²⁺ present
The ppt from above + heat	A gas with a characteristic choking smell is evolved. The gas turns wet red Litmus paper blue.	NH ⁺⁴ confirmed
Solution of X + NH _{3(aq)} in drops and in excess	Dirty-green gelatinous ppt which is insoluble in excess NH _{3(aq)} is formed. The ppt turns brownish on Exposure	Fe ²⁺ confirmed
Solution X + BaCl _{2(aq)}	White ppt is observed	SO ₄ ²⁻ , SO ₃ ²⁻ CO ₃ ²⁻ Suspected
Ppt from above +	The ppt is insoluble in dil. HCl _(aq)	SO ₄ ²⁻ confirmed

9.2 Volumetric Analysis

Question

All your burette readings (initial and final) as well as the size of your pipette must be recorded. No account of experimental procedure is required. All calculations must be done in your answer book.

A is dilute HCl containing 3.65g of HCl per dm³. B is a solution of impure Na₂CO₃ containing 1.500gm of the impure salt per 250cm³.

- a. Put A into the burette and titrate with 25cm³ portions of B using methyl orange as indicator. Record the volume of your pipette. Tabulate your burette readings and calculate the average volume of A used.

- b. From your results and information provided, calculate the:-
- Concentration of A in mol dm^{-3}
 - Concentration of B in mol dm^{-3}
 - Concentration of B in g dm^{-3}
 - The percentage (%) purity of Na_2CO_3 in the impure solid.

The Equation is:-



Answer

Volume of pipette = 25cm³

Burette Readings	TITRATIONS			
	Rough	1st	2nd	3rd
Final reading (cm ³)	25.50	25.50	30.50	8.50
Initial readings (cm ³)	0.00	1.00	6.00	4.00
Volume of acid used	25.50	24.50	24.50	24.50

$$\text{Average vol. of acid used} = \frac{24.50 + 24.50 + 24.50}{3} ^{\circ} 1 = 24.50 \text{ cm}^3$$

i. Concentration of A in mol dm^{-3}

$$\frac{\text{mass conc of A in } \text{g dm}^{-3}}{\text{Molar mass of A in } \text{g mol}^{-1}}$$

But mass conc. of A = 3.650gdm⁻³

$$\therefore \text{Conc of A in } \text{mol dm}^{-3} = 3.650 \text{ gdm}^{-3}$$

$$\begin{aligned} & 365 \text{ g mol}^{-1} \\ & = 0.01, \text{ mol dm}^{-3} \end{aligned}$$

ii. Conc. of B in mol dm^{-3}

$$\frac{\text{From } C_A V_A}{C_B V_B} = \frac{a}{b} \quad \frac{C_A \times 24.50}{C_B \times 25.0 \text{cm}^3} = \frac{2}{1}$$

$$\frac{0.100 \times 24.50}{C_B \times 25.0} = \frac{2}{1}$$

$$C_B = 0.100 \times 2450 \times 1$$

=

$$25.0 \times 2$$

$$= 0.049$$

$$\therefore \text{Conc of B} = 0.05 \text{ moldm}^{-3}$$

(iii) Molar mass of (B) $\text{Na}_2\text{CO}_3 = 106 \text{ g mol}^{-1}$

$$\text{Mass Conc of pure NaCO}_3 = 0.05 \text{ moldm}^{-3} \times 106 \text{ g mol}^{-1}$$

$$\text{Mass Conc of pure NaCO}_3 = 5.30 \text{ g dm}^{-3}$$

To find percentage of Na_2CO_3

$$\text{Mass Conc. of impure Na}_2\text{CO}_3 = \frac{\text{Mass m}}{\text{Volume v}}$$

$$\text{Mass m} = 1.500 \text{ g}$$

$$\text{Volume v} = 250 \text{ cm}^3 = 0.25 \text{ dm}^3$$

$$\text{Mass Conc} = \frac{m}{v} = \frac{1.500 \text{ g}}{0.25 \text{ dm}^3} = 6.00 \text{ g dm}^{-3}$$

$$\therefore \% \text{ purity of Na}_2\text{CO}_3 = \frac{5.30}{6.00} \times \frac{100}{1} = 88.3\%$$

Question

All your burette readings (initial and final), as well as the size of your pipette, must be recorded but no account of experimental procedure is required. All calculations must be done in your answer book.

A is a solution containing 0.104 moldm^{-3} of H_2SO_4 . B is a solution containing 2.90 g in 100 cm^3 of hydrated sodium trioxocarbonate(iv), $\text{Na}_2\text{CO}_3 \cdot \text{XH}_2\text{O}$.

Put A in the burette and titrate 25 cm^3 portion of B using methyl orange as indicator. Tabulate your burette readings and calculate the average volume of acid used.

From your result and information provided, calculate the

- Write a balanced equation of the reaction
- (i) Concentration of anhydrous Na_2CO_3 in solution B in mol dm^{-3}
- (ii) Mass concentration of anhydrous Na_2CO_3
- (iii) The value of X
- (iv) The percentage of water of crystallization of solution B

(H = 1, Na = 23, C = 12, O = 16).

Answer

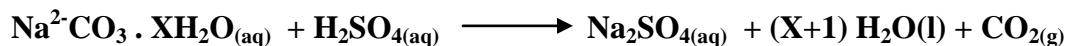
$$\text{Volume of pipette used} = 25.00 \text{ cm}^3$$

Titration Table

Burette Readings (in cm^3)	TITRATIONS			
	Rough	1st	2nd	3rd
Final reading (cm^3)	24.01	29.06	34.06	44.06
Initial readings (cm^3)	0.00	5.00	10.00	20.00
Volume of acid used (cm^3)	24.01	24.06	24.06	24.06

$$\begin{aligned}\text{Average vol. of acid used} &= \frac{24.06 + 24.06 + 24.06}{3} \\ &= 24.06 \text{ cm}^3\end{aligned}$$

Equation for the reaction:



X in $\text{Na}_2\text{CO}_3 \cdot \text{XH}_2\text{O}$ represented the number of water of crystallization.

In $(\text{X}+1)\text{H}_2\text{O}$, X represents the number of water of crystallization while I represents one mole of water which would have been produced by anhydrous sodium trioxocarbonate (iv).

- Conc. of anhydrous Na_2CO_3 in solution B in mol dm^{-3} . From equation for the reaction, the ratio of amount of acid to base is 1:1

$$\frac{C_a V_a}{C_b C_b} = \frac{1}{1} \quad (\text{ratio of amount})$$

Substituting known values in (1) then

$$\frac{0.104 \text{ moldm}^{-3} \times 24.06 \text{ cm}^3}{C_b \times 25 \text{ cm}^3} = \frac{1}{1}$$

$$C_b \times 25 \text{ cm}^3 \times 1 = 0.104 \text{ moldm}^{-3} \times 24.06 \text{ cm}^3 \times 1$$

$$C_b = \frac{0.104 \text{ moldm}^{-3} \times 24.06 \text{ cm}^3 \times 1}{25 \text{ cm} \times 1}$$

$$C_b = 0.10 \text{ moldm}^{-3}$$

(ii) Mass Concentration of anhydrous Na_2CO_3

$$\text{Conc in moldm}^{-3} = \frac{\text{Concentration in gdm}^{-3}}{\text{Molar Mass}}$$

$$\text{Molar Mass of } \text{Na}_2\text{CO}_3 = (23 \times 2) + 12 + (16 \times 3)$$

$$46 + 12 + 48 = 106 \text{ g mol}^{-1}$$

$$\therefore 0.100 \text{ moldm}^{-3} = \frac{\text{Conc of anhydrous } \text{Na}_2\text{CO}_3 \text{ in dm}^{-3}}{106 \text{ g mol}^{-1}}$$

$$\text{Conc of anhydrous } \text{Na}_2\text{CO}_3 = 0.10 \text{ moldm}^{-3} \times 106 \text{ g mol}^{-1}$$

$$= 10.6 \text{ g dm}^{-3}$$

(ii) Number of water of crystallization . X using the formula:

$$\frac{\text{Mass Conc. of hydrated } \text{Na}_2\text{CO}_3}{\text{Mass Conc. of anhydrous } \text{Na}_2\text{CO}_3} = \frac{\text{Molar Mass of } \text{Na}_2\text{CO}_3 \cdot \text{XH}_2\text{O}}{\text{Molar Mass of } \text{Na}_2\text{CO}_3}$$

But the conc. of hydrated Na_2CO_3 is 2.90g in 100 cm^3

To find the conc. of $\text{Na}_2\text{CO}_3 \cdot \text{XH}_2\text{O}$ in g dm^{-3}

100 cm^3 contain 2.90g

$$1000 \text{ cm}^3 (1 \text{ dm}^3) \text{ will ,, } 2.90 \times \frac{1000}{100} = 29.0 \text{ g dm}^{-3}$$

$$\frac{29.0 \text{ g dm}^{-3}}{10.6 \text{ g dm}^{-3}} = \frac{\text{Molar Mass of } \text{Na}_2\text{CO}_3 \cdot \text{XH}_2\text{O}}{\text{Molar Mass of } \text{Na}_2\text{CO}_3}$$

(Where X is the number of water of crystallization)

$$\frac{29.0 \text{ gdm}^{-3}}{10.6 \text{ gdm}^{-3}} = \frac{106 + 18x}{106}$$

$$3074 = 1123.6 + 190.8X$$

$$1950.4 = 190.8X$$

$$X = \frac{1950.4}{190.8} = 10$$

- (iv) The percentage of water of crystallization in hydrated Na_2CO_3 is calculated by applying the formula:

$$\frac{\text{Mass of water}}{\text{Mass of hydrated } \text{Na}_2\text{CO}_3} \times \frac{100 \%}{1}$$

Mass of water = Mass of hydrated

$$\begin{aligned} \text{Mass of H}_2\text{O} &= \text{Mass of hydrated } \text{Na}_2\text{CO}_3 - \text{Mass of anhydrous } \text{Na}_2\text{CO}_3 \\ &= 29.0 \text{ gdm}^{-3} - 10.6 \text{ gdm}^{-3} = 18.40 \text{ gdm}^{-3} \end{aligned}$$

$$\begin{aligned} \% \text{ of water of crystallization} &= \frac{18.40}{29.00} \times \frac{100}{1} \\ &= 63.45\% \end{aligned}$$

Question:

Solution A contains 2.2g tetraoxosulphate(vi) acid, H_2SO_4 per 500cm^3

Solution.

Solution B is prepared by dissolving 6.5g of impure Na_2CO_3 per dm^3 solution.

Put A in the burette and titrate with 25cm^3 portion of B using methyl orange as indicator. Record the volume of your pipette. Tabulate your burette readings and calculate the average volume of A used.

From your result calculate

- (a) Concentration of A in mole dm^{-3}

(b) Concentration of B in gdm^{-3}

(c) Percentage purity of B.

Equation for the reaction



Volume of pipette = 25cm^3

Solution in pipette = B($\text{Na}_2\text{CO}_{3(\text{aq})}$)

Solution in Burette = A ($\text{H}_2\text{SO}_{4(\text{aq})}$)

Indicator used = Methyl orange

Colour of indicator in Na_2CO_3 = Yellow

Colour at end-point = orange

Burette Readings (in cm^3)	TITRATIONS			
	Rough	1st	2nd	3rd
Final reading (cm^3)	25.50	25.30	26.30	27.30
Initial readings (cm^3)	0.00	0.00	1.00	2.00
Volume of acid used (cm^3)	25.50	25.30	25.30	25.30

$$\begin{aligned}\text{Average vol. of acid used } \text{cm}^3 &= \frac{25.30 + 25.30 + 25.30}{3} \\ &= 25.30 \text{cm}^3\end{aligned}$$

a. To find the conc. of A in mol dm^{-3} using the formula;

$$\text{No. of mole} = \frac{\text{Mass in grams}}{\text{Molar mass}}$$

$$\text{No. of mole} = \underline{2.2} = 0.024$$

∴ 500cm³ contains 0.0224 mole

$$\therefore 1000\text{cm}^3 \text{ will contain } 0.0224 \times \frac{1000}{500}$$

$$\text{Molar conc. of A} = 0.0448 \text{ moledm}^{-3}$$

b. To find the concentration of B in gdm⁻³

First find the conc. of B in moldm-3 by using the volumetric equation as follows:

$$\frac{C_A V_A}{C_B V_B} = \frac{A}{B}$$

Where C_A = Conc. of acid (0.0448 moldm⁻³)

V_A = Volume of acid (25.30cm³)

A = No. of mole of acid (1)

C_B = Conc. of base in mole dm⁻³

V_B = Volume of base (25cm³)

B = No. of mole of base (1)

$$\frac{0.0448 \times 25.30}{C_B \times 25} = \frac{1}{1}$$

$$C_B = 0.0453 \text{ mole dm}^{-3}$$

$$\therefore \text{Molar Conc. of B} = 0.0453 \text{ moledm}^{-3}$$

To find the concentration of B in gdm⁻³,

We use:

$$\text{Molar Conc. in mole dm}^{-3} = \frac{\text{Mass Conc. in gdm}^{-3}}{\text{Molar Mass in gmol}^{-1}}$$

$$\text{Mass Conc. in gdm}^{-3} = \text{Molar Conc. in moldm}^{-3} \times \frac{\text{Molar Mass in g mol}^{-1}}{1000}$$

$$= 0.0453 \times 106$$

$$\therefore \text{Mass Conc. of B} = 4.80 \text{ gdm}^{-3}$$

c. To find the percentage purity of X, 6.50g of impure B contain 4.80g pure Na₂CO₃

$$\therefore \frac{4.80}{6.50} \times \frac{100}{1} = 73.85\%$$

$$\% \text{ purity of B} = 73.85\%$$

Question

To calculate the amount of ions and the actual number of ions in solutions.

One mole of any substance contains Avagadro's Number of particles. The particles may either be atoms, molecules or ions. The number is constant and has a value of 6.02×10^{23}

Example:

Solution A contains 4.9g H₂SO₄ dm⁻³

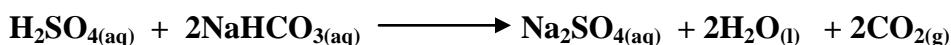
Solution B is a solution of NaHCO₃.

Put A into the burette and titrate as usual.

From your results and the information provided, calculate:

- (i) The concentration of B in mol dm⁻³
- (ii) The amount of Na⁺ ions and HCO₃⁻ ions that reacted.
- (iii) The number of Na⁺ ions and HCO₃⁻ ions that reacted.
- (iv) The mass conc. of B in g dm⁻³.

The equation for the reaction of A and B is:



Volume of pipette = 25.0 cm³

Titration Table

Burette Readings (in cm ³)	TITRATIONS			
	Rough	1st	2nd	3rd
Final reading (cm ³)	23.00	45.50	22.50	42.50
Initial readings (cm ³)	0.00	23.00	0.00	20.00
		469		

Volume of acid used (cm³) 23.00 22.50 22.50 22.50

$$\text{Average vol. of acid used cm}^3 = \frac{22.50 + 25.50 + 22.50}{3} \\ = 22.50 \text{ cm}^3$$

(i) Conc. of B in moldm⁻³

First find the Conc. of acid in moldm⁻³ by using the formular:

$$\text{Conc. of acid in moldm}^{-3} = \frac{\text{Mass Conc}}{\text{Molar mass}} = \frac{4.9 \text{ gdm}^{-3}}{98 \text{ gmol}^{-1}} \\ = 0.05 \text{ moldm}^{-3}$$

$$\text{Using volumetric formular } \frac{C_A V_A}{C_B V_B} = \frac{\text{Amount of A}}{\text{Amount of B}}$$

$$= \frac{0.05 \text{ moldm}^{-3} \times 22.50 \text{ cm}^3}{C_B \times 25.0 \text{ cm}^3} = \frac{1}{2}$$

$$C_B = \frac{0.05 \text{ moldm}^{-3} \times 22.50 \text{ cm}^3 \times 2}{25.0 \text{ cm}^3 \times 1} = 0.09 \text{ moldm}^{-3}$$

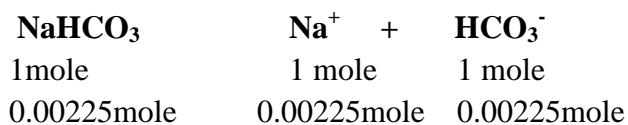
The Conc. of B = 0.09 moldm⁻³

(ii) Amount of NaHCO₃ = Conc. of NaHCO₃ in moldm⁻³ X Vol. of NaHCO₃ in dm³

$$\text{But Volume of NaHCO}_3 \text{ in dm}^3 = \frac{25}{1000} = 0.025 \text{ dm}^3$$

$$\begin{aligned} \text{Amount of NaHCO}_3 &= 0.09 \text{ moldm}^{-3} \times 0.025 \text{ dm}^3 \\ &= 0.00225 \text{ mol HCO}_3^- \\ &= 2.25 \times 10^{-3} \text{ mole HCO}_3^- \end{aligned}$$

Ionization of NaHCO₃



1. mole of NaHCO₃ gives 1 Mole of Na⁺ and 1 mole of HCO₃⁻

∴ 0.00225 mole will yield 0.00225 mole of Na^+ ions and 0.00225 mole of HCO_3^-

iii. 1 mole of Na^+ contains 6.02×10^{23} ions (Avagadro Number)

$$\begin{aligned}\therefore 0.00225 \text{ mole will contain } & 6.02 \times 10^{23} \times \underline{0.00225} \\ & 1 \\ & = 1.35 \times 10^{21} \text{ Na}^+ \text{ ions}\end{aligned}$$

Similarly;

1 mole of HCO_3^- contain 6.02×10^{23} ions

$$\begin{aligned}0.00225 \text{ mole will contain } & 6.02 \times 10^{23} \times \underline{0.00225} \\ & 1\end{aligned}$$

$$= 1.35 \times 10^{21} \text{ HCO}_3^- \text{ ions}$$

ii. Mass Conc of B in gdm^{-3} = molar conc of B in mol dm^{-3} x Molar Mass of B

$$\text{Molar Mass of B} = 23 + 1 + 12 + 48 = 84 \text{ g mol}^{-1}$$

$$\therefore \text{Mass Conc.} = 0.09 \text{ mol dm}^{-3} \times 84 \text{ g mol}^{-1} = 7.56 \text{ g dm}^{-3}$$

CHAPTER 10

PAST QUESTION : ALTERNATIVE TO PRACTICAL: W.A.S.S.C.E. 2002: NOV.

1. A student was required to standardize a solution of HCl using $0.125 \text{ mol dm}^{-3}$ Na_2CO_3 solution.

The student titrated the acid against 20.0 cm^3 portions of the Na_2CO_3 solution using phenolphthalein indicator; and tabulated the results as indicated below:

	Rough	1st	2nd	3rd
Final reading (cm ³)	24.00	22.75	22.5	22.40
Initial readings (cm ³)	0.00	0.00	0.00	20.00
Volume of acid used (cm ³)	24.00	22.75	22.5	22.40

Average titre value = $\frac{22.75 + 22.5 + 22.40}{3}$

- (a) (i) Name the piece of apparatus used for measuring the portions of Na₂CO₃ solution.
(ii) Identify three errors in the result presented.

[4 marks]

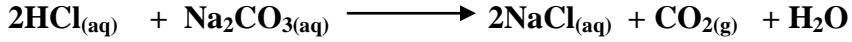
- (b) (i) State the colour of phenolphthalein in the Na₂CO₃ solution
(ii) Is phenolphthalein a suitable indicator for this titration?

Give the reason for your answer. [3 marks]

- (ii) After the burette had been washed and rinsed with water, what final cleaning step should be taken before filling it with the acid?
(c) (i) What effect would the acid have on litmus paper?

[2 marks]

- (d) Using an assumed titre value of 25.45cm³, determine the concentration of the HCl solution in mol dm⁻³. The equation for the reaction is



[4 marks]

- (e) Calculate the mass of anhydrous Na₂CO₃ required to prepare 250cm³ of 0.125mol dm⁻³ Na₂CO₃ solution.

[C = 12.0; O = 16.0; Na = 23.0] [5 marks]

2. (a) (i) State what is observed when ammonia is tested with concentrated HCl.
(ii) Describe briefly one chemical test to distinguish dilute HCl from dilute HNO₃.
(iii) Give one chemical test for the -COOH functional group. [6 marks]

- (b) In a laboratory procedure for preparing ZnSO_4 , a solid X was added with stirring to warm, dilute H_2SO_4 until no more of it dissolved. Excess of X was removed from the reaction mixture. The solution of ZnSO_4 obtained was concentrated to a small volume and then set aside.
- Suggest three substances which X could be.
 - Name the method used for removing excess of X.
 - What method was used to concentrate the $\text{ZnSO}_{4(\text{aq})}$?
 - State what later appeared in the concentrated $\text{ZnSO}_4(\text{aq})$ that was set aside. [6 marks]

- (c) Copy and complete the table below as appropriate.

Test	Observation
(i) $\text{ZnSO}_{4(\text{aq})} + \text{BaCl}_{2(\text{aq})}$ and then + dil. HCl in excess	
(ii) $\text{ZnSO}_{4(\text{aq})} + \text{NaOH}_{(\text{aq})}$ in Drops and then in excess	
(iii) $\text{ZnSO}_{4(\text{aq})} + \text{NH}_{3(\text{aq})}$ in Drop and then in excess	

[6 marks]

3. (a) A gas Y was passed over heated copper (II) oxide in order to convert it to the metal.
- Identify Y.
 - State the colour change during the experiment.
 - If moderately concentrated HNO_3 were added to the metal, what would be observed? [5 marks]
- (b) Give the reason for each of the following laboratory practices:
- Hydrogen peroxide is stored in dark (amber) reagent bottles;
 - In the preparation of dilute acid solutions, water is not added to the concentrated acid;
 - Hydrochloric acid is prepared from hydrogen chloride by bubbling the gas into water through an inverted funnel. [4 marks]

- (c) State what is observed in each case, on treating lime water with:
- (i) a limited supply of CO₂ and then excess CO₂;
 - (ii) aqueous NaOH solution in drops and then in excess;
 - (iii) about 10cm³ of standard soap solution. [5 marks]

Answer: Alternative to Practical: W.A.S.S.E: 2002

1. A solution of HCl was titrated against 20.0cm³ portions of 0.125 moldm⁻³ Na₂CO_{3(aq)}
- (a) (i) The piece of apparatus used for measuring the portions of Na₂CO₃ solution is pipette.
- (ii) The errors in the result as presented are as follows:

- Unit of measurement was not stated/ No unit.
- One of the titres averaged was not concordant / was divergent
- The burette readings were inconsistent

- (b) (i) Phenolphthalein is pink in $\text{Na}_2\text{CO}_3(\text{aq})$
- (ii) Phenolphthalein is not a suitable indicator for this titration.

Reason: The titration is that of a strong acid and a weak base.

- (c) (i) The acid would turn blue litmus red, no effect on red litmus
- (ii) After washing and rinsing the burette with water, before filling it with the acid, it should be rinsed with the acid.
- (d) To determine the concentration of the HCl using an assumed titre value of 25.45cm^3 .

$$\frac{\text{C}_A \text{V}_A}{\text{C}_B \text{V}_B} = \frac{2}{1}$$

$$\frac{\text{C}_A \times 25.45}{0.125 \times 20} = \frac{2}{1}$$

$$\text{C}_A = \frac{2 \times 0.125 \times 20}{25.45} = 0.196 \text{ mol dm}^{-3}$$

Alternatively, from first principles,

Amount of Na_2CO_3 in 20.0cm^3 of $0.125 \text{ mol dm}^{-3}$ solution

$$= \frac{20 \times 0.125 \text{ mole}}{100}$$

Since mole ratio of HCl to Na_2CO_3 is 2:1,

$$\text{Amount of HCl in } 25.45\text{cm}^3 \text{ solution} = \frac{20 \times 0.125 \times 2}{1000}$$

Hence amount of HCl in 1dm^3 i.e concentration of the

$$\text{HCl solution} = \frac{20 \times 0.125 \times 2}{1000} \times \frac{1000}{25.45}$$

$$\frac{20 \times 0.125 \times 2}{25.45}$$

0.196 mole per dm³ (to 3 sig. Figs.)

- (e) To calculate the mass of anhydrous Na₂CO₃ required to prepare 250cm³ of 0.125 moldm⁻³ Na₂O₃ solution:

$$\text{Molar mass of Na}_2\text{O}_3 = (23 \times 2) + 12 + (15 \times 3)$$

Or 106 gmol⁻¹

∴ (106 × 0.125)gdm⁻³ gives 0.125moldm⁻³ solution

Hence mass required to prepare 250cm³ of

$$0.125\text{moldm}^{-3} \text{ solution} = \frac{106 \times 0.125}{4} \text{ OR } \frac{106 \times 0.125 \times 250}{1000}$$

Correct evaluation 3 sig. Fig. gives 3.31g

(No penalty for not stating units as they are in question paper, but wrong unit does not score.

[4 marks] Total = 18 marks

2. (a)(i) Ammonia gives dense white fumes with concentrated HCl

(ii) To distinguish dilute HCl from dilute HNO₃, add AgNO_{3(aq)} or (CH₃COO)₂Pb_(aq) or Pb(NO₃)_{2(aq)}/solution to separate portions. White precipitate indicates HCl no precipitate indicates HNO₃.

OR

Add FeSO_{4(aq)} to separate portions in a test tube, followed by concentrated H₂SO₄ Brown ring at liquid junction indicates HNO₃, no brown ring indicates HCl.

(iii) Test for the - COOH functional group.

- Add saturated NaHCO₃/Na₂CO₃ solution

Effervescence/gas evolved

OR Add alcohol and warm with a few drops of conc. H₂SO₄ Sweet smell of ester

[6 marks]

- (b) In a procedure for preparing ZnSO₄, X was added with stirring to warm, dilute H₂SO₄ until no more of X dissolved.

X was added with stirring to warm, dilute H₂SO₄ until no more of X dissolved.

(i) X could be Zn; ZnO; ZnCO₃

(ii) Excess of X is removed by filtration

(Accept sedimentation and decantation).

(iii) The $\text{ZnSO}_4\text{(aq)}$ (filtrate) was concentrated by evaporation. (Accept distillation/boiling/heating).

(iv) ZnSO_4 crystals later appeared when the concentrated $\text{ZnSO}_4\text{(aq)}$ was set aside.

(Do not accept white solid/white ppt. For ZnSO_4^4 crystals.

(c)	TEST	OBSERVATION
	(i) $\text{ZnSO}_4\text{(aq)} + \text{BaCl}_2\text{(aq)}$ and then + dil HCl in excess	White precipitate Precipitate insoluble/remained
	(ii) $\text{ZnSO}_4\text{(aq)} + \text{NaOH(aq)}$ in drops and then in excess	White precipitate Precipitate dissolved /Colourless solution formed
	(iii) $\text{ZnSO}_4\text{(aq)} + \text{NH}_3\text{(aq)}$ in drops and then in excess	White precipitate Precipitate dissolved/colourless solution formed
Total = 18 marks.		

3. (a) A gas Y was passed over heated CuO in order to convert it to the metal.

(i) Y is H_2 or CO or name Hydrogen or Carbon (II) oxide

(ii) During the experiment, the colour of the solid changes from black brown

(b) (i) Hydrogen peroxide is stored in dark (amber) reagent bottles because it decomposes in sunlight/its decomposition is photocatalyzed bottles because it decomposes in sunlight/its decomposition is photocatalyzed.

(ii) Water is not added to a concentrated acid in order to dilute it because the solution process is highly exothermic

There could be a violent reaction which could cause bodily harm.

(Hence, the practice is to add the concentrated acid to water).

(iii) Hydrochloric acid is prepared from hydrogen chloride by bubbling the gas into water through an inverted funnel so as to prevent the acid sucking back

- (c) On treating lime water with
- (i) a limited supply of CO_2 and then excess CO_2 , a milkiness/ turbidity/ white precipitate first appears and it later disappears/dissolves in excess CO_2
 - (ii) aqueous NaOH solution in drops and then in excess, a white precipitate is formed which remained insoluble in excess.
 - (iii) about 10cm^3 of standard soap solution, curdiness/scum is formed.
(Accept lather is not readily produced).

PAST QUESTIONS: ALTERNATIVE A: W.A.S.S.E. 2004: JUNE

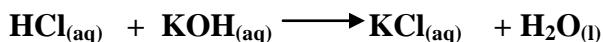
All your burette readings (initial and final) as well as the size of your pipette, must be recorded but no account of experimental procedure is required. All calculation must be done in your answer book.

A is a solution of HCl containing 5.0gdm^{-3} .

B is a solution of impure KOH containing 6.50gdm^{-3}

- (a) Put A into the burette and titrate it against 20.0cm³ or 25.03cm portions of B using methyl orange as indicator. Tabulate your burette readings and calculate the average volume of A used.

The equation for the reaction involved in the titration is:



[10marks]

- (b) From your results and the information provided above, calculate the:

(i) concentration of A in mol dm⁻³;

concentration of B in mol dm⁻³;

(ii) percentage purity of KOH in B.

[H = 1; Cl = 35.5; KOH = 56.0gmol⁻¹]

[10 marks]

2. Credit will be given for strict adherence to the instruction, for observations precisely recorded and for accurate inference. All tests, observations and inferences must be clearly entered in your answer book, in ink, at the time they are made

C is a sample of iron (II) tetraoxosulphate (VI).

D is a sample of Zinc trioxocarbonate (IV).

Carry out the following exercises on C and D. Record your observations and identify any gases evolved. State the conclusion you draw from the result of each test.

- (a) (i) Put all of C in a test tube and add about 5cm³ of distilled water. Stir and test with litmus paper.

Divide the solution into two portions.

(ii) To the first portion, add sodium hydroxide solution in drops and then in excess.

(iii) To the second portion, add few drops of barium chloride solution, followed by dilute hydrochloric acid in excess.

(iii) State one use of each of the following pieces of apparatus in the laboratory.

I. Wash bottle

II. Tripod stand

[5 marks]

(c) State what would be observed when:

- (i) $\text{Pb}(\text{NO}_3)_2$ is heated;
- (ii) concentrated HCl is added to MnO_2 .

[4 marks]

ANSWER: ALTERNATIVE A: W.A.S.S.E. 2004

(a) Averaging - $X \text{ cm}^3$

(b) (i) To calculate the concentration of solution A

$$\text{Molar mass of HCl} = 36.5 \text{ g mol}^{-1}$$

$$\therefore \text{Concentration of A (C}_A\text{)} = \frac{5.0}{36.5} \\ = 0.137 \text{ mol dm}^{-3}$$

(ii) To calculate the concentration of solution B

From the equation, mole ratio of acid to base is 1:1

$$\frac{\text{C}_A \text{V}_A}{\text{C}_B \text{V}_B} = \frac{1}{1}$$

$$\text{C}_B = \frac{\text{C}_A \text{V}_A}{\text{V}_B}$$

Alternative

1 mole of A is equivalent to 1 mole of B

$0.137 \text{ mol dm}^{-3}$ of A will neutralize 25 cm^3 of B

$$\begin{aligned} \text{Conc. of B (C}_B\text{)} &= \frac{0.137 \times \text{V}_A}{25} \\ &= b \text{ mol dm}^{-3} \end{aligned}$$

(iii) To calculate percentage purity of KOH in B

Molar mass of KOH = 56

Molar mass of KOH x molar concentration

$56 \times b (1) = z \text{ gdm}^{-3}$ say

Percentage purity = $\frac{\text{mass concentration} \times 100}{\text{mass of impure KOH}}$ or $\frac{z}{6.50} \times 100$

= y% say

C is FeSO₄; D is ZnCO₃

TEST	OBSERVATION	INFERENCE
(a)(i) C + H ₂ O + litmus paper	C dissolves, blue litmus paper turned red	soluble salt solution is acidic (Not acid)
(ii) Solution of C + NaOH(aq) in drops + excess	Dirty green ppt ppt insoluble	
(iii) Solution of C + BaCl _{2(aq)} + dil. HCl	white ppt ppt insoluble in Dil. HCl	

TEST	OBSERVATION	INFERENCE
(b)(i) D + heat	D changes from white to yellow and turned to white (1) on cooling colourless gas evolved, gas turned lime water milky	

(ii) Residue + dil. HCl	Effervescence/gas evolved, residue dissolves to form a colourless solution
(iii) Solution + NaOH from (b)(ii) in drops + excess	White gelatinous ppt, ppt soluble in excess or ppt dissolves
(iv) Solution (b)(ii) + NH ₃ (aq) + excess	White gelatinous ppt, ppt soluble in excess/

3. (a)(i) The dry gas does not contain $\text{H}_3\text{O}^+/\text{H}^+$ which changes the colour of litmus OR the ions - $\text{H}_3\text{O}^+/\text{H}^+/\text{OH}^-$ are produced only in presence of water.
- (ii) The reagent barium chloride can form a white ppt with other ions if the solution is neutral. Addition of a mineral acid removes the interfering ($\text{CO}_3^{2-}, \text{SO}_4^{2-}$) ions if they are present.
- (b) (i) Chloride, carbon (II) oxide, Hydrogen sulphide, hydrogen chloride/Names or formula.
- Sulphur (IV) oxide
- Nitrogen (IV) oxide, chlorine
- Ammonia (Any two)
- (ii) They should be prepared in the fume chamber
- (iii) I. For keeping solvents such as distilled water
- II. As a support for flask/beakers during heating

- (c) (i) $\text{Pb}(\text{NO}_3)_2$ - decomposes, brown fumes are given off, yellow residue / reddish brown.
- (ii) Greenish yellow gas evolved.

PAST QUESTION: CHEMISTRY I (PRACTICALS): NECO 2003 JUNE

Q.2 X is a mixture of NH₄Cl and CuCO₃. Carryout the following tests. Record your observations and inference.

2. X is a mixture of NH₄Cl and CuCO₃

TEST	OBSERVATION	INFERENCE
a. X + water + Filtered	Colourless Filtrate Green Residue	
	483	

b(i) Filtrate + dil HNO_3 (aq) + AgNO_3	White Ppt	Cl^- present, may be present,
(ii) Mixture in (i) above + NH_3 (aq) in Excess For "disappears"	Ppt dissolved/is soluble, No mark	<u>Cl⁻ confirmed</u> accept chloride ion.
(iii) Filtrate + NaOH (aq) and warm.	Colourless gas evolved with characteristic Choking/pungent/ Suffocating/irritating Smell, turned red litmus blue or formed dense white fumes with conc HCl vapour Note: No dense, no mark.	Gas is NH_3 From NH_4^+
c. Residue + dil HCl and warmed	Residue dissolved, liberating colourless, odourless gas which turned lime water Milky/turbid	Gas is CO_2 From CO_3^{2-} , or HCO_3^-
c(i) Solution from (c) above + NaOH (aq) in drops	Blue ppt	
In excess	Ppt remained/insoluble	Cu^{2+} present
C(ii) solution from (c) above + NH_3 (aq) in drops	Light (pale) blue ppt	
In excess	Ppt dissolved to give a deep blue solution. Do Not accept COLOURATION.	Cu^{2+} confirmed Accept Copper (II) ions

GENERAL HINTS

$$1. \text{ Amount of substance (mol)} = \frac{\text{Mass of substnace (g)}}{\text{Molar mass (gmol}^{-1}\text{)}}$$

OR

$$\text{Amount of substance (mol)} = \text{Molar conc. (Moldm}^{-3}\text{)} \times \text{Volume (dm}^3\text{)}$$

$$2. \text{ Molar Conc. (moldm}^{-3}\text{)} = \frac{\text{Mass Conc. (gdm}^{-3}\text{)}}{\text{Molar Mass (gmol}^{-1}\text{)}}$$

OR

$$\text{Molar Conc. (Mol dm}^{-3}\text{)} = \frac{\text{Amount (mol)}}{\text{Volume (dm}^3\text{)}} = \frac{n}{v}$$

$$3. \text{ Percentage Purity} = \frac{\text{Mass conc of pure substance}}{\text{Mass conc of impure substance}} \times \frac{100}{1}$$

$$4. \text{ Percentage Impurity} = \frac{\text{Mass conc of impurity}}{\text{Mass conc of impure substance}} \times \frac{100}{1}$$

OR

$$100 - \text{Percentage purity} = \% \text{ impurity}$$

$$5. \text{ Number of Water of crystallization in Na}_2\text{CO}_3 \cdot \text{XH}_2\text{O}$$

$$\frac{\text{Mass Conc. of hydrated Na}_2\text{CO}_3 \text{ in gdm}^{-3}}{\text{Mass Conc. of anhydrous Na}_2\text{CO}_3 \text{ in gdm}^{-3}} = \frac{\text{Molar mass of hydrated Na}_2\text{CO}_3}{\text{Molar mass of anhydrous Na}_2\text{CO}_3}$$

$$6. \text{ To find the percentage of water of crystallization in hydrated Na}_2\text{CO}_3$$

$$\frac{\text{Mass of water}}{\text{Mass of hydrated Na}_2\text{CO}_3} \times \frac{100}{1}$$

$$7. \text{ Mass of water} = \text{Mass of hydrated Na}_2\text{CO}_3 - \text{Mass of anhydrous Na}_2\text{CO}_3$$

$$8. \text{ Molar mass (gmol}^{-1}\text{)} = \frac{\text{Mass Conc. in gdm}^{-3}}{\text{Molar Conc. in moldm}^{-3}}$$

$$9. \text{ Mass (in gdm}^{-3}\text{)} = \text{Molar Mass (gmol}^{-1}\text{)} \times \text{Molar Conc (moldm}^{-3}\text{)}$$

$$10. \text{ Volumetric Equation : } \frac{C_A V_A}{C_B V_B} = \frac{\text{Amount of A}}{\text{Amount of B}}$$

NOTE: This is used only where C_A or C_B is given.

CHAPTER 11

NUCLEAR CHEMISTRY

Nuclear chemistry is concerned with chemical reaction in which the nuclei of atoms are involved. Isotopes are atoms of the same elements having the same atomic number but different atomic masses. For instance, the isotopes of:-

- i. Chlorine are: $^{35}_{17}\text{CL}$ and $^{37}_{17}\text{CL}$

ii. Carbon are: $^{12}_6\text{C}$ $^{13}_6\text{C}$ and $^{14}_6\text{C}$

iii. Oxygen are: $^{16}_8\text{O}$ $^{17}_8\text{O}$ $^{18}_8\text{O}$

These isotopes are stable because the number of both the neutrons and protons in each are about equal or they have a comparable ratio. The neutron - proton ratio of the

- i. Two isotopes of chlorine are 17:18 and 17:20 respectively
- ii. Three isotopes of carbon are 1:1, 7:6 and 4:3 respectively.
- iii. Three isotopes of oxygen are 1:1, 9:8 and 5:4 respectively.

11.1 RADIOACTIVITY, DISCOVERY, CHARACTERISTICS OF RADIOACTIVITY

In nuclear chemistry, it was observed that the neutron-proton ratio in the nuclei of some elements is not comparable. The neutron-proton ratio is disproportionate. The number of neutrons in the nucleus of an atom may be too many when compared to the number of protons and vice versa. When the situation occurs, the resulting isotope are said to be unstable and are called the unstable isotopes. Unstable isotopes are also known as the radioactive elements because they are radioactive in nature. Unstable or radioactive isotopes naturally want to become stable by emitting radiations or by disintegrating spontaneously to equalise or normalise their neutron - proton ratio. Unstable isotope is a radioactive isotope or element which spontaneously emits radiation. Radioactivity is therefore defined as the spontaneous emission of radiation by radioactive elements. Or it can be defined as the spontaneous disintegration of radioactive elements.

Radioactivity was discovered by chance in 1896 by Henry Becquerel, He placed some photographic plates wrapped in black paper in the drawer (Cardboard) where some uranium salt was kept. He noticed after some days that the photographic plates have been blackened.

After some controlled experiments, he concluded that the uranium salt were responsible for emitting an unknown radiation which blackened the film. The discovery was very similar to the accidental discovery of x-ray by Roentgen a year before. In 1898, Marie and Pierre Curie discovered that thorium, Polonium and Radium also produced radiation. The invisible rays produced by the radioactive elements are called RADIATION.

A great quantity of energy known as the nuclear energy is released during radioactivity. Radioactive elements emits radiations. The rate at which each radioactive element emits radiation is neither temperature nor pressure dependent. The invisible ray (radiation) affects photographic plates like the visible light rays, but unlike the visible light rays, the invisible rays can pass through opaque material. The invisible rays can cause gases to ionize.

11.2 TYPES OF RADIATION AND THEIR DETECTION

Radioactive elements emit or produce three types of radiation which can be separated and distinguished by their behaviour in an electrostatic field.

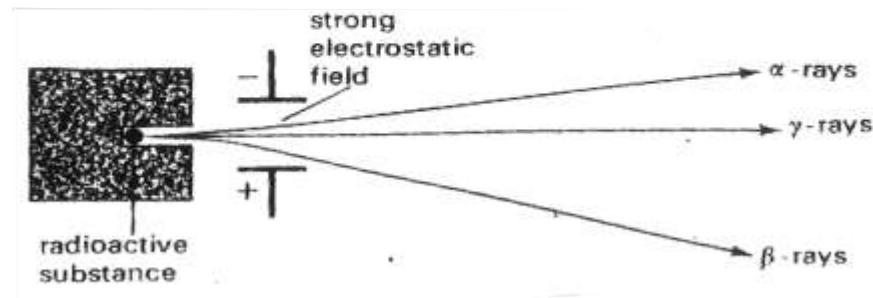


Fig. 11.1 Types of radiation

The three radiations are:

- i. Alpha rays or Alpha particles (α -rays)
- ii. Beta rays or Beta particles (β -rays).
- iii. Gamma rays (γ -rays)

ALPHA RAYS (α -rays): These rays are flow of fast-moving positively charged particles. α -Particle is a helium nucleus (${}_{2}^{4}\text{He}^{2+}$). It has +2 charges. α -particles are deflected slightly towards the negative plate in an electric or electromagnetic field. They travel few centimeters in air and when they collide with gas molecule can produce the ionization of the gas. α - particle is stopped or absorbed by a thin piece of paper.

BETA RAYS (β -rays): They consists of a fast moving flow of electrons. Each particle is represented with the symbol β^- . It has a mass number of zero and a -1 charge. β^- - particle can also produce ionization in gas. They have higher penetration power than α -rays. They can penetrate a thin sheet of metal like aluminum foil. They are deflected towards the positive plate in an electric field.

GAMA RAYS (γ -rays): γ -rays are not particles, they are electromagnetic waves. They resemble x-rays and visible light rays but have very short wavelengths. γ -ray has the least ionization power of the three radiations. They have no charge and are not affected by an electromagnetic field. They have the most penetrating power. They can pass through an iron or a lead block.

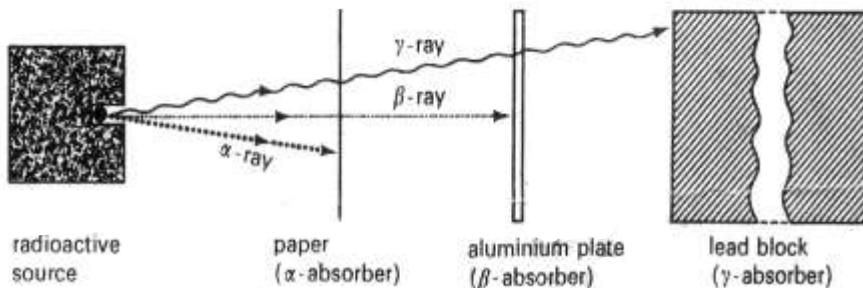


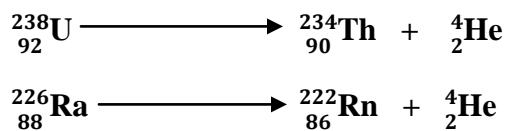
Fig. 11.2

Radiations produced by radioactive substance have some important scientific uses but sometimes they can be very harmful. It is therefore very important to detect and measure radioactions. The three most common methods used in detecting and measuring radioation are:

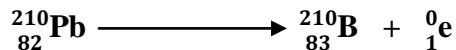
- i. Geiger - Muller counter
- ii. Scintillation counter
- iii. Diffusions cloud chamber

EFFECT OF α , β and γ emission (RADIOACTIVE DECAY):

When an atom loses an α -particle (He its atomic number decreases by 2 and the mass number decreases by 4. Uranium 238 losses an α - particle to become an atom of thorium - 234.



A radioactive isotope of lead can lose a β - particle to produce an atom of bismuth.



These types of change continue till a stable atomic nucleus is formed.

Emission of γ -rays do not change the mass number or atomic number of an element.

11.3 HALF-LIFE, $t_{1/2}$, AS A MEASURE OF STABILITY OF THE NUCLEUS

Half life is the rate of decay. The decay of an isotope can be measured by its half-life. Half-life is the time taken for half of the total mass of the isotope to decay. No method has been found to change the rate of the radioactive decay of an atom.

Large variation in temperature and pressure will not speed up or slow down the process. Half-life, ($t^{1/2}$) is constant for any particular isotope. It does not depend on the mass involved. Uranium has the half-life of 1620 years. Thorium has the half-life of 24 days. If 10g of a sample of the isotope (thorium) was taken, after 24 days, only 5g would remain. In the next 24 days, this 5g of thorium would have decayed by half (ie only 2.5g would remain). Therefore the greater the half-life of a radioactive element, the greater the stability of that element or isotope. When the change in percentage mass of a radioactive isotope is plotted against time in days, the curve shows an exponential decay which is a characteristic of first order reaction.

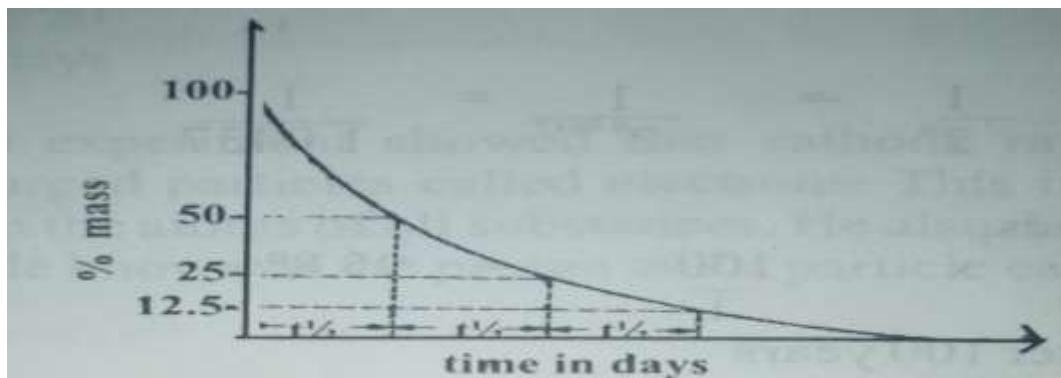


Fig. 11.3: Decay curve for radioisotope

Question 1

If 10g of an unstable isotope with half-life of 2 days is left for 10 days, what is the mass of the isotope remaining.

Answer

Using derived nuclear reaction equation we have

$$\frac{\text{original mass}}{2^n} = \text{residual mass}$$

Where $n = \frac{\text{number of activity}}{\text{half-life}} = \frac{\text{exposure time}}{\text{half-life}}$

Since exposure time is 10 days and half-life is 2 days

$$n = \frac{10\text{day}}{2\text{ days}} = 5$$

$$\therefore \frac{\text{Original mass}}{2^n} = \text{residual mas, Rm}$$

$$Rm = \frac{10g}{2^5} = \frac{10g}{3^2} = 0.3125g$$

Question 2

Isotope of an element, ^{226}R has half-life of 1620 years. What percentage of it remains after 100 years?

Answer

Number of activity, $n = \frac{\text{exposure (decay) time}}{\text{half-life}}$

$$= \frac{100 \text{ yrs}}{1620 \text{ yrs}} = 0.0617$$

$$\text{Fraction remaining} = \frac{1}{2^n} = \frac{1}{2^{0.0617}} = \frac{1}{1.0437}$$

In percentage this becomes,

$$\frac{1}{1.0437} \times \frac{100}{1} = 95.8\%$$

95.8% of R remains after 100 years

Question 3

A radioactive material has half-life of 20 days. How many days will it take to reduce the nuclei Y to ${}^{1/32}\text{Y}$ using the derived nuclear reaction equation.

Answer

Using the derived nuclear equation we have

$$\frac{\text{Original mass}}{2^n} = \text{Residual mass}$$

$$n = \frac{\text{exposure time (Duration of material)}}{\text{half-life of material } (t^{1/2})}$$

Let the exposure time be X, $t^{1/2} = 20$ days

$$\therefore n = \frac{X}{20}$$

Original mass = Y and Residual mass = ${}^{1/32}\text{Y}$

$$\therefore \frac{Y}{2^n} = \frac{Y}{32}$$

$$2^n \times Y = 32 \times Y$$

Divide both sides by Y

$$\therefore 2^n = 32$$

$$2^n = 2^5$$

$$\therefore n = 5$$

$$\therefore 5 = \frac{x}{20}$$

$$\therefore X = 5 \times 20$$

$$= 100 \text{ days}$$

Rutherford's experiment showed that cathode rays consist of smallest negatively charged particles called electrons. This indicated that electrons were present in the atoms of all substances. He also discovered the positively charged particle known as the proton and a particle called neutron which had no charge.

11.4 NUCLEAR REACTIONS

The nucleus of a radioactive element disintegrates spontaneously because it is unstable. The stability of an atomic nucleus has close relationship to its neutron-proton ratio. The neutron-proton ratio for stable isotopes are from 1 to 1.5, but isotopes with neutron-proton ratio less than 1 or greater than 1.5 are unstable and undergo radioactivity. The naturally occurring light isotopes (elements) have stable nuclei and are not radioactive. Potassium - 40 and Carbon - 14 are the exceptions. They are the light isotopes but are radioactive in nature.

The notable nuclear reactions we shall discuss are:

i. Natural radioactivity

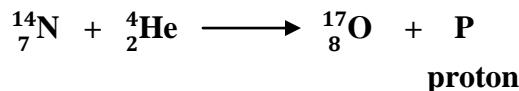
ii. Artificial radioactivity

iii. Nuclear fission

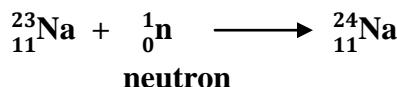
iv. Nuclear fusion

i. **Natural radioactivity** is the spontaneous disintegration of radioactive elements due to the instability of their nuclei. It is spontaneous and uncontrollable.

- ii. **Artificial radioactivity** is the nuclear transmutes (changes) brought about by bombardment of an isotope with either the α -particle or fast moving neutron. For instance, a nitrogen isotope can be transmuted into an isotope of oxygen by bombarding it with an α -rays.



Sodium -23 is transmuted to a radioactive sodium -24 by neutron bombardment



iii. Nuclear Fission:

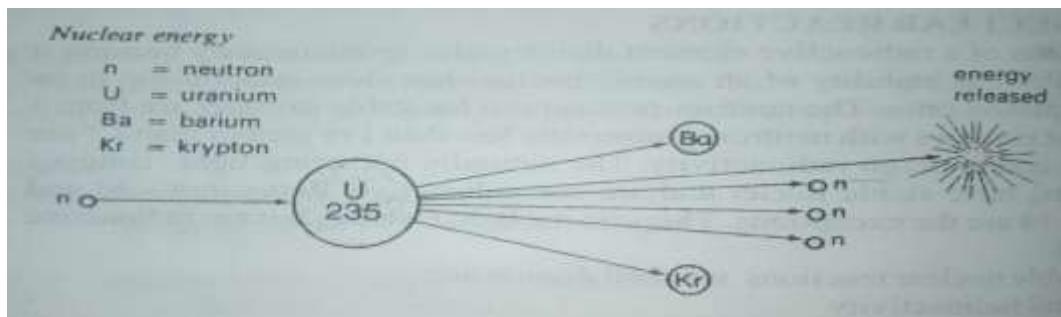
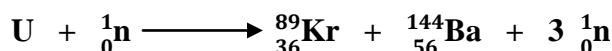


Fig. 11.4

When a nuclear reaction takes place in which a heavy nucleus is split into two or more atoms of similar size. The process is known as nuclear fission. For example when Uranium - 235 is bombarded with a neutron, it splits into krypton and Barium with a release of 3



The neutrons formed collide with more uranium nuclei. The fission process speeds up rapidly becoming a chain reaction. The large quantity of energy released from the chain reaction involving the fission of uranium -235 has been used in the atomic bomb. The energy release is calculated using the Einsteins equation.

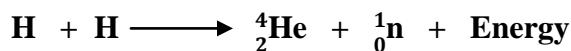
$$E = mc^2$$

Where E = energy released

m = mass of the isotope before and after the nuclear fission.

C = velocity of light

- iv. **Nuclear fission:** is the combining of two or more light nuclei to form a heavier nucleus. Greater energy is released during nuclear fission than during nuclear fission. In nuclear fission, the mass of the resulting nucleus is less than the total mass of the two or three combining nuclei. The loss in mass is equivalent to the energy released which accordance with the Einstein's equation, $E = mc^2$.



The sun and the stars obtained their energy from nuclear fission reaction.

11.5 DIFFERENCE BETWEEN NUCLEAR AND CHEMICAL REACTIONS

	Nuclear Reaction	Chemical Reaction
1.	It Occurs in unstable isotopes	It occurs in stable isotopes (elements).
2.	Invisible radiation are emitted during nuclear reaction	Invisible radiation are not emitted during chemical reaction.
3.	It occurs in the nucleus of atom	It does not occur in the nucleus of atom
4.	No donation and acceptance of electrons are involved because electron is not found in the nucleus.	Atoms involved in chemical reaction, donate and accept electrons.
5.	No bond is formed	Different types of bonds are formed
6.	Nuclear energy is generated	Chemical energy is generated.

Uses of radioactivity

- i. Cancer is sometimes cured by occasional exposure of the cancerous cells to the radiation known as the gamma-rays.
- ii. It is used as a tracer technique.

Exercises

1. If an atom of Uranium ${}_{92}^{235}\text{U}$ has a half-life of 1620 years, what percentage of it remains after 100 years?
2. Define these terms: (a) Half-life (b) Unstable isotope

(c) Radioactivity

3. What is the effect of emission of:
(a) α - particle? (b) β - particle?
4. What methods can be used to detect and measure radiations?
5. What is the difference between nuclear fission and nuclear fussion?
6. What are the differences between nuclear and chemical reactions?
7. If a radioactive material has a half-life of 14 days, what fraction of the original nuclei will remain after 42 days?

PERIODIC TABLE

GROUP		Atomic number	Symbol of element	Name of element	Relative atomic mass (approximate)																	
PERIOD						3	4	5	6	7	0											
1	H	1	Li	Lithium	6.94																	
2	Mg	2	B	Boron	10.81																	
3	Na	3	N	Nitrogen	14.01																	
4	K	4	C	Carbon	12.01																	
5	Rb	5	O	Oxygen	16.00																	
6	Cs	6	F	Fluorine	19.00																	
7	Ra	7	Ne	Neon	20.18																	
	Ac	8	He	Helium	4.00																	
Transition Elements																						
	Sc	9	V	Vanadium	50.94																	
	Ti	10	Cr	Chromium	51.99																	
	Cr	11	Mn	Manganese	54.94																	
	Mn	12	Fe	Iron	55.85																	
	Fe	13	Co	Cobalt	58.93																	
	Co	14	Ni	Nickel	58.73																	
	Ni	15	Pd	Palladium	69.73																	
	Pd	16	Au	Gold	196.97																	
	Au	17	Tl	Thallium	204.37																	
	Tl	18	Ag	Silver	107.87																	
	Ag	19	Pt	Platinum	195.08																	
	Pt	20	Hg	Mercury	200.59																	
	Hg	21	Pb	Lead	207.20																	
	Pb	22	Tl	Thallium	204.37																	
	Tl	23	Bi	Bismuth	208.98																	
	Bi	24	Po	Polonium	210.00																	
	Po	25	Al	Aluminum	26.98																	
	Al	26	Rn	Radon	222.00																	
	Rn	27	Ra	Radium	226.00																	
	Ra	28	Ac	Actinium	227.00																	

Lanthide Series		La	Lu	Ce	Pr	Nd	Eu	Sm	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	
		Lanthanum	173	Cerium	144	Neptunium	144	Europium	144	Terbium	159.5	Ytterbium	159.5	Thulium	173	Yttrium	173
1	Actinide Series	Ac	Ac	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lu/LW	

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